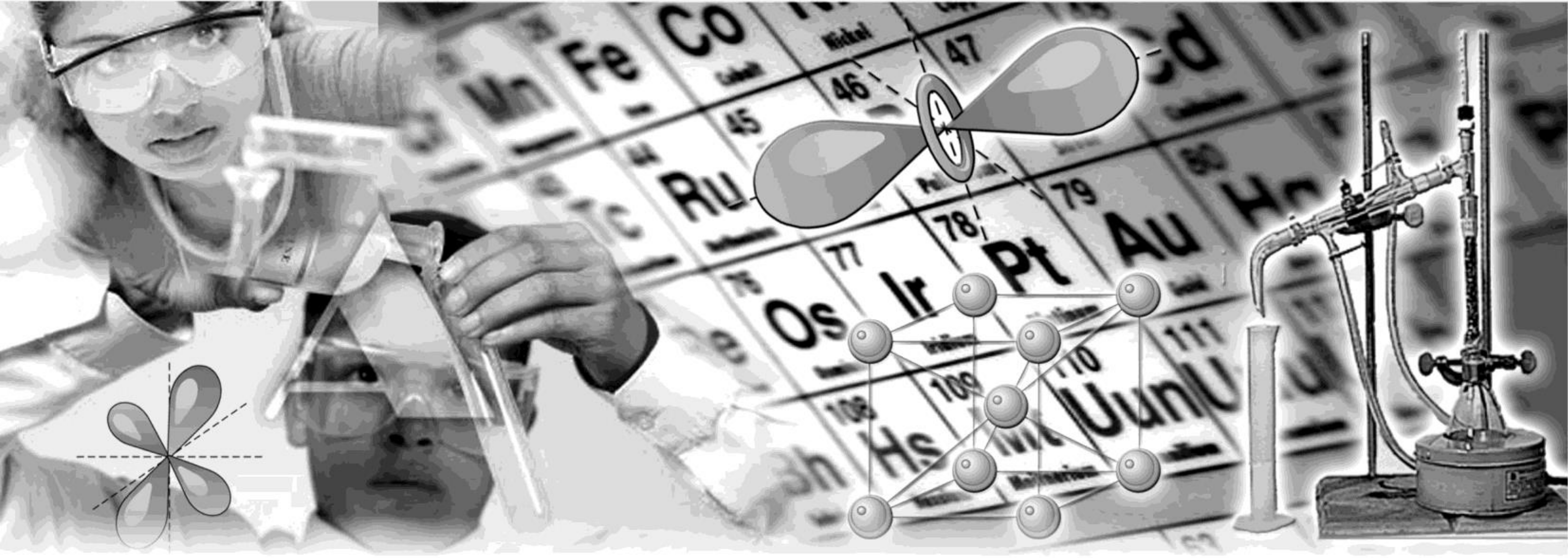


Chapter 15

Chemical Periodicity

Periodic table is an arrangement of elements with similar properties placed together. The **periodic table evolved largely as a result of experimental observations**.

Earlier attempt to classify elements

(1) **Dobereiner's law of triads** (1829) : It was the classification of elements into groups of three elements each with similar properties such that the atomic weight of the middle element was the arithmetic mean of the other two e.g. *Ca, Sr, Ba, Cl, Br, I* etc.

(2) **Telluric screw or Helix** (1862)**:** It was proposed by Chancourtois.

(3) **Newlands law of octaves** (1864) : It was an arrangement of elements in order of increasing atomic weights in which it was observed that every eighth element had properties similar to those of the first just like the eighth node of an octave of music.

(4) **Mendeleef's period law** (1869) : The first significant classification was given by Mendeleeff in the form of periodic table, commonly known as Mendeleeff's periodic table. His periodic table was based on periodic law, ''The physical and chemical properties of elements are periodic functions of their atomic weights.''

In Mendeleef's periodic table elements are arranged in order of their increasing atomic weights in such a way that elements with similar properties are placed in the same group. It consists of seven horizontal rows called periods. These are numbered from 1 to 7.

Mendeleef's original table consists of **8 vertical columns called groups**. These are numbered as **I, II III..... VIII**. However, 9th vertical column called Zero group was added with the discovery of inert gases. Except for group VIII and zero, each group is further divided into two sub-groups designated as *A* and *B*. Group VIII consists of 9 elements arranged in three sets each containing three elements.

(5) **Modern Periodic Law :** The recent work has established **that the fundamental property of an atom is atomic** number and not atomic weight. Therefore, atomic number is taken as the basis of the classification of the elements. The modern periodic law was given by **Moseley, it may be stated as** : ''The properties of elements are periodic functions of their atomic number''.

When atomic number is taken as the basis for classification of elements, many anomalies of Mendeleef's table disappear, such as the,

(i) ***Position of hydrogen*** : Dual behaviour of hydrogen is explained on the fact that it has one electron in its outermost orbit. When it loses its electron it gives H^+ and behaves like alkali metals and when it gains an electron it gives H^- and behaves like halogens. Thus, it resembles with both the alkali metals and the halogens.

(ii) ***Dissimilar elements placed together*** : The lengths of periods are determined by the arrangement of electrons in different orbits. The period ends on the completion of last orbits (last members always being the inert gas). Different periods contain 2, 8, 18 or 32 elements. Now out of the two elements which every long period adds to the group, one resembles the typical elements while the other does not. This gives rise to formation of subgroups. This explains the inclusion of dissimilar elements in the same group but different subgroups.

(iii) ***Position of rare earth elements*** : The electronic arrangement of rare earths can be written as 2, 8, 18, (18 + *x*), 9, 2 where *x* varies from 0 to 13, *i.e.,* from Lanthanum to Lutecium. The number of electrons in valency shell, in case of all the elements remain the same although the atomic number increases. Since they possess the same number of valency electrons, the chemical behaviour is also similar. This justifies their positions in the same group and in the same place of the periodic table.

(iv) ***Anomalous pairs of elements*** : Now the basis of classification is atomic number, therefore, this anomaly disappears as the elements occupy their normal position in the new periodic table.

(v) ***Position of isotopes*** : Since the isotopes of same element possess same atomic number they should occupy one and the same position in the periodic table.

(vi) ***Position of VIII group elements*** : In long periods 18 elements are to be distributed among 8 groups; 1 to 7 groups get 2 elements each and zero group accommodates inert elements, the rest three elements are placed at one place in a new group, known as VIII group. This lack of space justifies the induction of VIII group in the periodic table.

(vii) ***Transuranic elements*** : These elements form a series known as **actinide series**, it begins from actinium and ends at lawrencium (89–103). This series has been placed outside the periodic table. The electronic configuration of these elements can be written as 2, 8, 18, 32, (18 + *x*), 9, 2, where *x* varies from zero (for actinium) to 14 (for lawrencium). The number of valency electrons remains the same for all these elements although atomic number increases. Therefore, their chemical behaviour is similar. This justifies their position outside the periodic table at one place.

Classification of element on the basis of electronic configuration

According to Bohr the element have been grouped into four types

(1) **Inert gases :** These atom have been full fill outer most *s* and *p* subshell and having the configuration ns^2np^6 (Exception helium having $1s^2$) due to stable configuration these element do not show chemical activity under normal condition and thus termed as inert gases. Under special conditions, the higher members do form some compounds with other elements and hence, the name has been changed to **Noble gases**. Six element He, Ne, Ar, Kr, Xe and Rn belongs to this group.

(2) **Representative or normal elements :** Element in which atom have all shells complete except outermost shell which is incomplete. The number of electrons in the outermost shell varies from 1 to 7 *i.e.* the configuration of the outermost shell varies from ns^1 to ns^2np^5. These consist of some metals, all non metals and metalloid. These are the elements which are found in nature in abundance and active in nature. On account of this, these elements are called representative elements. All the three types of valencies are observed in the case of these elements. These elements have one of the following configuration in their outermost shell.

$$\underset{1}{ns^1} \quad \underset{2}{ns^2} \quad \underset{2}{ns^2np^1} \quad \underset{4}{ns^2np^2} \quad \underset{5}{ns^2np^3} \quad \underset{6}{ns^2np^4} \quad \underset{7}{ns^2np^5}$$

The number of electrons present in the outermost shell signify the group to which these element belong.

(3) **Transition elements :** In the atoms of these elements the outermost shell and the penultimate shell (Next to the outermost) are incomplete. These elements have the general configuration $(n-1)d^{1-9}\ ns^{0 \text{ or } 1 \text{ or } 2}$ either in the ground state or in excited state. These elements are present in IV^{th}, V^{th}, VI^{th} and VII^{th} period of periodic table and called transition elements. There are four transition series every series consists 9 elements each.

(4) **Inner transition elements :** Atoms of these elements have three outermost shell incomplete. The general configuration is $(n-2)f^{1-14}(n-1)d^{0 \text{ or } 1}ns^2$. There are two series of elements.

(i) Lanthanides or rare earths from $Ce(58)$ to $Lu(71)$.

(ii) Actinides from $Th(90)$ to $Lr(103)$

Each series consists of 14 elements *i.e.* in the lanthanides, $4f$ is gradually filled up while in actinides, $5f$ is gradually filled up.

The properties of these elements are similar to transition elements. The members of actinide series are radioactive and majority of them are not found in nature. The elements from atomic number of 93 onwards are called transuranic elements and have been discovered artificial means.

In this classification the element, *Zn*, *Cd* and *Hg* have not been included in any of the four groups of elements. The elements *Lu* and *Lr* of the inner transition group have $(n-2)f$ shell complete consisting 14 electron hence their inclusion in this group is not justified.

Extended or long form of periodic table

Modern periodic table is also called long form of the periodic table or Bohr's table. In this table, the elements are arranged in order of their increasing atomic number. It consists of 4 blocks (*s*, *p*, *d* and *f*), 18 groups numbered from 1 to 18 and 7 periods numbered from 1 to 7.

Blocks : The periodic table is divided into four main blocks (*s, p, d* and *f*) depending upon the subshell to which the valence electron enters into.

(1) Elements of group 1 and 2 constitute *s*-Block.

(2) Elements of group 13, 14, 15, 16, 17, 18 constitute *p*-Block

(3) Elements of group 3, 4, 5, 6, 7, 8, 9, 10, 11, 12 constitute *d*-Block

(4) The f-Block elements comprise two horizontal rows placed at the bottom of the periodic table to avoid its unnecessary expansion.

Elements of *s*- and *p*-blocks are called normal or **representative elements**, those of *d*-block are called **transition elements** while the *f*-block elements are called **inner transition elements**.

Groups : The 18 vertical columns are called groups. The elements belonging to a particular group is known as a family and is usually named after the first number. Apart from this some of the groups are given typical names as examplified beneath,

(1) Elements of group 1 are called **Alkali-Metals**.

(2) Elements of group 2 are called **Alkaline Earths**.

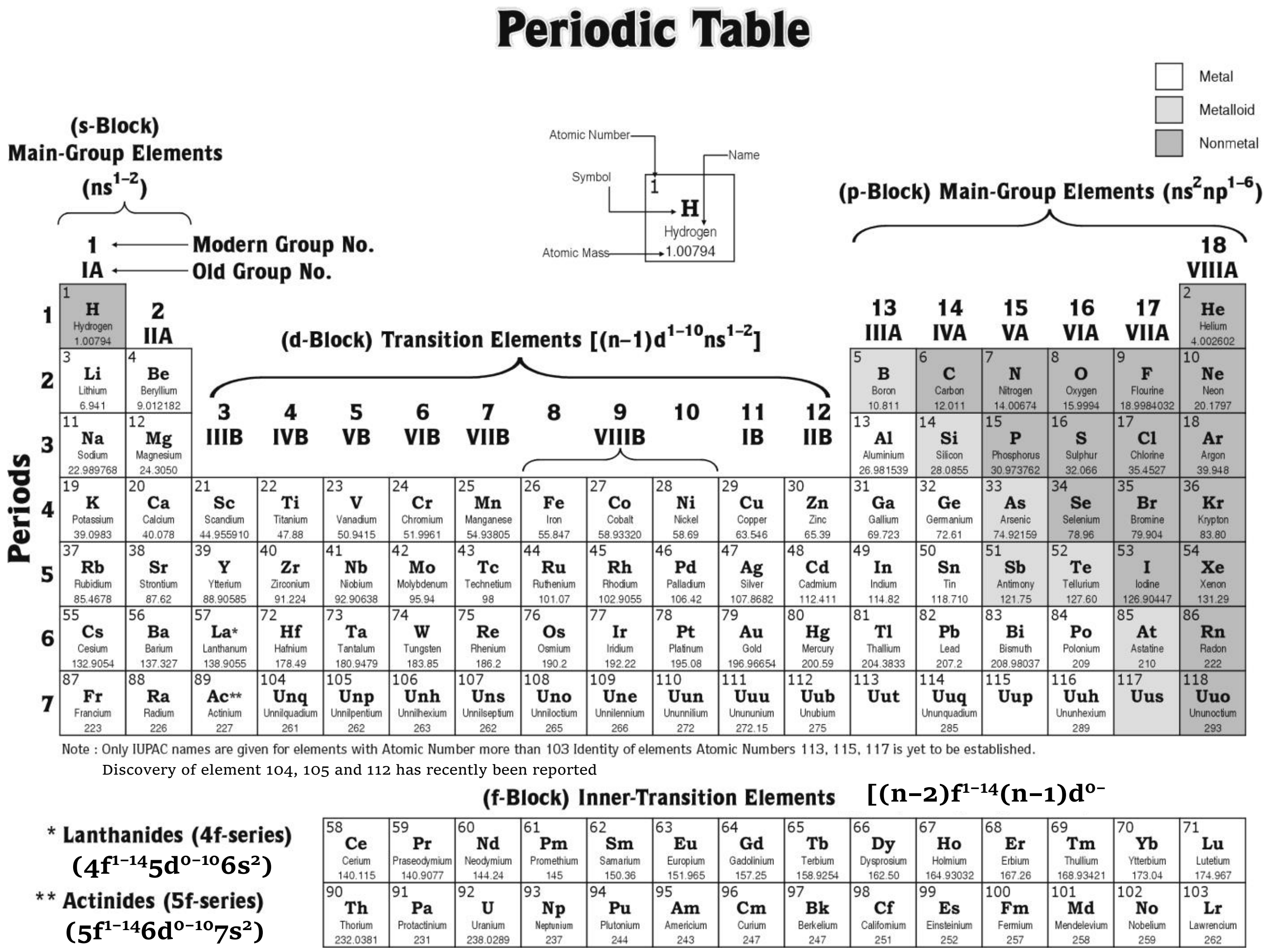
(3) Elements of group 3 are called **Pnicogens**.

(4) Elements of group 16 are called **Chalcogens**.

(5) Elements of group 17 are called **Halogens**.

(6) Elements of group 18 are called **Noble Gases or Aerogens**.

Periodic Table



Periods	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8	9 VIIIB	10	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 **H** Hydrogen 1.00794																	2 **He** Helium 4.002602
2	3 **Li** Lithium 6.941	4 **Be** Beryllium 9.012182											5 **B** Boron 10.811	6 **C** Carbon 12.011	7 **N** Nitrogen 14.00674	8 **O** Oxygen 15.9994	9 **F** Flourine 18.9984032	10 **Ne** Neon 20.1797
3	11 **Na** Sodium 22.989768	12 **Mg** Magnesium 24.3050											13 **Al** Aluminium 26.981539	14 **Si** Silicon 28.0855	15 **P** Phosphorus 30.973762	16 **S** Sulphur 32.066	17 **Cl** Chlorine 35.4527	18 **Ar** Argon 39.948
4	19 **K** Potassium 39.0983	20 **Ca** Calcium 40.078	21 **Sc** Scandium 44.955910	22 **Ti** Titanium 47.88	23 **V** Vanadium 50.9415	24 **Cr** Chromium 51.9961	25 **Mn** Manganese 54.93805	26 **Fe** Iron 55.847	27 **Co** Cobalt 58.93320	28 **Ni** Nickel 58.69	29 **Cu** Copper 63.546	30 **Zn** Zinc 65.39	31 **Ga** Gallium 69.723	32 **Ge** Germanium 72.61	33 **As** Arsenic 74.92159	34 **Se** Selenium 78.96	35 **Br** Bromine 79.904	36 **Kr** Krypton 83.80
5	37 **Rb** Rubidium 85.4678	38 **Sr** Strontium 87.62	39 **Y** Ytterium 88.90585	40 **Zr** Zirconium 91.224	41 **Nb** Niobium 92.90638	42 **Mo** Molybdenum 95.94	43 **Tc** Technetium 98	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.9055	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.8682	48 **Cd** Cadmium 112.411	49 **In** Indium 114.82	50 **Sn** Tin 118.710	51 **Sb** Antimony 121.75	52 **Te** Tellurium 127.60	53 **I** Iodine 126.90447	54 **Xe** Xenon 131.29
6	55 **Cs** Cesium 132.9054	56 **Ba** Barium 137.327	57 **La*** Lanthanum 138.9055	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.9479	74 **W** Tungsten 183.85	75 **Re** Rhenium 186.2	76 **Os** Osmium 190.2	77 **Ir** Iridium 192.22	78 **Pt** Platinum 195.08	79 **Au** Gold 196.96654	80 **Hg** Mercury 200.59	81 **Tl** Thallium 204.3833	82 **Pb** Lead 207.2	83 **Bi** Bismuth 208.98037	84 **Po** Polonium 209	85 **At** Astatine 210	86 **Rn** Radon 222
7	87 **Fr** Francium 223	88 **Ra** Radium 226	89 **Ac**** Actinium 227	104 **Unq** Unnilquadium 261	105 **Unp** Unnilpentium 262	106 **Unh** Unnilhexium 263	107 **Uns** Unnilseptium 262	108 **Uno** Unniloctium 265	109 **Une** Unnilennium 266	110 **Uun** Ununnilium 272	111 **Uuu** Unununium 272.15	112 **Uub** Unubium 275	113 **Uut**	114 **Uuq** Ununquadium 285	115 **Uup**	116 **Uuh** Ununhexium 289	117 **Uus**	118 **Uuo** Ununoctium 293

Note : Only IUPAC names are given for elements with Atomic Number more than 103 Identity of elements Atomic Numbers 113, 115, 117 is yet to be established.

Discovery of element 104, 105 and 112 has recently been reported

(f-Block) Inner-Transition Elements $[(n-2)f^{1-14}(n-1)d^{0-}$

*** Lanthanides (4f-series)** $(4f^{1-14}5d^{0-10}6s^2)$	58 **Ce** Cerium 140.115	59 **Pr** Praseodymium 140.9077	60 **Nd** Neodymium 144.24	61 **Pm** Promethium 145	62 **Sm** Samarium 150.36	63 **Eu** Europium 151.965	64 **Gd** Gadolinium 157.25	65 **Tb** Terbium 158.9254	66 **Dy** Dysprosium 162.50	67 **Ho** Holmium 164.93032	68 **Er** Erbium 167.26	69 **Tm** Thullium 168.93421	70 **Yb** Ytterbium 173.04	71 **Lu** Lutetium 174.967
**** Actinides (5f-series)** $(5f^{1-14}6d^{0-10}7s^2)$	90 **Th** Thorium 232.0381	91 **Pa** Protactinium 231	92 **U** Uranium 238.0289	93 **Np** Neptunium 237	94 **Pu** Plutonium 244	95 **Am** Americium 243	96 **Cm** Curium 247	97 **Bk** Berkelium 247	98 **Cf** Californium 251	99 **Es** Einsteinium 252	100 **Fm** Fermium 257	101 **Md** Mendelevium 258	102 **No** Nobelium 259	103 **Lr** Lawrencium 262

All the other groups are named after the first member of each group.

Periods : The horizontal rows are called periods. There are seven periods in the long form of the periodic table,

(1) Ist period $_1H \rightarrow _2He)$ contains 2 elements. It is the shortest period.

(2) 2nd period $(_3Li \rightarrow _{10}Ne)$ and 3rd period $(_{11}Na \rightarrow _{18}Ar)$ contains 8 elements each. These are short periods.

(3) 4th period $(_{19}K \rightarrow _{36}Kr)$ and 5th period $(_{37}Rb \rightarrow _{54}Xe)$ contains 18 elements each. These are long periods.

(4) 6th period $(_{55}Cs \rightarrow _{86}Ra)$ consists of 32 elements and is the longest period.

(5) 7th period starting with $_{87}Fr$ is incomplete and consists of 19 elements.

Periodicity in properties

"The repetition of similar electronic configuration after a definite period is the cause of periodicity of the properties of elements."

It can be explained with the help of electronic arrangement of elements. According to the modern views, the valency of an element is indicated by the number of electrons present in the outermost orbit. The chemical properties of elements are dependent on valency electrons. Variation in electronic arrangement leads to the variation in properties. After a definite interval, recurrence of similar electronic arrangement takes place when the number of valency electrons is the same. Thus, there is a regular gradation and repetition in the properties of elements.

Periodic properties are directly or indirectly related to their electronic configuration and show a regular gradation on moving from left to right in a period or from top to bottom in a group. Some period or from top to bottom in a group. Some important periodic properties are : **oxidation number, shielding effect, atomic radii, ionization energy, electron affinity, electro-negativity, valency, density, m.pt. and b.pt**.

The screening effect or shielding effect

A valence-electron in a multi-electron atom is attracted by the nucleus, and repelled by the electrons of inner-shells. The combined effect of this attractive and repulsive force acting on the valence-electron experiences less attraction from the nucleus. This is called ***shielding*** or ***screening effect***. The magnitude of the screening effect depends upon the number of inner electrons, *i.e.,* higher the number of inner electrons, greater shall be the value of screening effect. The screening effect constant is represented by the symbol 'σ' is determined by the ***Slater's rules.***

The magnitude of screening constant in the case of *s*- and *p*- block elements increases in a period as well as in a group as the atomic number increases.

Effective nuclear charge

Due to screening effect the valency electron experiences less attraction towards nucleus. This brings decrease in the nuclear charge (*Z*) actually present on the nucleus. The reduced nuclear charge is termed ***effective nuclear charge*** and is represented by Z^*. It is related to actual nuclear charge (*Z*) by the following formula,

$Z^* = (Z - \sigma)$ where σ is screening constant.

It is observed that ***magnitude of effective nuclear charge increases in a period when we move from left to right.***

In a subgroup of normal elements the magnitude of effective nuclear charge remains almost the same.

Covalent, Ionic and Vander waal's radii

The radius of an atom is the distance between the centre of its nucleus and electrons in the last orbit. However, according to quantum mechanics, there is no certainty about the exact position of electrons at any time. Theoretically, an electron, at one time, may be very close to the nucleus while at other time it may be far away from the nucleus. In spite of these limitations we need to have some operational definition of the term atomic radius. There are three operational concepts of atomic radius.

(1) If the bonding is covalent, the radius is called a *covalent radius*.

(2) If the bonding is ionic, the radius is called *ionic radius*.

(3) If the two atoms are not bonded by a chemical bond (as in noble gases) the radius is called *vander Waal's radius*.

(1) **Covalent radius** : *It is half of the distance between the nuclei of two like atoms bonded together by a single bond*. Thus covalent radius of carbon in a compound having *C* – *C* single bond can be determined by dividing the bond length by 2, *i.e.,*

$$r_c = \frac{C-C}{2} \qquad \therefore\ C-C = 2r_c \qquad \text{or} \qquad r_c + r_c$$

where, r_c is the single bond covalent radii (SBCR) of carbon. However, if atoms forming the covalent bond are different *i.e.,* one is more electronegative than the other then the atomic radius is determined by the relation

$$A-B=r_A+r_B-0.009\,(\chi_A-\chi_B)$$

where χ_A and χ_B are electronegativities of the atoms A and B respectively. This relation was given by ***Stevenson*** in 1941.

In a given period, atomic radius generally decreases from left to right and thus in any period, alkali metal is the largest and halogen is the smallest atom. For example, in second period elements the covalent radii decrease from Li to F.

$_3Li$	$_4Be$	$_5B$	$_6C$	$_7N$	$_8O$	$_9F$	$_{10}Ne$
1.23	0.89	0.80	0.77	0.74	0.74	0.72	1.6

The decrease in size along a period is due to the effect of successive increasing nuclear charge without addition of a new shell, *i.e.,* in each element of a given period a new electron is added in the same principal quantum number. For example, in the second period the nuclear charge increases from + 3 in Li to + 9 in F. The increased nuclear charge attracts the electrons more strongly to the nucleus and thus decreases the size of the atom. In case of noble gases, the atomic radii are only the vander Waal's radii which are naturally higher than the covalent radii of other elements.

In a given group, Atomic radius generally increases as one moves from top to bottom, e.g., in group 1 atomic size increases steadily from lithium to cesium, *i.e.* $r_{Cs}>r_{Rb}>r_K>r_{Na}>r_{Li}$

The increase in size on descending a group is due to addition of extra shell which outweighs the effect of increased nuclear charge. Remember that He and Fr are smallest and largest atom respectively.

(2) **Ionic radius** : *It is the effective distance from the nucleus of an ion upto which it has its influence on its electron cloud.*

A cation is always much smaller than the corresponding atom. Further, more the number of electrons removed smaller will be the size of the resulting positive ion. For example, $\underset{1.26}{r_{Fe}} > \underset{0.76}{r_{Fe^{2+}}} > \underset{0.64}{r_{Fe^{3+}}}$.

This is due to following two factors

(i) A cation formed by the loss of electrons may result in the complete disappearance of the outer shell and since the remaining inner shells do not extend so far in space, the cation is much smaller than the metal atom. For example,

$$\underset{(2,\,8,\,1)}{\text{Sodium atom }(Na)} \rightarrow \underset{(2,\,8)}{\text{Sodium ion }(Na^+)}$$

(ii) Whenever a cation is formed, the ratio of nuclear charge to the number of electrons (Z/e ratio) is increased with the result the effective nuclear charge is increased and the electrons are pulled towards the nucleus. Consequently, the cation becomes smaller. An anion is always larger than the corresponding atom. For example, Atomic radius of I = 1.23 ; Ionic radii of I^- = 2.16

This is again due to following two factors : (a) Since in the formation of an anion, one or more electrons are added, the electron cloud expands and the ionic size increases. (b) In the formation of anion, the effective nuclear charge decreases with the result the electrons get away from the nucleus and thus the anion becomes larger than the corresponding atom.

In any particular group, the ions of elements increase in size on moving from top to bottom.

In case of *isoelectronic ions* (*ions having same number of electrons but different nuclear charge*); the greater the nuclear charge, the greater is the attraction for electrons and smaller is ionic radius. Hence size of such ions decreases. Which is as follows

$$\underrightarrow{C^{-4}>N^{-3}>O^{-2}>F^{-1}>Ne>Na^{+}>Mg^{+2}>Al^{+3}>Si^{+4}>P^{+5}>S^{+6}>Cl^{+7}}$$

Size of ions (in decreasing order)

(3) **Vander Waal's radius** : *It is one–half of the distance between the nuclei of two adjacent atoms belonging to two neighbouring molecules of an element in the solid state.*

The covalent radius is always smaller than the corresponding vander Waal's radius. This is because of the fact that in the formation of a chemical bond, the two atoms have to come closer to each other. This also explains *why covalent bonds are much stronger than the vander Waal's forces.* It is important to note that since the noble gases ordinarily do not form any covalent bond, in crystals of noble gases, no chemical forces are operating between the atoms. Hence the vander Waal's forces are the only attractive forces in these cases. In other words, *the vander Waal's radii constitute the* *atomic radii of noble gases* and since vander Waal's radii are larger than covalent radii, *atomic radii of noble gases are largest in their respective periods* (anomaly).

Ionisation potential or ionisation energy

The electrons in an atom are attracted by the nucleus. When an electron is to be removed then work is done against this nuclear attraction. In other words energy is required to remove an electron from an atom. To understand the details of chemical behaviour of an element we must have an indication of the energy with which an atom binds its electrons. This is obtained by the measurement of ionisation potential or ionisation energy. It may be defined as **the** *energy required to remove an electron from the outermost orbit of an isolated gaseous atom in its ground state*. It is expressed in electron volts (eV) or kilo calories per gram atom. In an atom, the energy required to remove first electron from a gaseous atom is called first ionisation energy. The energy required to remove one electron from a unipositive ion to form a bipositive ion is called second ionisation energy. Second ionisation energy is higher than the first. The reason is that in unipositive ion left after the removal of one electron from the atom, the electrons are more firmly bound to the nucleus than in the atom. Hence more energy is needed to remove the second electron.

$$A \rightarrow A^{+} + e^{-1} \quad \text{(First } I.E.\text{)}$$

$$A^{+} \rightarrow A^{+2} + e^{-1} \quad \text{(Second } I.E.\text{)}$$

$$A^{+2} \rightarrow A^{+3} + e^{-1} \quad \text{(Third } I.E.\text{)}$$

Similarly, third ionisation energy is even more than second ionisation energy.

(1) **Variation of ionisation energy in periodic table**

(i) *Ionisation energy decreases in a group as the atomic number increases*. It is based on the fact that as we move down a group, the size of atom increases, and the outer electrons become farther away from the nucleus thus reducing the force of attraction and hence ionisation energy decreases.

Li	*Na*	*K*	*Rb*	*Cs*
5.4 *eV*	1. *eV*	4.3 *eV*	4.2 *eV*	3.9 *eV*

(ii) *Ionisation energy increases along a period with increase in atomic number*. This is due to the size of atom since it decreases along a period and outer electrons are most strongly attracted by the nucleus and hence more energy is required to remove the electron.

Li	*Be*	*B*	*C*	*N*	*O*	*F*	*Ne*
5.4*eV*	9.3*eV*	8.3*eV*	11.3*eV*	14.6 *eV*	13.6*eV*	17.0*eV*	21.6*eV*

(iii) The ionisation energies of inert gases are greater than that of their immediate neighbour. It is due to their complete octet ns^2p^6 configuration which is highly stable. Therefore, it is very difficult to remove an electron from the outermost orbit of an inert gas.

(2) **Factors affecting the value of ionisation energy**

(i) *Size of atom* : With an increase in atomic size, the ionisation potential is reduced, since the distance of the outermost electron from the nucleus increases and hence the force of attraction decreases.

(ii) *The charge on the nucleus* : With an increase in the nuclear charge, there is an increase in force of attraction of nucleus for electrons making the removal of the electrons more difficult. Thus an increase in nuclear charge increases the ionisation potential.

(iii) *The shielding or screening effect of inner shells* : The valence electrons in a multi–electron atom are pulled by the nucleus but are repelled by the electrons of the inner shells. The valence electrons, therefore, do not experience the total pull of the nucleus. Instead the total pull of the nucleus is reduced by the electrons in inner shells. This effect of reducing the force of attraction of nucleus by the inner shells is called *screening effect*. This effect is exhibited maximum by s^2p^6 (the most stable) shell. Therefore, the ionisation energy of *K* is much less than *Cu*, however, both have one electron in their fourth shell.

$$K^{19} = 1s^2, 2s^2p^6, 3s^2p^6, 4s^1$$

$$Cu^{29} = 1s^2, 2s^2p^6, 3s^2p^6d^{10}, 4s^1$$

The ionisation energy of *K* is 4.33 *eV* while that of *Cu* is 7.72 *eV*. This is due to a large screening effect of s^2p^6, penultimate orbit in K while $s^2p^6d^{10}$, penultimate orbit in *Cu* which exhibit little screening effect.

(iv) ***Type of electrons involved*** : Ionisation energy also depends upon the type, *i.e., s, p, d* or *f*, electrons which are to be removed, *s* –electrons are closer to the nucleus and are more tightly held as compared to *p, d* or *f* electrons. Hence, ionisation energy decreases in the order of $s > p > d > f$ orbitals.

(v) ***Completely filled or half filled sub-shells*** **:** According to Hund's rule, completely filled or half filled orbitals are more stable. Therefore, it is comparatively difficult to remove the electrons from these shells. The ionisation energy of Be (9.3 eV) is more than B (8.3 eV) because Be has $2s^2$ configuration of the outermost orbit which is fully filled. Similarly, nitrogen (14.6 eV) has more ionisation energy than oxygen (13.6 eV) because nitrogen has outermost shell configuration as $2s^2p^3$ in which p shell is half filled and is more stable. Similarly, ionisation energy of Mg is more than Al and that of P is more than S.

(3) **Relative ionisation energies**

- IE_1 and IE_2 of the 2nd period elements

IE_1 : $Li < B < Be < C < O < N < F < Ne$

IE_2 : $Be < C < B < N < F < O < Ne < Li$

- IE_1 and IE_2 of the 3rd period elements

IE_1 : $Na < Al < Mg < Si < S < P < Cl < Ar$

IE_2 : $Mg < Si < Al < P < S < Cl < Ar < Na$

- IE_1 of elements with very high values

$Cl < H < O < Kr < N < Ar < F < Ne < He$

(4) **Importance of ionisation energy**

(i) Lower is the ionisation potential of an element, more would be its reducing power and also reactivity.

(ii) It gives rough estimate about the basic character of the elements.

(iii) The relative values of ionisation potential and electron affinity of two elements are related to the nature of bond formed during their combination.

(iv) The ionisation potentials provide an indication about the number of valence electron (s) in an atom; the abnormally high value indicates that the electron removed is other than the valence electron.

For example, IE_1, IE_2 and IE_3 values are 5.39, 75.62 and 122.42 eV. Since the values shows sudden jump, it indicates that the number of valence electron in its atom is one. Similarly, values of IE_1, IE_2 and IE_3 as 9.32, 18.21 and 153.85 eV indicate that the number of valence electrons in its atom is two.

Electron affinity

Those atoms whose nuclear forces are not completely screened by electronic shells, offer attraction for electrons. Such atoms capture electrons if these are available with in their effective fields to neutralise the electrostatic forces of the nucleus. Energy is always liberated whenever there is a force of attraction offered by an atom or ion, and this energy is called electron affinity. This may be defined as, ''*the energy released when an extra electron is added to a neutral gaseous atom*''.

When first electron is added in a neutral atom then some energy is released that is called first electron affinity but in case of second electron affinity energy will be absorbed due to electronic repulsion. Example as

(a) $O_{(g)} + e^- \to O^-_{(g)} + E_1 (\Delta H = -ve, \text{Exoenergic})$

(b) $O^-_{(g)} + e^- \to O^{-2} + E_2 \,(\Delta H = +ve, \text{Endoenergic})$

Thus, higher the energy released in the process of taking up an extra electron, the higher will be the electron affinity. Higher the value of electron affinity of an atom, the more is its tendency to change into anion. It is very difficult to determine the electron affinity experimentally. The values have been calculated on the basis of thermodynamic concepts. It is expressed in electron volts. The values of inert gases are assumed to be zero because they have stable ns^2p^6 configuration and unable to accept any electron. The values for alkali metals are between zero and one.

The electron affinities of Be, Mg and zero since they have complete ns^2 configuration which can not accommodate extra electron. Similarly, the values for N and P are very low because they also have completely half–filled p orbitals (ns^2p^3) and are more stable.

(1) **Factors affecting the value of electron affinity**

(i) ***Atomic size*** **:** The value of electron affinity decreases with the increase in the size of atom since the nuclear attraction decreases down a group as the atomic number increases. Its value increases as we move along a period since the size of atoms decreases along a period. The lower value of F than Cl is due to the very small size of F in which negative charge is highly concentrated and repels the incoming electron thereby reducing the force of attraction of the nucleus towards the adding electron and hence decreasing the electron affinity. Thus, chlorine has a highest value of electron affinity.

(ii) ***Nuclear charge*** **:** The value of electron affinity increases with increasing nuclear charge. Thus, its value increases with increase in nuclear charge along a period.

(iii) ***Screening or shielding effect*** **:** The value of electron affinity increases with the decrease in shielding effect of inner electrons. Besides, the value of electron affinity also depends to some extent upon the type of orbital in which electron is added. The value is greater when electron enters '*s*' orbital and decreases successively for *p*, *d* and *f* orbitals.

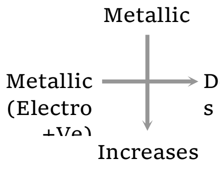
(2) **Importance of electron affinity :** Certain properties of the elements are predicted on the basis of values of electron affinity.

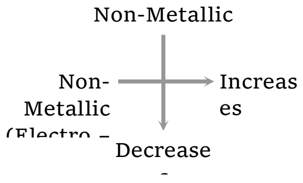
(i) The elements having high value high values of electron affinity are capable of accepting electron easily. They form anions and electrovalent compounds. These elements are electronegative in nature.

(ii) The elements having high values of electron affinity act as strong oxidising agents, for example, $F, Cl, Br, O, S,$ etc.

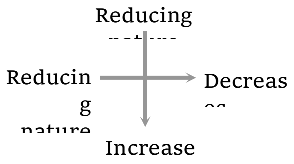
On the basis of the general trend of ionisation potential and electron affinity, the following properties can be predicted,

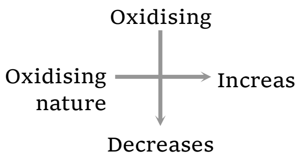
(a) Metallic nature decreases in a period while nonmetallic nature increases. Metallic nature increases in a group while non-metallic nature decreases. The arrow (↓) represents a group and (→) represents a period.



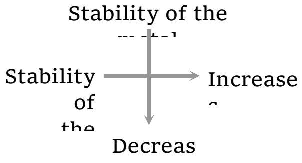


(b) Reducing nature decreases in a period while oxidising nature increases. The reducing nature increases in a group while oxidising nature decreases.



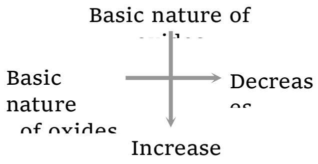


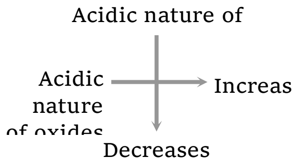
(c) Stability of metal increases while activity of the metal decreases in a period and in a group stability decreases while activity increases.



This trend is observed especially in IA, IIA and IIIA elements.

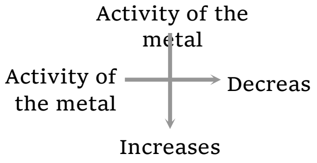
(d) The basic nature of the oxides decreases in a period while acidic nature increases. In a group, basic nature increases while acidic nature decrases.





Electronegativity

The tendency of an atom in a compound to attract a pair of bonded electrons towards itself is known as electronegativity of the atom. It is important to note that electron affinity and electronegativity both measure the electrons attracting power but the former refers to an isolated gaseous atom while the latter to an atom in a compound. Thus electron affinity is attraction for a single electron while electronegativity is for a pair of bonded electrons. Further electron affinity is energy while electronegativity is a tendency.



(1) **Factors affecting the value of electronegativity**

(i) The size of the atom.

(ii) Electronic configuration. Small atoms attract electrons more than the larger one and are therefore more electronegative. Secondly, atoms with *nearly* filled shell of electrons, will tend to have higher electronegativity than those sparsely occupied ones.

(2) **Variation of electronegativity in the periodic table**

(i) In a period, electronegativity increases from left to right. This is due to decrease in size and increase in nuclear charge. Thus the alkali metals possess the lowest value, while the halogens have the highest. Inert gases have zero electronegativity.

(ii) In a group, electronegativity decreases from top to bottom. This is due to increase in atomic size.

If an element exhibits various oxidation state, the atom in the higher oxidation state will be more negative due to greater attraction for the electron, *e.g., Sn* II (1.30) and *Sn* IV (1.90).

(3) Electronegativity may be expressed on the following three scales

(i) ***Mulliken's scale*** **:** Mulliken regarded electronegativity as the average value of ionization potential and electron affinity of an atom.

$$\text{Electronegativity} = \frac{\text{Ionization potential} + \text{Electron affinity}}{2}$$

(ii) ***Allred-Rochow scale*** **:** Allred and Rochow defined electronegativity as the electrostatic force exerted by the nucleus on the valence electrons. Thus $\chi = \frac{0.359\,Z}{r^2} + 0.744$ where Z is the effective nuclear charge and r is the covalent radius of the atom in Å.

(iii) ***Pauling scale*** **:** Pauling scale of electronegativity is the most widely used. It is based on excess bond energies. He determined electronegativity difference between the two atoms and then by assigning arbitrary values to few elements (*e.g.* 4.00 to fluorine, 2.5 to carbon and 2.1 to hydrogen), he calculated the electronegativity of the other elements. $\chi_A - \chi_B = 0.208\sqrt{\Delta E}$ where χ_A and χ_B are electronegativities of the atoms A and B respectively, the factor 0.208 arises from the conversion of kcal to electron volt (1 eV = 23.0 kcal/mole),

while $\Delta E = \text{Actual bond energy } - \sqrt{(E_{A-A} \times E_{B-B})}$

Pauling and Mulliken values of electronegativities are related as below χ (Pauling) = 0.34χ (Mulliken) – 0.2

(4) **Importance of electronegativity :** The following predictions can be made from value of electronegativity,

(i) Nature of the bond between two atoms can be predicted from the eelctronegativity difference of the two atoms.

(a) The difference $X_A - X_B = 0$, *i.e.,* $X_A = X_B$ the bond is purely covalent.

(b) The difference $X_A - X_B$ is small, *i.e.,* $X_A > X_B$, the bond is polar covalent.

(c) The difference $X_A - X_B$ is 1.7, the bond is 50% covalent and 50% ionic.

(d) The difference $X_A - X_B$ is very high, the bond is more ionic and less covalent. The molecule will be represented in such case as $BA\,(B^+A^-)$. Percentage ionic character may be calculated as,

Percentage of ionic character

$$= 16(X_A - X_B) + 3.5(X_A - X_B)^2$$

where X_A and X_B represents electronegativity of bonded atoms *A* and *B*.

This relation was given by A.L. Allerd (1961).

(ii) Greater the value of difference $(X_A - X_B)$ more stable will be the bond.

	$H-F$	$H-Cl$	$H-Br$	$H-I$
$(X_A - X_B)$	1.9	0.9	0.7	0.4

$$\xrightarrow{\hspace{8em}}$$

Stability decreases

Stability of compounds in which $X_A - X_B$ is very small are unstable in nature, $SiH_4(0.3), NCl_3(0.0)$, $PH_3(0)$, $AsH_3(0.1)$ are unstable.

(iii) $(X_O - X_A)$ difference predicts the nature of the oxides formed by the element *A*. X_O is the electronegativity of oxygen.

$X_O - X_A$ is large, the oxide shows basic nature, (*e.g.,* Na_2O).

$X_O - X_A$ is small, the oxide shows acidic nature, (*e.g.,* SO_2).

(iv) Ionic compounds having percentage ionic character less than 20% were found coloured, e.g.,

$AgCl$	$AgBr$	AgI	Ag_2S
22%	18%	11%	8%
White	Light yellow	Dark yellow	Black

Lesser the percentage ionic character, darker will be the colour.

Some other periodic properties

(1) **Atomic volume :** It is defined as the volume occupied by one gram atom of an element. Mathematically,

$$\text{Atomic volume} = \frac{\text{Gram atomic weight}}{\text{Density in solid state}}$$

Units of atomic volume are c.c./mole. Atomic volume signifies the volume occupied by one mole (Avogadro number) of atoms of the given element in solid state. Lower atomic volume generally leads to higher density, increased hardness and brittleness, higher melting and boiling points, less malleability and ductility.

(i) While descending a group, the atomic volume generally increases which is due to increase in the number of shells though the valence electrons in a given group remains constant.

(ii) While going left to right across a period the atomic volume first decreases to a minimum and then increases. Francium has the highest atomic volume and boron has lowest atomic volume.

(2) **Density :** The density of the elements in solid state varies periodically with their atomic numbers. At first, the density increases gradually in a period and becomes maximum somewhere for the central members and then starts decreasing afterwards gradually.

(3) **Melting and boiling points :** The melting points of the elements exhibit some periodicity with rise of atomic number. It is observed that elements with low values of atomic volumes have high melting points while elements with high values of atomic volumes have low melting points. In general, melting points of elements in any periodic at first increase and become maximum somewhere in the centre and thereafter begins to decreases.

Tungsten has the maximum melting point (3410°C) amongst metals and carbon has the maximum melting point (3727°C) amongst non-metals. Helium has the minimum melting point (–270°C). The metals, Cs, Ga and Hg are known in liquid state at 30°C.

The boiling points of the elements also show similar trends, however, the regularities are not so striking as noted in the case of melting points.

(4) **Oxidation state (Oxidation number, O.N.) :** *Oxidation number of an element in a compound is the total number of electrons it appears to have gained or lost* (negative and positive oxidation states respectively) *during the formation of that particular compound.*

Note : For detail see chapter redox reaction.

(5) **Magnetic properties :** Magnetic properties of matter depend on the properties of the individual atoms. A substance (atom, ion or compound) capable of being attracted into a magnetic field is known as **paramagnetic**. The paramagnetic substances have a net magnetic moment which in turn is due to the presence of unpaired electron(s) in atoms, ions or molecules. Since most of the transition metal ions have unpaired *d*–electrons, they show paramagnetic behaviour. The exceptions are $Sc^{3+}, Ti^{4+}, Zn^{2+}, Cu^{+}$, etc. which do not contain any unpaired electron and hence are diamagnetic.

On the other hand, a substance which is repelled by a magnetic field is known as ***diamagnetic***. Such substances do not have any net magnetic moment because they do not have any unpaired electron. Electrons determine the magnetic properties of matter in two ways,

- Each electron can be treated as a small sphere of negative charge spinning on its axis. The spinning of charge produces magnetic moment.
- An electron travelling in closed path around a nucleus will also produce magnetic moment just as does electric current travelling in a loop of wire.

The observed magnetic moment is therefore the sum of the two moments: the spin moment and the orbital moment. It is expressed in units called Bohr Magnetons (BM). In terms of *n* (number of unpaired electron), magnetic moment is given by the formula, $\mu = \sqrt{n(n+2)}$

Greater the number of unpaired electrons in a substance, the greater is the magnetic moment of the substance. The value of magnetic moment has been used to calculate the number of unpaired electrons in an ion. In some cases, even the structure of the molecule or complex is indicated by its magnetic moment.

Paramagnetism is generally measured by a simple device known as Guoy's balance which involves weighing the species in presence of a magnetic field.

Ferromagnetism is a special property observed in some substances in the solid state. Such *substances are strongly attracted to magnetic field and may retain the magnetic properties for some time even after the removal of the field.* The most common example is of *Fe* followed by *Co* and *Ni*.

(6) **Hydration and hydration energy**

(i) Hydration energy is the enthalpy change that accompanies the dissolving of 1 *mol* of gaseous ions in water.

(ii) Size of ions and its charge determines extent of hydration. Greater the charge smaller the size of the ion, greater the attraction for the lone pair of *O* of H_2O, hence greater the extent of hydration energy.

(a) Size of the hydration ion increases.

(b) Ionic mobility decreases *i.e.* heavier (hydrated) ions moves slower.

(7) **Acid-base-character of oxides**

(i) On moving across a period, the basic character of the oxides gradually changes first into amphoteric and finally into acidic character.

(ii) On moving down a group, reverse behaviour is observed i.e., from more acidic to more basic.

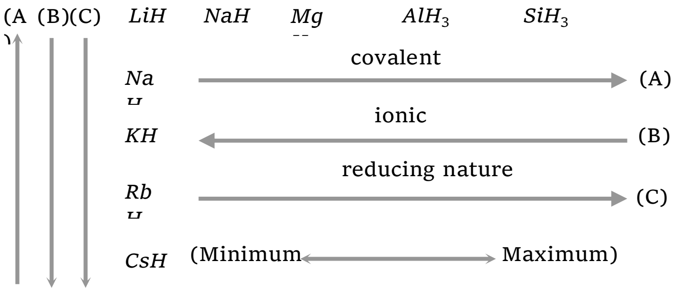
(iii) Stability of oxides decreases across a period.

(8) **Hydrides**

(i) Hydrogen combines with a number of other elements including metals and non-metals to form compounds called hydrides.

(ii) Covalent nature of hydrides increases across a period and decreases down the group.

(iii) Ionic hydride are better reducing agents than covalent hydride and reducing nature of hydride decreases across a period and increases down the group.

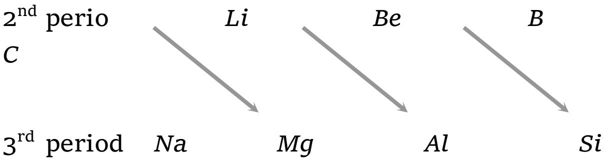


(iv) Covalent and ionic hydrides are classified as follows,

Diagonal relationship

Certain elements of 2nd period show similarity with their diagonal elements in the 3rd period as shown below :

Group 1 Group 2 Group 13 Group 14



Thus, *Li* resembles *Mg*, Be resembles *Al* and *B* resembles *Si*. This is called diagonal relationship and is due to the reason that these pairs of element have almost identical ionic radii and polarizing power (*i.e.* charge/size ratio). Element of second period are known as bridge elements.

***Anomalous behaviour of the first elements of a group*:** The first element of a group differs considerably from its congeners (i.e. the rest of the element of its group). This is due to *(i) small size (ii) high electronegativity and (iii) non availability of d-*orbitals for bonding. Anomalous behaviour is observed among the second row elements (*i.e. Li* to *F*).

Ordinary Thinking

Objective Questions

Extended or long form of periodic table

1. Which of the following statement is not correct for the element having electronic configuration $1s^2, 2s^2 p^6, 3s^1$
(a) It is a monovalent electropositive
(b) It forms basic oxide
(c) It is a non-metal
(d) It has low electron affinity

2. Which of these dose not reflect the periodicity of the elements **[UPSEAT 2001; BIT 1990; MP PMT 2001]**
(a) Bonding behaviour (b) Electronegativity
(c) Ionization energy (d) Neutron/proton ratio

3. If an atom has electronic configuration $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$, it will be placed in **[CBSE PMT 2002]**
(a) Second group (b) Third group
(c) Fifth group (d) Sixth group

4. All the s–block elements of the periodic table are placed in the groups ... **[Orissa JEE 2002]**
(a) IA and IIA (b) IIIA and IVA
(c) B sub groups (d) VA to VIIA

5. The electronic configuration of halogen is **[MP PET/PMT 1998; Pb. PMT 2001]**
(a) $ns^2 np^6$ (b) $ns^2 np^3$
(c) $ns^2 np^5$ (d) ns^2

6. Hydrogen by donating one electron forms H^+. In this property, it resembles with
(a) Transitional metals (b) Alkaline earth metals
(c) Alkali metals (d) Halogens

7. The tenth elements in the periodic table resembles with the **[CPMT 1988]**
(a) First period (b) Second period
(c) Fourth group (d) Ninth group

8. The element with quantum numbers $n = 2, l = 1, m = 1, s = -1/2$ has the following position in the periodic table
(a) Group VII-A, period II (b) Group 0, period II
(c) Group VII-A, period III (d)Group 0, period III

9. Who developed the long form of periodic table **[MP PET 1997]**
(a) Lothar Meyer (b) Niels Bohr
(c) Mendeleef (d) Moseley

10. 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 **[CBSE PMT 1995]**
(a) 33 (b) 34
(c) 31 (d) 49

11. In the periodic table, the element with atomic number 16 will be placed in the group **[MP PET/PMT 1998]**
(a) Third (b) Fourth
(c) Fifth (d) Sixth

12. The first element of rare–earth metals is **[AFMC 1992]**
(a) Cerium (b) Actinium
(c) Uranium (d) Lanthanum

13. The d-block elements consists mostly of **[MP PMT 1994]**
(a) Monovalent metals
(b) All non-metals
(c) Elements which generally form stoichiometric metal oxide
(d) Many metals with catalytic properties

14. "*The 6 properties of the elements are periodic function of their atomic numbers*." The statement was given by **[MNR 1995]**
(a) N. Bohr (b) J.W. Dobereiner
(c) D.I. Mendeleef (d) H.G.J. Moseley

15. The long form of periodic table has **[CPMT 1986; KCET 1998]**
(a) Eight horizontal rows and seven vertical columns
(b) Seven horizontal rows and eighteen vertical columns
(c) Seven horizontal rows and seven vertical columns
(d) Eight horizontal rows and eight vertical columns

16. The telluric helix was given by **[AFMC 1990]**
(a) De Chan Courtois (b) Newlands
(c) L. Meyer (d) Mendeleef

17. Which one of the following belongs to representative group of elements in the periodic table **[Kurukshetra CEE 1991]**
(a) Lanthanum (b) Argon
(c) Chromium (d) Aluminium

18. An element of atomic number 29 belongs to **[CPMT 1991; Kurukshetra CEE 1991; MP PET 2001]**
(a) s-block (b) p-block
(c) d-block (d) f-block

19. The element californium belongs to the family **[MNR 1987]**
(a) Actinide series (b) Alkali metal family
(c) Alkaline earth family (d) Lanthanide series

20. On moving from left to right across a period in the table the metallic character **[CPMT 1986]**
(a) Increases
(b) Decreases
(c) Remains constant
(d) First increases and then decreases

21. An element with atomic number 20 will be placed in which period of the periodic table **[MNR 1986; UPSEAT 19**
(a) 4 (b) 3
(c) 2 (d) 1

22. The electronic structure $(n-1)d^{1-10} ns^{0-2}$ is characteristic of **[CET Pune 1998]**
(a) Transition elements (b) Lanthanides
(c) Actinides (d) Rare gases

23. The elements with atomic number 10, 18, 36, 54 and 86 are all **[CPMT 1976]**
(a) Light metals (b) Inert gases
(c) Halogens (d) Rare–earths

24. Elements of atomic number 6 is placed in **[CPMT 1978]**

(a) IV group
(b) IV period
(c) VI group
(d) III group

25. Which of the following elements is a lanthanide (Rare–earth element) **[Manipal MEE 1995]**

(a) Cadmium
(b) Californium
(c) Cerium
(d) Cesium

26. Mendeleef's periodic law is based on

(a) Atomic weight
(b) Atomic number
(c) Number of neutrons
(d) None of the above

27. The heaviest atom amongst the following is **[CPMT 1976; NCERT 1976]**

(a) U
(b) Ra
(c) Pb
(d) Hg

28. Which of the following pairs has both members from the same group of the periodic table **[CPMT 1985; MP PET/PMT 1998]**

(a) $Mg - Ba$
(b) $Mg - Na$
(c) $Mg - Cu$
(d) $Mg - K$

29. Which of the following pairs has both members from the same period of the periodic table **[CPMT 1985; UPSEAT 2001; BHU 2003]**

(a) $Na - Ca$
(b) $Na - Cl$
(c) $Ca - Cl$
(d) $Cl - Br$

30. Diagonal relationship is shown by **[DPMT 1984]**

(a) Elements of first period
(b) Elements of second period
(c) Elements of third period
(d) (b) and (c) both

31. The elements having the electronic configuration, $[Kr]\ 4d^{10} f^{14}, 5s^2 p^6 d^2, 6s^2$ belongs to **[CPMT 1982]**

(a) s-block
(b) p-block
(c) d-block
(d) f-block

32. Chemical property of Li and Mg similar because **[RPMT 2002]**

(a) These belong to same group
(b) Both ionisation potential is same
(c) Shows diagonal relationship
(d) Both electron affinity is same

33. According to the periodic law of elements, the variation in properties of elements is related to their **[AIEEE 2003]**

(a) Atomic masses
(b) Nuclear masses
(c) Atomic numbers
(d) Nuclear neutron-proton number

34. The element with atomic number 36 belongs to ...... block in the periodic table **[KCET 2003]**

(a) p
(b) s
(c) f
(d) d

35. Which group of the periodic table contains only metals **[UPSEAT 2003]**

(a) IIA
(b) IB
(c) IA
(d) None of these

36. The elements in which s and p-orbitals are present

(a) Common elements
(b) Inert gases
(c) Halogens
(d) Transitional elements

37. Aluminium is diagonally related to (in periodic table) **[MP PET 1993]**

(a) Li
(b) C
(c) B
(d) Be

38. An element has the electronic configuration $1s^2, 2s^2 2p^6,\ 3s^2 3p^6 3d^5, 4s^1$. It is a

(a) s-block element
(b) p-block element
(c) d-block element
(d) Inert gas

39. Which of the following show diagonal relationship **[KCET 2003; MP PMT 2003]**

(a) B and Si
(b) B and Al
(c) B and Ga
(d) B and C

40. Which of the following dinegative anion is quite common **[CPMT 2000]**

(a) S^{2-}
(b) Se^{2-}
(c) Te^{2-}
(d) O^{2-}

41. An element has electronic configuration $1s^2 2s^2 2p^6 3s^2 3p^4$. Predict their period, group and block **[CPMT 2000]**

(a) Period = 3rd, block = p, group = 16
(b) Period = 5th, block = s, group = 1
(c) Period = 3rd, block = p, group = 10
(d) Period = 4th, block = d, group = 12

42. If the atomic number of an element is 33, it will be placed in the periodic table in the **[RPET 1999; UPSEAT 2**

(a) First gp
(b) Third gp
(c) Fifth gp
(d) Seventh gp

43. Which of the following is the atomic number of a metal **[AIIMS 2000]**

(a) 32
(b) 34
(c) 36
(d) 38

44. Which of the following statement is not correct regarding hydrogen atom **[AIIMS 2000]**

(a) It resembles halogens in some properties
(b) It resembles alkali metals in some properties
(c) It can be placed in 7th group of periodic table
(d) It can not be placed in first group of periodic table

45. Lithium shows similarities to magnesium in its chemical behaviour because **[AFMC 2000]**

(a) Similar size, same electronegativity and lower polarizing power
(b) Similar size, greater electronegativity and similar polarizing power
(c) Similar size, same electronegativity and similar high polarizing power
(d) None of these

46. On going left to right in a period, in transition metals, their atomic volumes **[MP PMT 2003]**

(a) Decrease
(b) Increase
(c) Remain same
(d) None of these of correct

47. Electronic configuration of chalcons in their outermost orbit is

(a) $s^2 p^3$
(b) $s^2 p^4$

(c) s^2p^5 (d) s^2p^6

48. Which configuration represents a noble gas **[DPMT 2000]**

(a) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$

(b) $1s^2 2s^2 2p^6 3s^2 3p^6$

(c) $1s^2 2s^2 2p^6 3p^6$

(d) $1s^2 2s^2 2p^6 3s^2$

49. Which of the following pair has elements containing same number of electrons in the outermost orbit

[Kurukshetra CEE 1998; AFMC 2000]

(a) N, O (b) Na, Ca

(c) As, Bi (d) Pb, Sb

50. Dobereiner traids is **[RPMT 1997]**

(a) Na, K, Rb (b) Mg, S, As

(c) Cl, Br, I (d) P, S, As

51. As per the modern periodic law, the physical and chemical properties of elements are periodic functions of their

[RPMT 1997; EAMCET 1998]

(a) Atomic volume

(b) Electronic configuration

(c) Atomic weight

(d) Atomic size

52. Elements after atomic number 103 have been discovered till now. If an element with atomic number 106 were ever discovered which of the following electronic configuration will it possess **[AIIMS 1982]**

(a) $[Rn]5f^{14}6d^4 7s^2$ (b) $[Rn]5f^{14}6d^5 7s^1$

(c) $[Rn]5f^{14}6d^6 7s^0$ (d) $[Rn]5f^{14}6d^1 7s^2 7p^3$

53. The element *X, Y, Z* and *T* have the indicated electronic configurations. Starting with the innermost shell, which is the most metallic element **[CPMT 1979, 93]**

(a) $X = 2,8,4$ (b) $Y = 2,8,8$

(c) $Z = 2,8,8,1$ (d) $T = 2,8,8,7$

54. Which pair of atomic numbers represents s-block elements

[EAMCET 1990; RPMT 1997; MP PET 2003]

(a) 7, 15 (b) 6, 12

(c) 9, 17 (d) 3, 12

55. Which pair of elements has same chemical properties

[EAMCET 1987]

(a) 13, 22 (b) 3, 11

(c) 4, 24 (d) 2, 4

56. Mosley's name is most closely associated with the discovery of

(a) Positron (b) Deutrons

(c) Atomic number (d) Atomic weight

57. In the periodic table going down in fluorine group

[CPMT 1981]

(a) Reactivity will increase

(b) Electronegativity will increase

(c) Ionic radius will increase

(d) Ionization potential will increase

58. Beryllium resembles much with **[CPMT 1988]**

(a) Zn (b) Al

(c) Li (d) Ra

59. The last member in each period of the periodic table is

[DPMT 2001]

(a) An inert gas element (b) A transition element

(c) A halogen (d) An alkali metal

60. Which one of the following combination represents a metallic element **[EAMCET 1979]**

(a) 2, 8, 7 (b) 2, 8, 8

(c) 2, 8, 4 (d) 2, 8, 2

61. The electronic configuration of an atom A is $1s^2$, $2s^2p^6$, $3s^2p^6d^{10}, 4s^2p^3$. The chemistry of A is therefore likely to be similar to that of **[MP PMT 1995]**

(a) Chlorine (b) Nitrogen

(c) Oxygen (d) Boron

62. The element having the electronic configuration $1s^2$, $2s^2\,2p^6, 3s^2\,3p^1$ is

(a) A transition element

(b) A representative element

(c) An inert gas

(d) An inner–transition element

63. The element with configuration $1s^2, 2s^2p^6, 3s^2$ would be

[CPMT 1986; MP PMT 1993]

(a) A metal (b) A non-metal

(c) An inert gas (d) A metalloid

64. The long form of periodic table is based on **[CPMT 1997]**

(a) Shape of the atom

(b) Mass of the atom

(c) Atomic number of the atom

(d) Electronegativity

65. Chloride of an element A gives neutral solution in water. In the periodic table, the element A belongs to

[AIIMS 1992; UPSEAT 2001]

(a) First group (b) Third group

(c) Fifth group (d) First transition series

66. The fundamental basis of the present-day Periodic Table is that elements are **[JIPMER 1999]**

(a) Arranged in the order of increasing atomic weights

(b) Grouped according to chemical properties

(c) Arranged in the order of increasing number of neutrons in the atomic nucleus

(d) Arranged in the order of increasing number of protons in the nucleus

67. All the elements in a group in the periodic table have the same **[NCERT 1974; MP PET 1996; MP PMT 1996]**

(a) Atomic number

(b) Electronic configuration

(c) Atomic weight

(d) Number of electrons in the outermost shell or number of electrons for bonding

68. The most predominantly ionic compounds will be obtained from the combination of elements belonging to

(a) 1 and 7 groups (b) 2 and 6 groups

(c) 3 and 5 groups (d) 0 and 7 groups

69. An atom with atomic number 21 belongs to the category of

[Kurukshetra CEE 1991]

(a) s-block elements (b) p-block elements
(c) d-block elements (d) f-block elements

70. Which metal has 2 electrons in the outermost orbit
(a) Na (b) Cu
(c) Au (d) Be

71. In the modern periodic table, elements are arranged in **[MP PMT 1990; MP PET 1995; CPMT 1971, 73, 78, 80]**
(a) Increasing mass
(b) Increasing volume
(c) Increasing atomic number
(d) Alphabetically

72. Alkali metals in each period have **[MP PMT 1995]**
(a) Smallest size
(b) Lowest ionization potential
(c) Highest ionization potential
(d) Highest electronegativity

73. The elements on the right side of the periodic table are **[CPMT 1976]**
(a) Metals (b) Metalloids
(c) Non-metals (d) Transition elements

74. The screening effect of d-electons is **[RPMT 2000]**
(a) Equal to that of p-electrons
(b) More than that of p-electrons
(c) Same as f-electrons
(d) Less than p-electrons

75. Chemical behaviour of an atom is determined by
(a) Atomic number (b) Mass number
(c) Binding energy (d) Number of isotopes

76. Which of the following is a inert element
(a) Na (b) Fe
(c) Li (d) He

77. The lightest metal is **[CPMT 1976; NCERT 1976; AFMC 1988]**
(a) Li (b) Mg
(c) Ca (d) Na

78. Choose the typical element
(a) K (b) Na
(c) Sc (d) He

79. Of the following pairs, the one containing example of metalloid elements in the periodic table is
(a) Sodium and potassium
(b) Fluorine and chlorine
(c) Calcium and magnesium
(d) Boron and silicon

80. The number of elements in each of the long periods in the periodic table is
(a) 2 (b) 8
(c) 18 (d) 32

81. In the long form of the periodic table, all the non-metals are placed under **[EAMCET 1988]**
(a) s-block (b) p-block
(c) d-block (d) f-block

82. Elements with outer electronic configuration ns^2np^6 are **[MP PET/PMT 1998]**
(a) Alkaline earth metals (b) Transition elements
(c) Chalcogenes (d) Noble gases

83. Highest density is of .... **[RPET 2000]**
(a) Ir (b) Os
(c) Pb (d) Hg

84. Lithium shows diagonal relationship with **[MP PET 1995, 96; EAMCET 1990]**
(a) Al (b) Mg
(c) Be (d) B

85. $1s^2 2s^2 2p^6 3s^2$ is the electronic configuration of the metal **[RPET 2000]**
(a) Na (b) Mg
(c) Fe (d) Al

86. Element having atomic number 17 is placed in **[MP PET 1995]**
(a) I-group (b) V-group
(c) $VIII$-group (d) VII-group

87. The most importasnt active step in the development of periodic table was taken by **[CPMT 1976]**
(a) Mendeleef (b) Dalton
(c) Avogadro (d) Cavendish

88. Who is called the father of chemistry **[CPMT 1972]**
(a) Faraday (b) Priestley
(c) Rutherford (d) Lavosier

89. The total number of rare–earth elements is **[CPMT 1993]**
(a) 8 (b) 32
(c) 14 (d) 10

90. Which is metalloid **[Bihar MEE 1997]**
(a) Pb (b) Sb
(c) Bi (d) Zn
(e) Mg

91. The element or elements whose position is anomalous in the periodic table is
(a) Halogens (b) Fe, Co and Ni
(c) Inert gases (d) Hydrogen

92. An element M has an atomic mass 19 and atomic number 9. Its ion is represented by
(a) M^+ (b) M^-
(c) M^{2+} (d) M^{2-}

93. The number of elements in the 5th period of the periodic table are
(a) 8 (b) 10
(c) 18 (d) 32

94. The element with atomic number 55 belongs to **[MP PMT 1995]**
(a) s-block (b) p-block
(c) d-block (d) f-block

95. Coloured salts are formed by **[Bihar MEE 1996]**
(a) Alkali metals (b) Lanthanides
(c) Actinides (d) Transition metals
(e) None of these

96. Which one of the following is an s - block element **[MP PMT 1999]**
(a) Aluminium (b) Chromium
(c) Niobium (d) Potassium

97. In the modern periodic table, the place of the element with atomic number 31 is in **[MP PMT 1999]**

(a) *s* - block (b) *d*- block

(c) *p* -block (d) *f* – block

98. Last element of group-IV is found to be **[DPMT 1996]**

(a) Strong metallic

(b) Weak metallic

(c) Strong non-metallic

(d) Weak non-metallic

99. Elements of *d* group are called **[DPMT 1996]**

(a) Transition elements (b) Transuranic elements

(c) Metals (d) Metalloids

100. Which of the following is a normal element

(a) Ce (b) He

(c) Li (d) Ar

101. Which of the following is metalloid **[BHU 1996; AMU 2000]**

(a) Pb (b) Zn

(c) As (d) None of these

102. Under normal condition which of the following electronic configuration is able to form dipositive ion **[RPET 2000]**

(a) $[Ar]4s^1$ (b) $[Ne]2s^2 3p^6$

(c) $[Ne]3s^2$ (d) None of these

103. The statement that is true for the long form of the periodic table is **[IIT 1988]**

(a) It reflects the sequence of filling the electrons in the order of sub-energy levels *s, p, d* and *f*

(b) It helps to predict the stable valency 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 bonds between any two elements

104. To which block is related an element having electronic configuration $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$ in the periodic table **[MP PMT 1995]**

(a) *s* - block (b) *p* - block

(c) *d* - block (d) *f* – block

105. $Ce-58$ is a member of

(a) *s*-block elements (b) *p*-block elements

(c) *d*-block elements (d) *f*-block elements

106. Atomic number of elements represent

(a) Number of protons in the nucleus

(b) Number of neutrons in the nucleus

(c) Number of protons and neutrons in nucleus

(d) The valency of an element

107. As we go from left to right in period two of the periodic table, gram atomic volume of the elements

(a) Will change indefinitely

(b) Increases at a constant rate

(c) First increases then decrease

(d) Decreases

108. The electronic configuration of the element which is just above the element with atomic number 43 in the same periodic group is **[MNR 1992; UPSEAT 1999, 2000, 02]**

(a) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$

(b) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$

(c) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^1$

(d) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1 4p^6$

109. The elements indicating following atomic numbers belong to same group **[RPMT 1997]**

(a) 11 and 37 (b) 19 and 15

(c) 39 and 88 (d) None of these

110. Elements in which $4f$ orbitals are progressively filled are called as **[MP PET 1996]**

(a) Transition elements (b) Lanthanides

(c) Actinides (d) Inert gases

111. Hydrogen can be put in halogen group because **[RPMT 2000]**

(a) It has deuterium and tritium as isotopes

(b) It forms hydrides like chlorides

(c) It contains one electron only

(d) It is light

112. In the main group elements (i) as we proceed down the same group in the periodic table and (ii) as we proceed from left to right in the same period, the atomic radius

(a) (i) Increase continuously; (ii) Decreases continuously

(b) (i) Decreases continuously; (ii) Increases continuously

(c) (i) Increases continuously; (ii) Decreases upto the group IV and then increases upto the end of the period.

(d) (i) Decreases continuously; (ii) Decreases upto the group IV and then increases upto the end of the period.

113. Cause of diagonal relationship is

(a) Similar electronic configuration of the elements

(b) Similar e/r ratio of the elements

(c) Same number of valency electrons in the elements

(d) Same atomic weights of the elements

114. From which of the following the hydration energy of Mg^{2+} is larger **[MP PET 2000]**

(a) Na^+ (b) Al^{3+}

(c) Be^{2+} (d) Cr^{3+}

115. Group comprising of all metals is **[RPET 2000]**

(a) III*A* (b) IV*A*

(c) VII*A* (d) II*A*

116. Whose name is not associated with the development of Periodic Table

(a) Prout's (b) Newlands

(c) Rutherford (d) Loother Meyer

117. Element of atomic number 23 is placed in the periodic table in **[MP PMT 1996]**

(a) *s* - block (b) *p* – block

(c) *d* - block (d) *f* - block

118. In which of the following groups all the three members are of the alkaline earth metals family

(a) Al, Sr, Ti (b) Li, Na, K

(c) Mg, Ba, Ca (d) Rb, Cs, Fr

119. Astatine is a **[RPET 2000]**

(a) Halogen

(b) Rare earth element

(c) Alkaline earth metal

(d) None of these

120. The nitride ion in lithium nitride is composed of **[CBSE PMT 2001]**

(a) $7P+7e$ (b) $10P+7e$

(c) $7P+10e$ (d) $10P+10e$

121. Which set has the same number of unpaired electrons in their ground state **[JIPMER 2000]**

(a) Cl^-, Fe^{3+}, Cr^{3+} (b) $Na^+,\ Mg^{2+},\ Al$

(c) $Na, P,\ Cl$ (d) N, P, V

122. Which of the following doesn't decompose on heating **[AMU 2002]**

(a) $MgCO_3$ (b) Na_2CO_3

(c) Li_2CO_3 (d) $Ca(HCO_3)_2$

123. Which of the following has smallest bond angle **[AMU 2002]**

(a) H_2O (b) NH_3

(c) CH_4 (d) CO_2

124. The metal-having highest melting point is **[AMU 2002]**

(a) Chromium (b) Tungston

(c) Diamond (d) Silver

125. The elements with atomic numbers 9, 17, 35, 53, 85 are all **[KCET 2004]**

(a) Noble gases (b) Halogens

(c) Heavy metals (d) Light metals

126. The atomic number of an element is derived from **[Kerala PMT 2004]**

(a) Number of electrons

(b) Number of protons

(c) Number of neutrons

(d) Number of isotopes

(e) Number of nucleons

127. Beryllium shows diagonal relationship with **[Pb.CET 2003]**

(a) Mg (b) Na

(c) B (d) Al

128. Which of the properties remains unchanged on descending a group in the periodic table **[MP PMT 1997; RPMT 2002]**

(a) Atomic size (b) Density

(c) Valence electrons (d) Metallic character

129. Which of the following element does not occur in liquid form **[RPMT 2002]**

(a) Hg (b) Li

(c) Ga (d) Br

130. The cause of periodicity of properties is

(a) Increasing atomic radius

(b) Increasing atomic weights

(c) Number of electrons in the valency orbit

(d) The re-occurrence of similar outer electronic configuration

131. The chemistry of lithium is very similar to that of magnesium even though they are placed in different groups **[NCERT 1982]**

(a) Both are found together in nature

(b) Both have nearly the same size

(c) Both have similar electronic configuration

(d) The ratio of their charge to size is nearly the same

Atomic and Ionic radii

1. The ratio between radii of He^+ ion and H atom is **[MP PET 1996]**

(a) $\frac{1}{2}$ (b) 1

(c) $\frac{3}{2}$ (d) 2

2. The smallest among the following ions is **[JIPMER 1999]**

(a) Na^+ (b) Mg^{+2}

(c) Ba^{2+} (d) Al^{3+}

3. Which is smallest in size **[RPMT 1997]**

(a) O^{2-} (b) C^{4-}

(c) F^- (d) N^{3-}

4. Which of the following has largest size **[MP PMT 1995, 2003; JIPMER (Med.) 2002]**

(a) Al (b) Al^+

(c) Al^{+2} (d) Al^{+3}

5. Of the following, the one with largest size is **[EAMCET 1997; BHU 1999]**

(a) Cl^- (b) Ar

(c) K^+ (d) Ca^{2+}

6. Which cation has smallest radius **[RPET 2000]**

(a) K^+ (b) Na^+

(c) Li^+ (d) Be^{2+}

7. The radii of F, F^-, O and O^{-2} are in the order of **[AIIMS 1999; CPMT 1999]**

(a) $O^{2-} > F^- > O > F$ (b) $O^{2-} > F^- > F > O$

(c) $F^- > O^{2-} > F > O$ (d) $O^{2-} > O > F^- > F$

8. Which of the following has the smallest size **[CBSE PMT 1996]**

(a) Na^+ (b) Mg^{+2}

(c) Cl^- (d) F^-

9. Which of the following is largest **[CBSE PMT 1996]**

(a) Cl^- (b) S^{2-}

(c) Na^+ (d) F^-

10. Which of the following property displays progressive increase down a group in the Bohr's periodic table

(a) Electronegativity (b) Electron affinity
(c) Ionization potential (d) Size of the atom

11. Atomic radii of fluorine and neon in angstrom units are respectively given by **[IIT 1987]**
(a) 0.762, 1.60 (b) 1.60, 1.60
(c) 0.72, 0.72 (d) None of these values

12. Which ion has greatest radius in the following
[CPMT 1976; NCERT 1977]
(a) H^- (b) F^-
(c) Br^- (d) I^-

13. Which has the maximum atomic radius
[CPMT 1975; AIIMS 1982; DPMT 1982]
(a) Al (b) Si
(c) P (d) Mg

14. Which one of the following ions has the highest value of ionic radius **[AIEEE 2004]**
(a) O^{2-} (b) B^{3+}
(c) Li^+ (d) F^-

15. On going down a main sub-group in the periodic table (example Li to Cs in IA or Be to Ra in IIA), the expected trend of changes in atomic radius is a
[CPMT 1981; NCERT 1979]
(a) Continuous increase
(b) Continuous decrease
(c) Periodic one, an increase followed by a decrease
(d) A decrease followed by increase

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

17. Which one is the correct order of the size of the iodine species
[Pb. CET 1986; CBSE PMT 1997; Kurukshetra CEE 1998; RPMT 1999; DCE 1999; MP PET 2000; MP PMT 2001; BCECE 2005]
(a) $I > I^+ > I^-$ (b) $I > I^- > I^+$
(c) $I^+ > I^- > I$ (d) $I^- > I > I^+$

18. Which one has larger radius **[CPMT 1997; KCET 2005]**
(a) Na^+ (b) F
(c) F^- (d) Na

19. In third row of periodic table the atomic radii from Na to Cl
[MP PMT 1986]
(a) Continuosly decreases
(b) Continuosly increases
(c) Remains constant
(d) Increases but not continuously

20. The size of the following species increases in the order
[IIT-JEE 1990; AFMC 1995]
(a) $Mg^{2+} < Na^+ < F^- < Al$
(b) $F^- < Al < Na^+ > Mg^{2+}$
(c) $Al < Mg < F^- < Na^+$
(d) $Na^+ < Al < F^- < Mg^{2+}$

21. In K^+F^- ionic radius of F^- is more while atomic radius of K^+ is **[CPMT 1997]**
(a) Less than F^- (b) More than F^-
(c) Equal of F^- (d) None of these

22. Which one of the following species possesses maximum size
[EAMCET 1993; MP PET 2001]
(a) Na^+ (b) F^-
(c) Ne (d) O^{2-}

23. The ionic radii of N^{3-}, O^{2-}, F^- and Na^+ follow the order **[MP PET/PMT 1998; MP PMT 2000]**
(a) $N^{3-} > O^{2-} > F^- > Na^+$
(b) $N^{3-} > Na^+ > O^{2-} > F^-$
(c) $Na^+ > O^{2-} > N^{3-} > F^-$
(d) $O^{2-} > F^- > Na^+ > N^{3-}$

24. On moving down a group of regular elements, both atomic and ionic radii increases with increasing **[BMEE 1995]**
(a) Atomic number (b) Atomic weight
(c) Atomic mass (d) None of these

25. Which one of the following indicates the correct order of atomic size **[EAMCET 1989]**
(a) $Be > F > C > Ne$ (b) $Be < C < F < Ne$
(c) $Be > C > F > Ne$ (d) $F < Ne < Be < C$

26. Which has the smallest size **[MP PET 1999]**
(a) Na^+ (b) Mg^{2+}
(c) Al^{3+} (d) P^{5+}

27. A sodium cation has a different number of electrons from
(a) O^{2-} (b) F^-
(c) Li^- (d) Al^{3+}

28. Which of the following statement concerning lanthanides elements is false **[CBSE PMT 1994]**
(a) Lanthanides are separated from one another by ion exchange method
(b) Ionic radii of trivalent lanthanides steadily increases with increase in the atomic number
(c) All lanthanides are highly dense metals
(d) More characteristic oxidation state of lanthanide elements is +3

29. The lanthanide contraction is responsible for the fact that
[CBSE PMT 1997]
(a) Zr and Y have about the same radius
(b) Zr and Nb have similar oxidation state
(c) Zr and Hf have about the same radius
(d) Zr and Zn have the same oxidation state

30. Elements of which group form anions most readily
[CBSE PMT 1992]
(a) Oxygen family (b) Nitrogen group
(c) Halogens (d) Alkali metals

31. The unit representing atomic radii and ionic radii is
(a) nm (b) cm
(c) Å (d) m

32. The atomic radii in periodic table among elements from right to left **[MP PET 1995]**
(a) Decreases
(b) Increases
(c) Remain constant
(d) First decreases and then increases

33. Of the following the ion with the smallest ionic radius is **[MP PET 1996]**

(a) K^+ (b) Ca^{2+}

(c) Ti^{3+} (d) Ti^{4+}

34. Which of the following does not represent the correct order of the property indicated **[CBSE PMT 1997]**

(a) $Sc^{3+} > Cr^{3+} > Fe^{3+} > Mn^{3+}$ ionic radii

(b) $Sc < Ti < Cr < Mn$ Density

(c) $Mn^{2+} > Ni^{2+} < Co^{2+} < Fe^{2+}$ ionic radii

(d) $FeO < CaO > MnO > CuO$ Basic nature

35. The order of magnitude of ionic radii of ions Na^+, Mg^{2+}, Al^{3+} and Si^{4+} is **[MP PMT 1996]**

(a) $Na^+ < Mg^{2+} < Al^{3+} < Si^{4+}$

(b) $Mg^{2+} > Na^+ > Al^{3+} > Si^{4+}$

(c) $Al^{3+} > Na^+ > Si^{4+} > Mg^{2+}$

(d) $Na^+ > Mg^{2+} > Al^{3+} > Si^{4+}$

36. The order of the magnitude of ionic radii of ions N^{3-}, O^{2-} and F^- is **[MP PMT 1996]**

(a) $N^{3-} > O^{2-} > F^-$ (b) $N^{3-} < O^{2-} < F^-$

(c) $N^{3-} > O^{2-} > F^-$ (d) $N^{3-} < O^{2-} > F^-$

37. Which statement is correct

(a) For potassium, the atomic radius < ionic radius; but for bromine, the atomic radius > ionic radius

(b) For potassium and bromine both, the atomic radii > ionic radii

(c) For potassium and bromine both, the atomic radii < ionic radii

(d) For potassium, the atomic radius > ionic radius but for bromine, the atomic radius < ionic radius

38. Which of the following ion is the smallest ion **[AIIMS 2001]**

(a) O_2^+ (b) O_2^-

(c) O_2 (d) O_2^{-2}

39. The correct order of radii is **[IIT-JEE (Screening) 2000]**

(a) $N < Be < B$ (b) $F^- < O^{2-} < N^{3-}$

(c) $Na < Li < K$ (d) $Fe^{3+} < Fe^{2+} < Fe^{4+}$

40. Which one of the following should be most stable **[MP PET 2000]**

(a) H_2^+ (b) H^+

(c) H (d) H^-

41. Which of the following is the correct order of ionic radii **[BHU 2002]**

(a) $F > Li > Na > K$ (b) $F > K > Na > Li$

(c) $Na > K > F > Li$ (d) $Li > Na > K > F$

42. Smallest among these species is **[KCET 2002]**

(a) Lithium ion (b) Hydrogen

(c) Lithium (d) Helium

43. Which of the following ionic radius would be maximum **[MP PET 1997]**

(a) C^{4-} (b) N^{3-}

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

44. Which is helpful in the formation of ionic bond

(a) Only small cation

(b) Only small anion

(c) Small cation and small anion both

(d) Low positive charge, large cation and small anion

45. Which of the following has largest ionic radius **[AFMC 1999; BHU 2003]**

(a) Cs^+ (b) Li^+

(c) Na^+ (d) K^+

46. Point out the wrong statement :

On moving horizontally from left to right across a period in the periodic table

(a) Metallic character decreases

(b) Electronegativity increases

(c) Gram atomic volume first decreases and then increases

(d) Size of the atoms increases for normal elements

47. Which of the following statements is correct **[MP PET 1997]**

(a) X^- ion is larger in size than X atom

(b) X^+ ion is larger in size than X atom

(c) X^+ ion is larger in size than X^- ion

(d) X^+ and X^- ions are equal in size

48. The atomic radius of elements of which of the following series would be nearly the same **[MP PET 1997]**

(a) $Na\ K\ Rb\ Cs$ (b) $Li\ Be\ B\ C$

(c) $Fe\ Co\ Ni\ Cu$ (d) $F\ Cl\ Br\ I$

49. The decreasing order of size of isoelectronic series K^+, Ca^{2+}, Cl^- and S^{2-} is **[Roorkee 1995]**

(a) $K^+ > Ca^{2+} > S^{2-} > Cl^-$

(b) $K^+ > Ca^{2+} > Cl^- > S^{2-}$

(c) $Ca^{2+} > K^+ > Cl^- > S^{2-}$

(d) $S^{2-} > Cl^- > K^+ > Ca^{2+}$

50. Which of the following sets of elements have the strongest tendency to form anions **[MP PET 1993]**

(a) *N, O, F* (b) *P, S, Cl*

(c) *As, Se, Br* (d) *Sb, Te, I*

51. Radius of the isoelectronic species **[MP PET 1994]**

(a) Increases with the increase of nuclear charge

(b) Decreases with the increase of nuclear charge

(c) Is the same for all

(d) First increases and then decreases

52. In which of the following pairs the difference between the covalent radii of the two metals is maximum **[MP PET 1994]**

(a) *K, Ca* (b) *Mn , Fe*

(c) *Co , Ni* (d) *Cr , Mn*

53. An atom of an element has electronic configuration 2, 8, 1. Which of the following statement is correct **[MP PMT 1994]**

(a) The element's valency is 7

(b) The element exists as a diatomic molecule

(c) The element is of non-metallic nature

(d) The element forms a basic oxide

54. Which of the following ions has the smallest radius

[KCET 1992]

(a) Be^{2+} (b) Li^{+}

(c) O^{2-} (d) F^{-}

55. Point out the *wrong* statement :

In a given period of the periodic table the *s* - block element has, in general, a lower value of **[MP PMT 1997]**

(a) Ionisation energy (b) Electronegativity

(c) Atomic radius (d) Electron affinity

56. Arrange the following in increasing order of their atomic radius : Na, K, Mg, Rb **[AFMC 1995, 97; CPMT 1999]**

(a) $Mg < K < Na < Rb$ (b) $Mg < Na < K < Rb$

(c) $Mg < Na < Rb < K$ (d) $Na < K < Rb < Mg$

57. In the isoelectronic species the ionic radii (Å) of N^{3-}, O^{2-} and F^{-} are respectively given by **[Pb. CET 1989]**

(a) 1.36, 1.40, 1.71 (b) 1.36, 1.71, 1.40

(c) 1.71, 1.40, 1.36 (d) 1.71, 1.36, 1.40

58. Al^{3+} has a lower ionic radius than Mg^{2+} because

[EAMCET 1992]

(a) Mg atom has less number of neutrons than Al

(b) Al^{3+} has a higher nuclear charge than Mg^{2+}

(c) Their electronegativities are different

(d) Al has a lower ionisation potential than Mg atom

59. When a neutral atom is converted into cation, there is

[EAMCET 1986]

(a) Decrease in the atomic number

(b) An increase in the atomic number

(c) A decrease in size

(d) An increase in size

60. A trend common to both groups I and VII elements in the periodic table as atomic number increases is

[NCERT 1981; EAMCET 1980]

(a) Oxidising power increases

(b) Atomic radius increases

(c) Maximum valency increases

(d) Reactivity with water increases

61. Increasing order of atomic radii is **[RPET 2003]**

(a) $Mg^{2+} < Na^{+} < Ne < F^{-} < O^{2-}$

(b) $Na^{+} < Mg^{++} < Ne < F^{-} < O^{2-}$

(c) $O^{2-} < F^{-} < Ne < Na^{+} < Mg^{2+}$

(d) $Ne < O^{2-} < F^{-} < Na^{+} < Mg^{2+}$

62. Chloride ion and potassium ion are isoelectronic. Then

[KCET 2002]

(a) Potassium ion is relatively bigger

(b) Depends on the other cation and anion

(c) Their size are same

(d) Chloride ion is bigger than potassium ion

63. Which of the following has the largest ionic radius

[Pb. PMT 2002; BHU 2003]

(a) Na^{+} (b) Ni^{+}

(c) Cs^{+} (d) Mg^{+2}

64. The ionic radii of Li^{+}, Na^{+}, K^{+} are in which of the following order **[MP PMT 2002]**

(a) $K^{+} > Na^{+} > Li^{+}$ (b) $K^{+} > Na^{+} < Li^{+}$

(c) $K^{+} < Na^{+} < Li^{+}$ (d) $Li^{+} > Na^{+} < K^{+}$

65. Which of the following has smallest size

[JIPMER (Med.) 2002]

(a) Mg^{2+} (b) Na^{+}

(c) Al^{3+} (d) Si^{4+}

66. Which one of the following is expected to have largest size

[UPSEAT 2004]

(a) F^{-} (b) O^{-2}

(c) Al^{+3} (d) N^{-3}

67. The trivalent ion having largest size in lanthanide series is

[Pb.PMT 2004]

(a) Ti (b) Zr

(c) Hf (d) La

68. Which of the following alkali metal ions has lowest ionic mobility in aqueous solutions **[DPMT 2004]**

(a) Rb^{+} (b) Cs^{+}

(c) Li^{+} (d) Na^{+}

69. Ionic radii are **[CBSE PMT 2003, 04]**

(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.

70. The correct sequence of increasing covalent character is represented by **[CBSE PMT 2005]**

(a) $LiCl < NaCl < BeCl_2$ (b) $BeCl_2 < NaCl < LiCl$

(c) $NaCl < LiCl < BeCl$ (d) $BeCl_2 < LiCl < NaCl$

71. Correct energy value order is **[Orissa JEE 2004]**

(a) $ns\ np\ nd(n-1)f$ (b) $ns\ np(n-1)d\ (n-2)f$

(c) $ns\ np(n-1)d\ (n-1)f$ (d) $ns(n-1)d\ n(n-1)f$

72. The ionic conductance of following cation in a given concentration are in the order **[Orissa JEE 2004]**

(a) $Li^{+} < Na^{+} > K^{+} < Rb^{+}$

(b) $Li^{+} > Na^{+} > K^{+} > Rb^{+}$

(c) $Li^{+} < Na^{+} > K^{+} > Rb^{+}$

(d) $Li^{+} = Na^{+} < K^{+} < Rb^{+}$

Ionisation energy

1. The incorrect statement among the following is

[IIT-JEE 1997]

(a) The first ionisation potential of Al 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 the third ionisation potential of Al

2. The second ionisation potential of an element M is the energy required to **[JIPMER 1997]**

(a) Remove one mole of electron from one mole of gaseous anion

(b) Remove one mole of electron from one mole of gaseous cation of the element

(c) Remove one mole of electron from one mole of monovalent gaseous cation of the element

(d) Remove 2 moles of electrons from one mole of gaseous atoms

3. The ionization energy of an element is

(a) The same as the electron affinity of the element

(b) Equal in magnitude but of opposite sign to the electron affinity of the element

(c) The energy released when an electron is added to an atom of the element

(d) The energy required to remove the outermost electron of an atom of the element

4. The first ionisation energies of alkaline earth metals are higher than those of the alkali metals. This is because **[MP PET 1996]**

(a) There is increase in the nuclear charge of the alkaline earth metals

(b) There is a decrease in the nuclear charge of the alkaline earth metals

(c) There is no change in the nuclear charge

(d) None of the above

5. The statement that is not correct for the periodic classification of elements is **[IIT-JEE 1992]**

(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 along a period do not vary in a regular manner with increase in atomic number

(d) For transition elements the d-sub-shells are filled with electrons monotonically with increase in atomic number

6. Choose the correct statement

(a) Ionization energy and electron affinity increases across a period

(b) Ionization energy increases but electron affinity decreases along a period

(c) Ionization energy decreases but electron affinity increases

(d) Both decreases along a period

7. In halogens, with the increase of atomic number which habit is found

(a) Habit to loose electrons decreases

(b) Ionic radii decreases

(c) Ionization potential decreases

(d) In MX_2 ($M =$ metal and $X =$ halogen), covalent properties decreases

8. Ionization potential is lowest for **[CPMT 1989; MP PET 2001]**

(a) Halogens (b) Inert gases

(c) Alkaline earth metals (d) Alkali metals

9. Which of the following explanation is best for not placing hydrogen in either the group of alkali metals or halogens **[NCERT 1978]**

(a) The ionization energy of hydrogen is to high for group of alkali metals, but too low of halogen group

(b) Hydrogen can form compounds with all other elements

(c) Hydrogen is much lighter element than the alkali metals or the halogens

(d) None of the above

10. The ionization energy of nitrogen is more than that of oxygen because **[MP PET 1993]**

(a) Nitrogen has half filled p-orbitals

(b) Nitrogen is left to the oxygen in the same period of the periodic table

(c) Nitrogen contains less number of electrons

(d) Nitrogen is less electronegative

11. The energy required to remove an electron of a gaseous atom from its ground state is called **[CPMT 1989, 9**

(a) Potential energy (b) Ionization energy

(c) Electrode potential (d) Activation energy

12. The first ionization energy of boron is less than that of beryllium because

(a) Boron has higher nuclear charge

(b) Atomic size of boron is more than that of beryllium

(c) Boron has only one electron in p-sub-shell

(d) Atomic size of boron is less than that of beryllium

13. $A \to A^+ + e, E_1$ and $A^+ \to A^{2+} + e, E_2$. The energy required to pull out the two electrons are E_1 and E_2 respectively. The correct relationship between two energy would be

(a) $E_1 < E_2$ (b) $E_1 = E_2$

(c) $E_1 > E_2$ (d) $E_1 \neq E_2$

14. Which of the following element has maximum, first ionisation potential **[AIIMS 2001]**

(a) V (b) Ti

(c) Cr (d) Mn

15. Highest energy will be absorbed to eject out the electron in the configuration **[RPMT 2000]**

(a) $1s^2 2s^2 2p^1$ (b) $1s^2 2s^2 2p^3$

(c) $1s^2 2s^2 2p^2$ (d) $1s^2 2s^2 2p^4$

16. In which of the following process highest energy is absorbed **[RPET 2000]**

(a) $Cu \to Cu^+$ (b) $Br \to Br^-$

(c) $I \to I^-$ (d) $Li \to Li^+$

17. The first ionization potential of Na, Mg, Al and Si are in the order **[IIT 1988; MP PMT 2000]**

(a) $Na < Mg > Al < Si$ (b) $Na > Mg > Al > Si$

(c) $Na < Mg < Al > Si$ (d) $Na > Mg > Al < Si$

18. How many ionisation energies can carbon have

(a) 1 (b) 2

(c) 4 (d) 6

19. Which of the following gaseous atoms has highest value of IE **[JIPMER 1997; CPMT 1997; AIIMS 2000]**

(a) P (b) Si

(c) Mg (d) Al

20. Hydrogen has high ionization energy than alkali metals, due to its **[AIIMS 1999]**

(a) Large size
(b) Small size
(c) Ionic bond
(d) Covalent bond

21. The first ionization potentials (eV) of Be and B respectively are **[CBSE PMT 1998]**
(a) $8.29 eV$, $9.32 eV$
(b) $9.32 eV$, $9.32 eV$
(c) $8.29 eV$, $8.29 eV$
(d) $9.32 eV$, $8.29 eV$

22. Which ionisation potential (IP) in the following equations involves the greatest amount of energy **[Pune CET 1998]**
(a) $Na \rightarrow Na^{+} + e^{-}$
(b) $K^{+} \rightarrow K^{2+} + e^{-}$
(c) $C^{2+} \rightarrow C^{3+} + e^{-}$
(d) $Ca^{+} \rightarrow Ca^{2+} + e^{-}$

23. Which of the following has maximum ionization potential
[MH CET 1999]
(a) K
(b) Na
(c) Al
(d) Mg

24. The first four ionization energy values of an element are 191, 578, 872 and 5962 *kcal*. The number of valence electrons in the element is
(a) 1
(b) 2
(c) 3
(d) 4

25. Which of the following has least ionization potential
[CPMT 1982, 93]
(a) Li
(b) Cs
(c) Cl
(d) I

26. Which of the following element has the lowest ionization potential **[CPMT 1976; RPMT 2002]**
(a) Fe
(b) H
(c) Li
(d) He

27. As one moves along a given row in the periodic table, ionization energy
[CPMT 1976, 89; NCERT 1978; EAMCET 1985]
(a) Remains same
(b) Increases from left to right
(c) First increases, then decreases
(d) Decreases from left to right

28. Ionization energy is highest for **[AFMC 2001; BVP 2003]**
(a) Noble gases
(b) Platinum metals
(c) Transition elements
(d) Inner–transition elements

29. Which one of the following elements has the highest ionisation energy **[IIT-JEE 1990]**
(a) $[Ne]3s^2 3p^1$
(b) $[Ne]3s^2 3p^2$
(c) $[Ne]3s^2 3p^3$
(d) $[Ar]3d^{10} 4s^2 4p^2$

30. Which of the following elements has the lowest ionistion potential **[EAMCET 1993]**
(a) N
(b) O
(c) F
(d) Ne

31. Which of the following has lowest first ionisation potential
[CPMT 1993]
(a) B
(b) C
(c) N
(d) O

32. If first orbit energy of He^{+} is – 54.4 *eV*, then the second orbit energy will be **[Roorkee 1995]**
(a) – 54.4 *eV*
(b) – 13.6 *eV*
(c) – 27.2 *eV*
(d) + 27.2 *eV*

33. The screening effect of inner electrons of the nucleus causes
[MP PMT 1994]
(a) A decrease in the ionisation potential
(b) An increase in the ionisation potential
(c) No effect on the ionisation potential
(d) An increase in the attraction of the nucleus to the electrons

34. Which of the following has highest first ionization energy
[MP PET 1994]
(a) Sulphur
(b) Oxygen
(c) Nitrogen
(d) Phosphorus

35. The second ionization potential is
[Bihar CEE 1995; CET Pune 1998]
(a) Less than the first ionization potential
(b) Equal to the first ionization potential
(c) Greater than the first ionization potential
(d) None of these

36. When the first ionization energies are plotted against atomic number the peaks are occupied **[CET Pune 1**
(a) Alkali metals
(b) Halogens
(c) Rare gases
(d) Transition elements

37. Among the following which has the highest first ionization energy
(a) K
(b) Na
(c) B
(d) Kr

38. The first ionisation potential will be maximum for
[CPMT 2000]
(a) Lithium
(b) Hydrogen
(c) Uranium
(d) Iron

39. Arrange S, P, As in order of increasing ionisation energy
[JIPMER (Med.) 2002]
(a) $S < P < As$
(b) $P < S < As$
(c) $As < S < P$
(d) $As < P < S$

40. With reference to concept of ionisation potential, which one of the following sets are correct **[Kurukshetra C**
(a) $U > K > Cs$
(b) $B > U > K$
(c) $Cs > U > B$
(d) $Cs < U < K$

41. Which among the following species has the highest ionisation potential **[KCET 2001]**
(a) B
(b) Li
(c) Ne
(d) F

42. The set representing the correct order of first ionisation potential is **[IIT-JEE (Screening) 2001]**
(a) $K > Na > Li$
(b) $Be > Mg > Ca$
(c) $B > C > N$
(d) $Ge > Si > C$

43. Among the following options, the sequence of increasing first ionisation potential will be
[AIIMS 2000; MP PMT 2002]
(a) $B < C < N$
(b) $B > C > N$
(c) $C < B < N$
(d) $N > C > B$

44. The decreasing order of the ionisation potential in the following elements is **[MP PMT 2001]**
(a) $Ne > Cl > P > S > Al > Mg$
(b) $Ne > Cl > P > S > Mg > Al$
(c) $Ne > Cl > S > P > Mg > Al$
(d) $Ne > Cl > S > P > Al > Mg$

45. Which is the correct order of the first ionization potential of N, O and C **[AMU 2000]**
(a) $C > N > O$
(b) $C < N > O$
(c) $O > N > O$
(d) $C > N \sim O$

46. Which of the following order is wrong **[CBSE 2002]**

(a) $NH_3 < PH_3 < AsH_3$ -acidic nature

(b) $Li^+ < Na^+ < K^+ < Cs^+$ -ionic radius

(c) $Al_2O_3 < MgO < Na_2O < K_2O$ -basic

(d) $Li < Be < B < C$ -1st ionisation potential

47. Which of the following has the least ionization potential **[MP PET 2002]**

(a) Lithium (Li) (b) Helium (He)

(c) Nitrogen (N) (d) Zinc (Zn)

48. The first ionisation energy of lithium will be **[EAMCET 1990]**

(a) Greater than Be (b) Less than Be

(c) Equal to that of Na (d) Equal to that of F

49. Spectrum of Li^{2+} is similar to that of **[AIIMS 2002]**

(a) H (b) He

(c) Be (d) Ne

50. Highest ionisation energy stands for **[DPMT 2000]**

(a) He (b) C

(c) N (d) H

51. Which of the following electrons should have the highest value of ionisation energy (for the same value of the principal quantum number)

(a) s (b) p

(c) d (d) f

52. The correct sequence of elements in decreasing order of first ionisation energy is **[MP PET 1997]**

(a) $Na > Mg > Al$ (b) $Mg > Na > Al$

(c) $Al > Mg > Na$ (d) $Mg > Al > Na$

53. Correct order of polarising power is **[MP PMT 2003; BHU 2003]**

(a) $Cs^+ < K^+ < Mg^{2+} < Al^{3+}$

(b) $K^+ < Cs^+ < Mg^{2+} < Al^{3+}$

(c) $Cs^+ < K^+ < Al^{3+} < Mg^{2+}$

(d) $K^+ < Cs^+ < Al^{3+} < Mg^{2+}$

54. Correct increasing order of first ionistion potential is **[UPSEAT 2003]**

(a) $Na < Mg > Al < Si$ (b) $Na < Mg < Al < Si$

(c) $Na > Mg > Al > Si$ (d) $Na < Mg < Al > Si$

55. The ionisation potential of hydrogen from ground state to the first excited state is **[DCE 2001]**

(a) $-13.6\ eV$ (b) $13.6\ eV$

(c) $-3.4\,eV$ (d) $3.4\ eV$

56. In view of their low ionisation energies the alkali metals are **[MP PMT 2002]**

(a) Weak oxidising agents

(b) Strong reducing agents

(c) Strong oxidising agents

(d) Weak reducing agents

57. Of the following iso-electronic ions, the one which has the lowest ionisation potential is **[AMU 1999]**

(a) Na^+ (b) Mg^{++}

(c) F^- (d) O^{--}

58. Ionisation energy in group I-A varies in the decreasing order as **[Orissa JEE 2005]**

(a) $Li > Na > K > Cs$ (b) $Na > Li > K > Cs$

(c) $Li > Cs > K > Na$ (d) $K > Cs > Na > Li$

59. Which of the following relation is correct with respect to first (I) and second (II) ionization potentials of sodium and magnesium **[CPMT 1999]**

(a) $I_{Mg} = II_{Na}$ (b) $I_{Na} > I_{Mg}$

(c) $II_{Mg} > II_{Na}$ (d) $II_{Na} > II_{Mg}$

60. The order of the magnitude of first ionisation potentials of Be, B, N and O is **[MP PMT 1996]**

(a) $N > O > Be > B$ (b) $N > Be > O > B$

(c) $Be > B > N > O$ (d) $B > Be > O > N$

61. Which of the following transitions involves maximum amount of energy **[AIIMS 1992]**

(a) $M^-(g) \to M(g)$ (b) $M(g) \to M^+(g)$

(c) $M^+(g) \to M^{2+}(g)$ (d) $M^{2+}(g) \to M^{3+}(g)$

62. Which of the following species has lowest ionization potential **[KCET 1996]**

(a) O (b) O_2

(c) O_2^+ (d) O_2^-

63. Which of the following has minimum ionization energy **[JIPMER 1999]**

(a) Ge (b) Se

(c) As (d) Br

64. First I.P. of Mg is ...... than Al **[CPMT 1997]**

(a) Less (b) More

(c) Equal (d) None of these

65. The element with highest value of ionization potential is

(a) Potassium (b) Helium

(c) Hydrogen (d) Xenon

66. Which has the highest second ionisation potential **[AIIMS 1991]**

(a) Nitrogen (b) Carbon

(c) Oxygen (d) Fluorine

67. In ionisation of hydrogen, the energy required is **[CPMT 1996]**

(a) $13.6eV$ (b) $> 13.6eV$

(c) $< 13.6eV$ (d) $1.5eV$

68. Which of the following elements will have the lowest first ionisation energy **[KCET 1992]**

(a) Mg (b) Rb

(c) Li (d) Ca

69. In the long form of periodic table, the element having lowest ionisation potentials are present in**[EAMCET**

(a) I group (b) IV group

(c) VII group (d) Zero group

70. The process requiring the absorption of energy is **[Roorkee 1990]**

(a) $F \to F^-$ (b) $Cl \to Cl^-$

(c) $O \to O^{2-}$ (d) $H \to H^-$

71. In a period from Li to F, ionization potential **[CPMT 1982]**

(a) Increases (b) Decreases

(c) Remains same (d) None of the above

72. Ionization energy increases in the order

(a) Be, B, C, N (b) B, Be, C, N

(c) C, N, Be, B (d) N, C, Be, B

73. A neutral atom will have the lowest ionization potential when its electronic configuration is **[NCERT 1978; CBSE PMT 1991]**

(a) $1s^1$ (b) $1s^2, 2s^2p^6$

(c) $1s^2, 2s^2p^2$ (d) $1s^2, 2s^2p^6, 3s^1$

74. Which has maximum first ionization potential **[IIT 1982; EAMCET 1997; KCET (Med.) 1999; KCET 2000]**

(a) C (b) N

(c) B (d) O

75. Which one of the following elements has the highest ionisation energy

(a) Na (b) Mg

(c) C (d) F

76. Order of first ionization potentials of elements *Li, Be, B, Na* is **[Kerala CET 2005]**

(a) $Li > Be > B > Na$ (b) $Be > B > Li > Na$

(c) $Na > Li > B > Be$ (d) $Be > Li > B > Na$

(e) $B > Be > Li > Na$

77. The ionization energy of nitrogen is larger than that of oxygen because of **[RPMT 1997; DCE 1999]**

(a) Greater attraction of electrons by the nucleus

(b) The size of nitrogen atom being smaller

(c) The half-filled p-orbitals possess extra stability

(d) Greater penetration effect

78. If the IP of Na is 5.48 *eV*, the ionisation potential of K will be **[EAMCET 1988]**

(a) Same as that of Na (b) 5.68 *eV*

(c) 4.34 *eV* (d) 10.88 *eV*

79. *Mg* and *Li* are similar in their properties due to **[AFMC 2004]**

(a) Same *e/m* ratio (b) Same electron affinity

(c) Same group (d) Same ionic potential

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

$O_{(g)} + e^- = O^-_{(g)} \Delta H^0 = -142\ kJmol^{-1}$

$O^-_{(g)} + e^- = O^{2-}_{(g)} \Delta H^0 = 844\ kJmol^{-1}$

This is because **[AIEEE 2004]**

(a) O^- ion will tend to resist the addition of another electron

(b) Oxygen has high electron affinity

(c) Oxygen is more electronegative

(d) O^- ion has comparatively larger size than oxygen atom

81. Which is correct about ionisation potential **[MHCET 2003]**

(a) It is independent of atomic radii

(b) It increases with increase in atomic radii

(c) It remains constant with increase in atomic radii

(d) It decreases with increase in atomic radii

82. Flourine is the best oxidising agent because it has **[CPMT 2004]**

(a) Highest electron affinity

(b) Highest E^0_{red}

(c) Highest E^0_{oxid}

(d) Lowest electron affinity

83. Which among the following elements have lowest value of IE_1 **[CPMT 2004]**

(a) Pb (b) Sn

(c) Si (d) C

84. In a given shell, the order of screening effect is **[Kerala PMT 2004]**

(a) $s > p > d > f$ (b) $f > d > p > s$

(c) $p < d < s < f$ (d) $d > f < s > p$

(e) $f > p > s > d$

85. Which of the following has the highest first ionisation energy **[BHU 2004]**

(a) Li (b) Be

(c) B (d) C

86. Which one of the following sets of ions represents the collection of isoelectronic species **[AIEEE 2004]**

(a) $K^+, Cl^-, Mg^{2+}, Sc^{3+}$ (b) $Na^+, Ca^{2+}, Sc^{3+}, F^-$

(c) $K^+, Ca^{2+}, Sc^{3+}, Cl^-$ (d) $Na^+, Mg^{2+}, Al^{3+}, Cl$

87. The correct order of reactivity of halogens is **[MHCET 2003]**

(a) $F > Cl > Br > I$ (b) $F < Cl > Br < I$

(c) $F < Cl < Br < I$ (d) $F < Cl < Br > I$

88. The first ionisation potential is maximum for **[CPMT 2004**

(a) B (b) N

(c) O (d) Be

89. The correct order of ionisation energy for comparing carbon, nitrogen and oxygen atoms is **[UPSEAT**

(a) $C > N > O$ (b) $C > N < O$

(c) $C < N > O$ (d) $C < N < O$

Electron affinity

1. Electron affinity depends on **[MP PMT 2002]**

(a) Atomic size

(b) Nuclear charge

(c) Atomic number

(d) Atomic size and nuclear charge both

2. Increasing order of electron affinity is **[RPET 2003]**

(a) $N < O < Cl < Al$ (b) $O < N < Al < Cl$

(c) $Al < N < O < Cl$ (d) $Cl < N < O < Al$

3. The correct order of electron affinity of *B, C, N, O* is **[MP PET 1997; J & K 2005]**

(a) $O > C > N > B$ (b) $B > N > C > O$

(c) $O > C > B > N$ (d) $O > B > C > N$

4. Which one has maximum electron affinity **[Roorkee 1995]**

(a) N (b) Be

(c) B (d) Cl

5. The electron affinity for the inert gases is **[Kurukshetra CEE 1998; MP PMT 2002]**

(a) Zero (b) High

(c) Negative (d) Positive

6. The electron affinities of halogens are $F = 322$, $Cl = 349$, $Br = 324$, $I = 295\,kJ\,mol^{-1}$. The higher value for Cl as compared to that of F is due to **[MP PMT 1**

(a) Weaker electron-electron repulsion in Cl
(b) Higher atomic radius of F
(c) Smaller electronegativity of F
(d) More vacant P - subshell in Cl

7. Which one of the following is an incorrect statement **[MP PMT 2001]**

(a) The ionisation potential of nitrogen is greater than that of oxygen
(b) The electron affinity of fluorine is greater than that of chlorine
(c) The ionisation potential of beryllium is greater than that of boron
(d) The electronegativity of fluorine is greater than that of chlorine

8. Electron affinity is the **[MP PMT 1993]**

(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

9. The electron affinity values for the halogens show the following trend **[Kerala PET 2002]**

(a) $F < Cl > Br > I$ (b) $F < Cl < Br < I$
(c) $F > Cl > Br > I$ (d) $F < Cl > Br < I$

10. Which element has maximum electron affinity

(a) Na (b) S
(c) Mg (d) Al

11. Which of the following has the least electron affinity in $kJmol^{-1}$ **[AFMC 2000]**

(a) Oxygen (b) Carbon
(c) Nitrogen (d) Boron

12. Fluorine has low electron affinity than chlorine because of **[CPMT 1997]**

(a) Smaller radius of fluorine, high density
(b) Smaller radius of chlorine, high density
(c) Bigger radius of fluorine, less density
(d) Smaller radius of chlorine, less density

13. For electron affinity of halogens which of the following is correct **[AIIMS 2004]**

(a) $Br > F$ (b) $F > Cl$
(c) $Br < Cl$ (d) $F > I$

14. Ionic compounds are formed most easily with **[DPMT 2005]**

(a) Low electron affinity, high ionisation energy
(b) High electron affinity, low ionisation energy
(c) Low electron affinity, low ionisation energy
(d) High electron affinity, high ionisation energy

15. In comparison with alkali metals, the electron affinity of halogens is

(a) Very high (b) Very low
(c) Nearly same (d) Exactly same

16. The electron affinity of **[CPMT 1994]**

(a) Carbon is greater than oxygen
(b) Sulphur is less than oxygen
(c) Iodine is greater than bromine
(d) Bromine is less than chlorine

17. The amount of energy which is released due to addition of extra electron to the outermost orbit of gaseous atom is called **[BHU 1996]**

(a) Electron capacity (b) Electron affinity
(c) Ionisation potential (d) Electronegativity

18. Which of the following species has the highest electron affinity **[KCET 1996]**

(a) F (b) O
(c) O^- (d) Na^+

19. The electron affinity values (in $kJ\,mol^{-1}$) of three halogens X, Y and Z are respectively – 349, – 333 and – 325. Then X, Y and Z are respectively**[EAMCET**

(a) F_2, Cl_2 and Br_2 (b) Cl_2, F_2 and Br_2
(c) Cl_2, Br_2 and F_2 (d) Br_2, Cl_2 and F_2

20. Nitrogen has lower electron affinity than its preceeding element carbon because

(a) Electron affinity decreases along a period
(b) Electron affinity generally increases along a period
(c) Nitrogen atom has half filled p-orbital
(d) Nitrogen is a p-block element

21. Electron affinity is the lowest for

(a) Nitrogen (b) Carbon
(c) Oxygen (d) Sulphur

22. Which one of the elements has the maximum electron affinity

[CPMT 1986; AFMC 1992, 95; Bihar MEE 1996; BHU 1997; CBSE PMT 1996, 99; MP PET 1995, 2001; AMU 2000]

(a) F (b) Cl
(c) Br (d) I

23. Which among the following factors is the most important in making fluorine the strongest oxidizing halogen **[AIEEE 2004]**

(a) Hydration enthalpy
(b) Ionization enthalpy
(c) Electron affinity
(d) Bond dissociation energy

24. Which of the following pairs show reverse properties on moving along a period from left to right and from top to down in a group **[DCE 2003]**

(a) Nuclear charge and electron affinity
(b) Ionisation energy and electron affinity
(c) Atomic radius and electron affinity
(d) None of these

25. Which of the following properties show gradual decrease with increase in atomic number across a period in the periodic table **[Pb. CET 2003]**

(a) Electron affinity (b) Ionization potential
(c) Electronegativity (d) Size of atom

26. Order of electron affinity of F, Cl, Br and I is .....

[AFMC 1999; Orissa JEE 2004,05]

(a) $F < Cl > Br > I$ (b) $F > Cl > Br > I$
(c) $F < Cl < Br < I$ (d) $F > Cl < Br > I$

27. Which one of the following arrangements represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species. **[CBSE PMT 2005]**

(a) $Cl < F < S < O$ (b) $O < S < F < Cl$

(c) $S < O < Cl < F$ (d) $F < Cl < O < S$

Electronegativity

1. Between HF, HCl, HBr and HI, HF has the highest ionic character because

(a) F has the highest electron affinity

(b) In HF, electronegativity difference is highest

(c) F^- ion has the highest value of ionic radius

(d) Atomic orbitals of H and F have almost similar energy

2. On going from right to left in a period in the periodic table the electronegativity of the elements

[MP PET/PMT 1998; MP PMT 2002]

(a) Increases

(b) Decreases

(c) Remain unchanged

(d) Decreases first then increases

3. On Pauling scale which of the following does not have electronegativity $\geq$ 3.0 **[MP PET 1994]**

(a) Oxygen (b) Nitrogen

(c) Chlorine (d) Bromine

4. Which one of the following represents the electronic configuration of the most electropositive element

[AIIMS 1982; CPMT 1994; MP PMT 2000]

(a) $[He]2s^1$ (b) $[Xe]6s^1$

(c) $[He]2s^2$ (d) $[Xe]6s^2$

5. An atom with high electronegativity has

[Kerala (Med.) 2003]

(a) Large size

(b) High ionisation potential

(c) Low electron affinity

(d) Low ionisation potential

6. Two elements whose electronegativities are 1.2 and 3.0 the bond formed between them would be**[MP PET 2002]**

(a) Ionic (b) Covalent

(c) Coordinate (d) Metallic

7. The solubilities of carbonates decreases down the magnesium group due to a decrease in**[AIEEE 2003]**

(a) Lattice energies of solids

(b) Hydration energies of cations

(c) Inter-ionic attraction

(d) Entropy of solution formation

8. Which element has the highest electronegativity

or

Which of the following is the most electronegative

[CPMT 1981; Roorkee 1995; MP PMT 2003; EAMCET 1980; CPMT 1989; MNR 1994; MP PMT 1999]

(a) F (b) He

(c) Ne (d) Na

9. Which element has the highest electronegativity

[MP PET/PMT 1998]

(a) C (b) Mg

(c) O (d) S

10. Keeping in view the periodic law and the periodic table suggest which of the following elements should have the maximum electronegative character **[MNR 1985]**

(a) P (b) As

(c) Bi (d) Sb

11. The outermost electronic configuration of the most electronegative element is

[MP PET 1996; RPMT 1997; MP PET 2004]

(a) ns^2np^3 (b) ns^2np^4

(c) ns^2np^5 (d) ns^2np^6

12. Going from fluorine to chlorine, bromine and iodine, the electronegativity **[MP PMT 2000]**

(a) Increases

(b) Decreases

(c) First decreases then increases

(d) Changes randomly

13. Of the following elements, which one has highest electro-negativity**[CPMT 1988; CBSE PMT 1991; BHU 1996; Kurukshetra CET 2002; Pb. PMT 2004]**

(a) I (b) Br

(c) Cl (d) F

14. Which of the following is most electronegative**[CPMT 1999**

(a) Carbon (b) Silicon

(c) Lead (d) Tin

15. The property of attracting electrons by the halogen atom in a molecule is called **[CPMT 1996]**

(a) Ionisation potential (b) Electron affinity

(c) Electronegativity (d) Electronic attraction

16. In third row of periodic table from Na to Cl

[MP PET 1986]

(a) Electronegativity increases

(b) Electronegativity decreases

(c) Ionization energy decreases

(d) Atomic volume increases

17. Which of the following is the most electropositive element

[AIIMS 1998]

(a) Aluminium (b) Magnesium

(c) Phosphorus (d) Sulphur

18. Which of the following sets of atoms is arranged in order of increasing electronegativity

(a) S, Si, P (b) S, P, Si

(c) Si, P, S (d) Si, S, P

19. Which of the following property displays progressive increase with the rise in atomic number across a period in the periodic table

(a) Electronegativity (b) Electron affinity

(c) Ionization potential (d) Size of the atom

20. With respect to chlorine, hydrogen will be

[NCERT 1978; MP PMT 2003]

(a) Electropositive (b) Electronegative

(c) Neutral (d) None of the above

21. The correct order of electropositive nature of Li, Na and K is

(a) $Li > Na > K$ (b) $Li > K > Na$

(c) $Na > K > Li$ (d) $K > Na > Li$

22. Electronegativity is a measure of the capacity of an atom to

[CPMT 1989]

(a) Attract electrons (b) Attract protons

(c) Repel electrons (d) Repel protons

23. With increasing atomic number in a certain period **[MP PMT 1987]**

(a) The chemical reactivity decreases
(b) The chemical reactivity increases
(c) The electropositive character increases
(d) The electronegative character increases

24. Which of the following have maximum electronegativity **[CPMT 1982]**

(a) Al (b) S
(c) Si (d) P

25. Which element has the lowest electronegativity **[CPMT 1976]**

(a) Li (b) F
(c) Fe (d) Cl

26. The attraction that an atom exerts on a pair of electrons that are being shared between that atom and another atom to which it is bonded by a covalent bond is referred to as its **[Manipal MEE 1995]**

(a) Electron affinity (b) Electronegativity
(c) Ionisation energy (d) Valence

27. The electronegativity of the following elements increases in the order **[IIT 1987]**

(a) *C, N, Si, P* (b) *N, Si, C, P*
(c) *Si, P, C, N* (d) *P, Si, N, C*

28. Choose the correct statement

(a) Electronegativity increases down a group
(b) Electronegativity decreases down a group
(c) Electronegativity decreases from left to right along a period
(d) Electronegativity changes along a group but remains constant along a period

29. In C, N, O and F the electronegativity **[DPMT 2001]**

(a) Decreases from carbon to fluorine
(b) Increases from carbon to fluorine
(c) Increases from carbon to oxygen and then decreases
(d) Decreases from carbon to oxygen and then increases

30. Which is the correct order of electronegativities **[EAMCET 1990]**

(a) $F > N < O > C$ (b) $F > N > O > C$
(c) $F < N < O < C$ (d) $F > N > O < C$

31. In the following, the element with the highest electropositivity is **[MP PET/PMT 1998]**

(a) Copper (b) Caesium
(c) Barium (d) Chromium

32. Which one of the following has the highest electronegativity **[UPSEAT 2004]**

(a) Br (b) Cl
(c) P (d) Si

33. Which or these have no unit **[AFMC 2004]**

(a) Electronegativity (b) Electron affinity
(c) Ionisation energy (d) Excitation potential

34. The polarising ability of which one of the following is highest **[DCE 2003]**

(a) Small highly +ve ion (b) Large +ve ion
(c) Small highly –ve ion (d) Large –ve ion

35. Among Al_2O_3, SiO_2, P_2O_3 and SO_2 the correct order of acid strength is **[AIEEE 2004]**

(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) $Al_2O_3 < SiO_2 < P_2O_3 < SO_2$

Valency and oxidation state

1. Which one of the following oxides is neutral **[IIT-JEE 1996]**

(a) CO (b) SnO_2
(c) ZnO (d) SiO_2

2. All element in 3rd period have **[JIPMER 1997]**

(a) An atomic number 3
(b) 3 complete sub-shells
(c) Valence electrons shell
(d) 3 electrons less than the octet

3. Which shows variable valency **[RPMT 1997]**

(a) *s* - block elements (b) *p* - block elements
(c) *d* - block elements (d) Radioactive elements

4. Most reducing agent is **[UPSEAT 1999]**

(a) K (b) Mg
(c) Al (d) Ba

5. Acidity of pentoxides in VA group **[CPMT 1982]**

(a) Decreases (b) Increases
(c) Remains same (d) None

6. If the valency shell electronic structure for an element is ns^2np^5, this element will belong to the group of **[CBSE PMT 1992]**

(a) Alkali metals (b) Inert metals
(c) Noble gases (d) Halogens

7. The order in which the following oxides are arranged according to decreasing basic nature is **[CPMT 19**

(a) Na_2O, MgO, Al_2O_3, CuO
(b) MgO, Al_2O_3, CuO, Na_2O
(c) Al_2O_3, MgO, CuO, Na_2O
(d) CuO, Na_2O, MgO, Al_2O_3

8. Strongest reducing agent is **[RPMT 1997]**

(a) Cl_2 (b) Cl^-
(c) Br^- (d) I^-

9. Metallic nature and basic nature of the oxides .... as we move along a period

(a) Increases
(b) Decreases
(c) First increases then decreases
(d) Remains constant

10. The correct order of increasing order of oxidising power is **[DCE 2000]**

(a) $F_2 < Cl_2 < Br_2 < I_2$ (b) $F_2 < Br_2 < Cl_2 < I_2$
(c) $Cl_2 < Br_2 < F_2 < I_2$ (d) $I_2 < Br_2 < Cl_2 < F_2$

11. The most basic among these hydroxides, is**[MP PMT 2003]**

(a) $Be(OH)_2$ (b) $Mg(OH)_2$
(c) $Ca(OH)_2$ (d) $Ba(OH)_2$

12. In any period the valency of an element with respect to oxygen **[Kerala (Med.) 2003]**
(a) Increases one by one from I*A* to VII*A*
(b) Decreases one by one form I*A* to VII*A*
(c) Increases one by one from I*A* to IV*A* and then decreases from V*A* to VII*A* one by one
(d) Decreases one by one from I*A* to IV*A* and then increases from V*A* to VII*A* one by one

13. Which will show maximum non-metallic character **[UPSEAT 2003]**
(a) B (b) Be
(c) Mg (d) Al

14. Which of the following halogen acids is least acidic **[RPET 2003]**
(a) HI (b) HCl
(c) HF (d) HBr

15. Pentavalency in phosphorus is more stable when compared to that of nitrogen even through they belong to same group is due to **[KCET 2002]**
(a) Reactivity of phosphorus
(b) Inert nature of nitrogen
(c) Dissimilar electronic configuration
(d) Larger size of phosphorus atom

16. In the ground state of cobalt atom (Z = 27) there are ……. unpaired electrons and thus the atom is…….
(a) 2, diamagnetic (b) 2, paramagnetic
(c) 3, diamagnetic (d) 3, paramagnetic

17. Variable valency in general, is exhibited by **[CPMT 1974, 84, 89; DPMT 1981, 82; MP PET 2001]**
(a) Transition elements (b) Gaseous elements
(c) Non-metals (d) s-block elements

18. 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

19. Which of the following oxides is most basic **[MP PET 1994]**
(a) Na_2O (b) Al_2O_3
(c) SiO_2 (d) SO_2

20. In the periodic table, the metallic character of elements **[MP PET 1993]**
(a) Decreases from left to right across a period and on descending a group
(b) Decreases from left to right across a period and increases on descending a group
(c) Increases from left to right across a period and on descending a group
(d) Increases from left to right across a period and decreases on descending a group

21. The halogen that most easily reduced is **[MP PMT 2000]**
(a) F_2 (b) Cl_2
(c) Br_2 (d) I_2

22. Which of the following is the correct order of gradually decreasing basic nature of the oxides **[MP PMT 19**
(a) Al_2O_3, MgO, Cl_2O_7, SO_3
(b) MgO, Al_2O_3, SO_3, Cl_2O_7
(c) Cl_2O_7, SO_3, Al_2O_3, MgO
(d) SO_3, Cl_2O_7, MgO, Al_2O_3

23. The correct order of reactivity of halogen is **[BHU 2000]**
(a) Flourine > bromine > chlorine > iodine
(b) Flourine > chlorine > bromine > iodine
(c) Iodine > bromine > chlorine > flourine
(d) Bromine > chlorine > flourine > iodine

24. Elements A and B with their respective electronic configurations $3d^{10}4s^1$ and $4d^{10}5s^1$ in their outermost shell are
(a) Both non-metals
(b) Both coinage metals
(c) A is a non-metal and B is coinage metal
(d) A is a coinage metal and B is non-metal

25. Which is the best reducing agent **[MP PET 2000]**
(a) F^- (b) Cl^-
(c) Br^- (d) I^-

26. Which of the following group of elements eliminates electron easily
(a) *N, P, As* (b) *O, S, Se*
(c) *Li, Na, K* (d) *Cl, Ba, I*

27. The maximum valency of an element with atomic number 7 is **[AFMC 2002]**
(a) 2 (b) 5
(c) 4 (d) 3

28. Which of the following metals exhibits more than one oxidation state **[MP PET 1999]**
(a) Na (b) Mg
(c) Fe (d) Al

29. Out of the following elements which one do you expect to be most reactive chemically **[CPMT 1983]**
(a) Mg (b) Ca
(c) Sr (d) Ba

30. Thalium shows different oxidation states because **[AIIMS 1982]**
(a) It is a transition element
(b) Of inert pair effect
(c) Of its amphoteric character
(d) Of its higher reactivity

31. Oxidising action increases in halogen in the following order **[DPMT 1990]**
(a) $Cl < Br < I < F$ (b) $Cl < I < Br < F$
(c) $I < F < Cl < Br$ (d) $I < Br < Cl < F$

32. Fluorine, chlorine, bromine and iodine are placed in the same group (17) of the periodic table, because **[KCET (Med.) 1999]**
(a) They are non-metals
(b) They are electronegative
(c) Their atoms are generally univalent
(d) They have 7 electrons in the outermost shell of their atom

33. Which of the following sequence correctly represents the decreasing acid nature of oxides **[AMU 2000]**

(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) $B_2O_3 > CO_2 > N_2O_3 > Li_2O > BeO$

34. Which of the following aqueous acid is most acidic **[AMU 2000]**

(a) HCl (b) HF

(c) HI (d) HBr

35. The correct order of the increasing ionic character is **[MP PET 2000]**

(a) $BeCl_2 < MgCl_2 < CaCl_2 < BaCl_2$

(b) $BeCl_2 < MgCl_2 < BaCl_2 > CaCl_2$

(c) $BeCl_2 < BaCl_2 < MgCl_2 < CaCl_2$

(d) $BaCl_2 < CaCl_2 < MgCl_2 < BeCl_2$

36. Which of the following elements is found in native state **[RPET 1999]**

(a) Al (b) Au

(c) Cu (d) Na

37. The basis of keeping the elements in the group of a periodic table is **[RPET 1999]**

(a) Ionisation potential

(b) Electronegativity

(c) Electron affinity

(d) Number of electrons in the valence shell

38. Which of the following electronic configurations in the outermost shell is characteristic of alkali metals **[Bihar CEE 1992]**

(a) $(n-1)s^2p^6, ns^2p^1$ (b) $(n-1)s^2p^6d^{10}, ns^1$

(c) $(n-1)s^2p^6, ns^1$ (d) $ns^2p^6d^1$

39. On moving down the group gradually increase

(a) Oxidising property (b) Electronegativity

(c) Acidic property (d) Metallic property

40. An ion which has 18 electrons in the outermost shell is

(a) K^+ (b) Ca^{2+}

(c) Na^+ (d) Cu^+

41. Increasing order of acid strength of halogen acid is **[DCE 2000]**

(a) $HF < HCl < HBr < HI$

(b) $HCl < HBr < HI < HF$

(c) $HF < HI < HBr < HCl$

(d) None of these

42. Which is the weakest base **[KCET 1993]**

(a) $NaOH$ (b) KOH

(c) $Ca(OH)_2$ (d) $Zn(OH)_2$

43. Which of the following elements shows maximum number of different oxidation states in its compounds **[CBSE PMT 1998]**

(a) Eu (b) La

(c) Gd (d) Am

44. The valency shell of calcium contains **[JIPMER 2000]**

(a) 8 electrons (b) 6 electrons

(c) 4 electrons (d) 2 electrons

45. 3 and 6 electrons are present in the outermost orbit of A and B respectively. The chemical formula of its compound will be

(a) A_3B_2 (b) A_2B_3

(c) A_2B (d) AB

46. Which of the following halogens doesn't exhibit positive oxidation state in its compounds **[MH CET 1999]**

(a) Cl (b) Br

(c) I (d) F

47. The most basic element is **[MP PET 2000; JIPMER 2000]**

(a) Fluorine (b) Iodine

(c) Chlorine (d) Bromine

48. Which of the following set has the strongest tendency to form anions **[AFMC 1999]**

(a) Ga, In and Te (b) Na, Mg and Al

(c) N, O and F (d) V, Cr and Mn

49. An element X which occurs in the first short period has an outer electronic structure s^2p^1. What are the formula and acid-base character of its oxides **[DCE 1999]**

(a) XO_3, basic (b) X_2O_3, basic

(c) X_2O_3, amphoteric (d) XO_2, acidic

50. Which of the following gas does not have an octet or eight electrons in the outer shell **[CBSE PMT 2001]**

(a) Ne (b) Ar

(c) Rn (d) He

51. Beryllium and aluminium exhibit many properties which are similar. But, the two elements differ in **[AIEEE 2**

(a) Forming covalent halides

(b) Forming polymeric hydrides

(c) Exhibiting maximum covalency in compounds

(d) Exhibiting amphoteric nature in their oxides

Critical Thinking

Objective Questions

1. Which of the following statement is correct with respect to the property of elements with an increase in atomic number in the carbon family (group 14) **[BHU 2004]**

(a) Atomic size decrease

(b) Ionization energy increase

(c) Metallic character decrease

(d) Stability of +2 oxidation state increase

2. The pair of amphoteric hydroxides is **[AIIMS 2005]**

(a) $Al(OH)_3, LiOH$

(b) $Be(OH)_2, Mg(OH)_2$

(c) $B(OH)_3, Be(OH)_2$

(d) $Be(OH)_2, Zn(OH)_2$

3. Which of the following oxides is amphoteric in character **[AIEEE 2005]**

(a) CaO (b) CO_2

(c) SiO_2 (d) SnO_2

4. Which has highest melting point **[RPMT 1997]**

(a) $LiCl$ (b) $BeCl_2$

(c) BCl_3 (d) CCl_4

5. Arrange S, O and Se in ascending order of electron affinity **[Roorkee 1990]**

(a) $Se < S < O$ (b) $Se < O < S$

(c) $S < O < Se$ (d) $S < Se < O$

6. Which of the following is not the correct increasing order of ionisation energy **[RPMT 2000]**

(a) $Cl^- < Ar < K^+$ (b) $Au < Ag < Cu$

(c) $Cs < Rb < K$ (d) $K < Ca < Sc$

7. In which of the following arrangements the order is NOT according to the property indicated against it **[AIEEE 2005]**

(a) $Al^{3+} < Mg^{2+} < Na^+ < F^-$: Increasing ionic size

(b) $B < C < N < O$: Increasing first ionization enthalpy

(c) $I < Br < F < Cl$: Increasing electron gain enthalpy (with negative sign)

(d) $Li < Na < K < Rb$: Increasing metallic radius

8. Which element has the greatest tendency to loose electrons **[NCERT 1980]**

(a) F (b) S

(c) Fe (d) Be

9. Strongest acid is **[RPMT 1997]**

(a) Al_2O_3 (b) MgO

(c) Na_2O (d) CaO

10. Which one of the following arrangements represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species **[CBSE PMT 2005]**

(a) $Cl < F < S < O$ (b) $O < S < F < Cl$

(c) $S < O < Cl < F$ (d) $F < Cl < O < S$

11. Increasing order of electronegativity is **[RPET 2003]**

(a) $Bi < P < S < Cl$ (b) $P < Bi < S < Cl$

(c) $S < Bi < P < Cl$ (d) $Cl < S < Bi < P$

12. What will be the order of Ist ionisation energy **[BHU 2005]**

(a) $Li > Na > K$ (b) $K > Li > Na$

(c) $Na > Li > K$ (d) $Li > K > Na$

13. Which of the following configurations represents atoms of the elements having the highest second ionization energy **[Pb. PMT 1998]**

(a) $1s^2 2s^2 2p^4$ (b) $1s^2 2s^2 2p^6$

(c) $1s^2 2s^2 2p^6 3s^1$ (d) $1s^2 2s^2 2p^6 3s^2$

14. The first ionization potentials in electron volts of nitrogen and oxygen atoms are respectively given by **[IIT 1987]**

(a) 14.6, 13.6 (b) 13.6, 14.6

(c) 13.6, 13.6 (d) 14.6, 14.6

15. The elements which occupy the peaks of ionisation energy curve, are **[CBSE 2000]**

(a) Na, K, Rb, Cs (b) Na, Mg, Cl, I

(c) Cl, Br, I, F (d) He, Ne, Ar, Kr

16. Which is the correct order of ionic sizes (At. No. : Ce = 58, Sn = 50, Yb = 70 and Lu = 71) **[AIEEE 2002]**

(a) $Ce > Sn > Yb > Lu$ (b) $Sn > Ce > Lu > Yb$

(c) $Lu > Yb > Sn > Ce$ (d) $Sn > Yb > Ce > Lu$

17. A sudden large jump between the values of second and third ionisation energies of an element would be associated with the electronic configuration **[CBSE PMT 1992; AFMC 1998; CPMT 1999]**

(a) $1s^2, 2s^2p^6, 3s^1$ (b) $1s^2, 2s^2p^6, 3s^2p^1$

(c) $1s^2, 2s^2p^6, 3s^2p^2$ (d) $1s^2, 2s^2p^6, 3s^2$

18. Which element having following electronic configurations has minimum ionization potential **[NCERT 1978; KCET 1991; CBSE PMT 1991; Pb. PET 1999; BHU 2000]**

(a) $1s^1$ (b) $1s^2, 2s^2\, 2p^6$

(c) $1s^2, 2s^2\, 2p^6, 3s^1$ (d) $1s^2, 2s^2\, 2p^2$

19. Arrange F, Cl, O, N in the decreasing order of electronegativity

(a) $O > F > N > Cl$ (b) $F > N > Cl > O$

(c) $Cl > F > N > O$ (d) $F > O > N > Cl$

20. Ionic radii of **[IIT-JEE 1999]**

(a) $Ti^{4+} < Mn^{7+}$ (b) ${}^{35}Cl^- < {}^{37}Cl^-$

(c) $K^+ > Cl^-$ (d) $P^{3+} > P^{5+}$

21. Which of the following have high electron affinity **[BHU 2000, 05]**

(a) F (b) Cl

(c) N (d) O

22. In which block 106th element belongs **[DCE 2000]**

(a) *s*-block (b) *p*-block

(c) *d*-block (d) *f*-block

Assertion & Reason

For AIIMS Aspirants

Read the assertion and reason carefully to mark the correct option out of the options given below :

(a) If both assertion and reason are true and the reason is the correct explanation of the assertion.

(b) If both assertion and reason are true but reason is not the correct explanation of the assertion.

(c) If assertion is true but reason is false.

(d) If the assertion and reason both are false.

(e) If assertion is false but reason is true.

1. Assertion : Positive ions will be wider than parent ato

Reason : Nuclear charge pulls them closer

2. Assertion : Dinegative anion of oxygen (O^{2-}) is quite common but dinegative anion of sulphur (S^{2-}) is less common

Reason : Covalency of oxygen is two **[AIIMS 2002]**

3. Assertion : The atomic radii of calcium is smaller than sodium.

Reason : Calcium has a lower nuclear charge than sodium **[AIIMS 1999]**

4. Assertion : The first ionization energy of *Be* is greater than that of *B*

Reason : $2p$ orbital is lower in energy than $2s$ of **[IIT-JEE Screening 2000]**

5. Assertion : $LiCl$ is predominantly a covalent compound

Reason : Electronegativity difference between Li and Cl is too small **[IIT-JEE 1998]**

6. Assertion : F atom has a less negative electron affinity than Cl atom

Reason : Additional electrons are repelled more effectively by $3p$ electrons in Cl atom than by $2p$ electrons in F atom **[IIT-JEE 1998]**

7. Assertion : Noble gases have maximum electron affinity. **[AIIMS 1995]**

Reason : High electron affinity shows that the electron is loosely bonded to the atom.

8. Assertion : The first ionisation energy of *Be* is greater than boron **[AIIMS 2002]**

Reason : $2p$ orbitals have lower energy than $2s$ orbitals.

9. Assertion : Atomic number of the element unumbium is 112.

Reason : Name for digits 1 and 2 is un-and bi-respectively in latin words.

10. Assertion : Chemistry of Actinoids is more complicated than Lanthanoids.

Reason : Actinoid elements are radioactive.

11. Assertion : Ionization enthalpy is always negative.

Reason : Energy is always released when electrons are removed.

12. Assertion : Shielding effect increases as we go down the group.

Reason : More is the number of electrons in the penultimate shell, more is shielding.

13. Assertion : Ionization potential across the period is $Na < Al < Mg < Si$.

Reason : Ionization potential decreases with decrease in atomic size.

14. Assertion : More is the electron affinity greater is the reducing character.

Reason : Reducing character depends on number of electrons gained.

15. Assertion : Ground state configuration of *Cr* is $3d^5, 4s^1$.

Reason : A set of half filled orbitals containing one electron each with their spin parallel provides extra stability to the system.

16. Assertion : I.E. of $_7N$ is more than that of $_8O$ as well as $_6C$.

Reason : This is due to difference in reactivity towards oxygen.

17. Assertion : NO^- ion is isoelectronic with CN^- ion.

Reason : Isoelectronic ions have same number of elelctrons.

18. Assertion : Outermost electronic configuration of most electropositive elements is $ns^2\ np^3$.

Reason : $ns^2\ np^3$ is stable due to half filled subshell.

19. Assertion : First ionization energy for nitrogen is lower than oxygen.

Reason : Across a period effective nuclear charge decreases. **[AIIMS 2005]**

ANSWERS

Extended or long form of periodic table

1	c	2	d	3	c	4	a	5	c
6	c	7	b	8	b	9	b	10	a
11	d	12	a	13	d	14	d	15	b
16	a	17	d	18	c	19	a	20	b
21	a	22	a	23	b	24	a	25	c
26	a	27	a	28	a	29	b	30	d
31	c	32	c	33	c	34	a	35	b
36	b	37	d	38	c	39	a	40	d
41	a	42	c	43	d	44	d	45	c
46	a	47	b	48	b	49	c	50	c
51	b	52	b	53	c	54	d	55	b
56	c	57	c	58	b	59	a	60	d
61	b	62	b	63	a	64	c	65	a
66	d	67	d	68	a	69	c	70	d
71	c	72	b	73	c	74	c	75	a
76	d	77	a	78	b	79	d	80	c
81	b	82	d	83	b	84	b	85	b
86	d	87	a	88	d	89	c	90	b
91	d	92	b	93	c	94	a	95	d
96	d	97	c	98	b	99	a	100	c
101	c	102	c	103	abcd	104	c	105	d
106	a	107	d	108	a	109	a	110	b
111	b	112	a	113	b	114	a	115	d
116	c	117	c	118	c	119	a	120	c
121	d	122	b	123	a	124	b	125	b
126	b	127	d	128	c	129	b	130	d
131	d								

Atomic and Ionic radii

1	b	2	d	3	c	4	a	5	a
6	d	7	a	8	b	9	b	10	d
11	a	12	d	13	d	14	a	15	a
16	d	17	d	18	d	19	a	20	a
21	b	22	d	23	a	24	a	25	c
26	d	27	c	28	b	29	c	30	c
31	c	32	b	33	d	34	a	35	d
36	a	37	d	38	a	39	b	40	d
41	b	42	a	43	a	44	d	45	a
46	d	47	a	48	c	49	d	50	a
51	b	52	a	53	d	54	a	55	c
56	b	57	c	58	b	59	c	60	b
61	a	62	d	63	c	64	a	65	d
66	d	67	d	68	c	69	c	70	c
71	b	72	a						

Ionisation energy

1	b	2	c	3	d	4	a	5	c
6	a	7	c	8	d	9	a	10	a
11	b	12	c	13	a	14	d	15	b
16	a	17	a	18	c	19	a	20	b
21	d	22	b	23	d	24	c	25	b
26	c	27	b	28	a	29	c	30	b
31	a	32	b	33	a	34	c	35	c
36	c	37	d	38	b	39	c	40	a
41	c	42	b	43	a	44	b	45	b
46	d	47	a	48	b	49	a	50	a
51	a	52	d	53	a	54	a	55	c
56	b	57	a	58	a	59	d	60	a
61	d	62	d	63	a	64	b	65	b
66	c	67	a	68	b	69	a	70	c
71	a	72	b	73	d	74	b	75	d
76	b	77	c	78	c	79	d	80	a
81	d	82	b	83	b	84	a	85	d
86	c	87	a	88	b	89	c		

Electron affinity

1	d	2	c	3	c	4	d	5	a
6	a	7	b	8	b	9	a	10	b
11	c	12	a	13	c	14	b	15	a
16	d	17	b	18	a	19	b	20	c
21	b	22	b	23	d	24	c	25	d
26	a	27	b						

Electronegativity

1	b	2	b	3	d	4	b	5	b
6	a	7	b	8	a	9	c	10	a
11	c	12	b	13	d	14	a	15	c
16	a	17	b	18	c	19	a	20	a
21	d	22	a	23	d	24	b	25	a
26	b	27	c	28	b	29	b	30	a
31	b	32	b	33	a	34	a	35	d

Valency and oxidation state

1	a	2	c	3	c	4	a	5	a
6	d	7	a	8	d	9	b	10	d
11	d	12	c	13	a	14	c	15	d
16	d	17	a	18	c	19	a	20	b
21	a	22	b	23	b	24	b	25	d
26	c	27	b	28	c	29	d	30	b
31	d	32	d	33	b	34	c	35	a
36	b	37	d	38	c	39	d	40	d
41	a	42	d	43	d	44	d	45	b
46	d	47	b	48	c	49	c	50	d
51	d								

Critical Thinking Questions

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

Assertion & Reason

1	d	2	b	3	c	4	c	5	c
6	c	7	d	8	a	9	a	10	b
11	d	12	a	13	c	14	e	15	a
16	c	17	e	18	e	19	b		

Answers

Extended or long form of periodic table

1	c	2	d	3	c	4	a	5	c
6	c	7	b	8	b	9	b	10	a
11	d	12	a	13	d	14	d	15	b
16	a	17	d	18	c	19	a	20	b
21	a	22	a	23	b	24	a	25	c
26	a	27	a	28	a	29	b	30	d
31	c	32	c	33	c	34	a	35	b
36	b	37	d	38	c	39	a	40	d
41	a	42	c	43	d	44	d	45	c
46	a	47	b	48	b	49	c	50	c
51	b	52	b	53	c	54	d	55	b
56	c	57	c	58	b	59	a	60	d
61	b	62	b	63	a	64	c	65	a
66	d	67	d	68	a	69	c	70	d
71	c	72	b	73	c	74	c	75	a
76	d	77	a	78	b	79	d	80	c
81	b	82	d	83	b	84	b	85	b
86	d	87	a	88	d	89	c	90	b
91	d	92	b	93	c	94	a	95	d
96	d	97	c	98	b	99	a	100	c
101	c	102	c	103	abcd	104	c	105	d
106	a	107	d	108	a	109	a	110	b
111	b	112	a	113	b	114	a	115	d
116	c	117	c	118	c	119	a	120	c
121	d	122	b	123	a	124	b	125	b
126	b	127	d	128	c	129	b	130	d
131	d								

Atomic and Ionic radii

1	b	2	d	3	c	4	a	5	a
6	d	7	a	8	b	9	b	10	d
11	a	12	d	13	d	14	a	15	a
16	d	17	d	18	d	19	a	20	a
21	b	22	d	23	a	24	a	25	c
26	d	27	c	28	b	29	c	30	c
31	c	32	b	33	d	34	a	35	d
36	a	37	d	38	a	39	b	40	d
41	b	42	a	43	a	44	d	45	a
46	d	47	a	48	c	49	d	50	a
51	b	52	a	53	d	54	a	55	c
56	b	57	c	58	b	59	c	60	b
61	a	62	d	63	c	64	a	65	d
66	d	67	d	68	c	69	c	70	c
71	b	72	a						

Ionisation energy

1	b	2	c	3	d	4	a	5	c
6	a	7	c	8	d	9	a	10	a
11	b	12	c	13	a	14	d	15	b
16	a	17	a	18	c	19	a	20	b
21	d	22	b	23	d	24	c	25	b
26	c	27	b	28	a	29	c	30	b
31	a	32	b	33	a	34	c	35	c
36	c	37	d	38	b	39	c	40	a
41	c	42	b	43	a	44	b	45	b
46	d	47	a	48	b	49	a	50	a
51	a	52	d	53	a	54	a	55	c
56	b	57	a	58	a	59	d	60	a
61	d	62	d	63	a	64	b	65	b
66	c	67	a	68	b	69	a	70	c
71	a	72	b	73	d	74	b	75	d
76	b	77	c	78	c	79	d	80	a
81	d	82	b	83	b	84	a	85	d
86	c	87	a	88	b	89	c		

Electron affinity

1	d	2	c	3	c	4	d	5	a
6	a	7	b	8	b	9	a	10	b
11	c	12	a	13	c	14	b	15	a
16	d	17	b	18	a	19	b	20	c
21	b	22	b	23	d	24	c	25	d
26	a	27	b						

Electronegativity

1	b	2	b	3	d	4	b	5	b
6	a	7	b	8	a	9	c	10	a
11	c	12	b	13	d	14	a	15	c
16	a	17	b	18	c	19	a	20	a
21	d	22	a	23	d	24	b	25	a
26	b	27	c	28	b	29	b	30	a

31	b	32	b	33	a	34	a	35	d

Valency and oxidation state

1	a	2	c	3	c	4	a	5	a
6	d	7	a	8	d	9	b	10	d
11	d	12	c	13	a	14	c	15	d
16	d	17	a	18	c	19	a	20	b
21	a	22	b	23	b	24	b	25	d
26	c	27	b	28	c	29	d	30	b
31	d	32	d	33	b	34	c	35	a
36	b	37	d	38	c	39	d	40	d
41	a	42	d	43	d	44	d	45	b
46	d	47	b	48	c	49	c	50	d
51	d								

Critical Thinking Questions

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

Assertion & Reason

1	d	2	b	3	c	4	c	5	c
6	c	7	d	8	a	9	a	10	b
11	d	12	a	13	c	14	e	15	a
16	c	17	e	18	e	19	b		

Answers and Solutions

Extended or long form of periodic table

2. (d) n/p ratio is a cause of radioactivity.

5. (c) Halogens has 7 electrons in his valance shell (ns^2np^5).

6. (c) As alkali metals have tendency to loose e^-.

7. (b) Each period consists of a series of elements whose atom have the same principal quantum no. (n) of the outer most shell i.e. In second period $n=2$, this shell has four orbitals (one 2s and three $2p$) which can have eight electrons, hence second period contains 8 elements from atomic no. 3 to 10.

9. (b) Neils Bohr developed the long form of periodic table on the basis of Mosley's principle.

10. (a) $33 - 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$

11. (d) 16-$1s^2 2s^2 2p^6 3s^2 3p^4$ there are $6e^-$ in outer most shell therefore its group is VIthA.

13. (d) Many metals with catalytic properties because

(i) They provide surface area for reaction to occur

(ii) They decreases the ionisation energy.

(iii) They have vacant d-orbitals.

17. (d) Aluminium. As it belongs to p-block element.

18. (c) $Cu_{29} - [Ar]\ 3d^{10}4s^1$.

21. (a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ principal quantum no. is 4 so it belongs to 4th period.

23. (b) Inert gases, these have ns^2np^6 configuration.

24. (a) $1s^2 2s^2 2p^2 -$ there are $4e^-$ in valence shell therefore it goes to IV- group.

27. (a) $U > Ra > Pb > Hg$

28. (a) $Mg - Ba$. Both belongs to II-A group.

29. (b) $Na - Cl$. Both belongs to III period.

30. (d) Elements of second and third period

Diagonal relationship

$$\begin{tikzcd} \text{II} & Li \arrow[dr] & Be \arrow[dr] & B \arrow[dr] & C \arrow[dr] & N \arrow[dr] & O \arrow[dr] & F \\ \text{III} & Na & Mg & Al & Si & P & S & Cl \end{tikzcd}$$

31. (c) d-Block because the last electron enters d-subshell.

34. (a) Kr has atomic no. 36 which is a noble gas and all noble gases are included in the p-block.

38. (c) d-block. As the last e^- enters in d-subshell.

40. (d) Due to its vacant p-orbital.

41. (a) By obserbing principal quantum number (n), Orbital (s, p, d, f) and equating no. of e^-'s we are able to find the period, block and group of element in periodic table.

42. (c) $33 : 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^3$

In its valence shell $5e^-$ are present so it is fifth (A) group element.

43. (d) 38 is the atomic no. of strontium (Sr) which is s-block element and all the elements of s-block are metals.

44. (d) Hydrogen resembles alkali metals in some properties so it can be placed in the first group of periodic table.

47. (b) Chalcons are oxygen family.

49. (c) Both belongs to VA group.

50. (c) According to Dobernier law of triads the atomic mass of the central element was nearly the arithmetic mean of atomic masses of other two elements.

Cl	*Br*	*I*	Arithmetic mean
31	75	120	$\frac{120+31}{2}=75.5$

53. (c) Z = 2,8,8,1. ∵ it would donate e^- more easily.

54. (d) Last electron goes to *s*-subshell.

55. (b) Because they belong to same group.

57. (c) Ionic radius will increase as number of shells increases

58. (b) *Al*. Due to diagonal relationship.

60. (d) 2,8,2. ∵ it would donate e^- more easily.

62. (b) A representative element as last e^- enters *p*-orbital.

63. (a) The configuration represents on alkaline earth metals.

65. (a) First group

e.g. $NaCl + H_2O \rightleftharpoons NaOH + HCl$

68. (a) Ionic bond is formed when there is large difference of electro-negativities between the atoms.

69. (c) *d*-block - $[Ar]3d^1 4s^2$

70. (d) Be : $1s^2\, 2s^2$

71. (c) Increasing atomic number. Mosley found that atomic no. was better fundamental property than atomic weight.

72. (b) Lowest ionisation energy due to largest size.

73. (c) Elements on the right side of the periodic table are *p*-block. Mostly non-metals.

74. (c) Screening effect of *d* and *f* block elements is nearly same.

77. (a) Li because of its smallest size.

78. (b) In third group *Na* is a typical element.

85. (b) On equating no. of e^-'s atomic no. is 12 which is for Mg .

86. (d) $17 - 1s^2 2s^2 2p^6 3s^2 3p^5$.

89. (c) Lanthanide's are called rare earth metals.

91. (d) It show similarities with both alkali metals as well as halogens.

92. (b) M^- After gaining an e^- the metal attains stable configuration.

95. (d) Due to presence of vacant *d*-orbitals and they show *d*-*d* transition.

96. (d) Potassium, $K - [Ar]4s^1$.

97. (c) *p*-block; ${}_{31}Ga \rightarrow [Ar]3d^{10} 4s^2 p^1$.

102. (c) *Mg* has only two electrons in the 3*s*-orbital and hence its I.E. is lowest, i.e. it has the maximum tendency to form di-positive ions.

103. (a,b,c,d) It reflects trends in physical and chemical properties of the elements.

104. (c) As last e^- goes to *d*-subshell.

107. (d) First decreases to a minimum and then increases

108. (a) ${}_{25}Mn - 3d^5 4s^2$.

111. (b) Hydrogen, forms hydrides like halides, e.g. *HCl*.

114. (a) Hydration energy increases along the period.

115. (d) In IIA group all elements are metal while in IIIA, IVA and VIIA groups non-metallic elements are also present.

118. (c) Mg, Ba, Ca have ns^2 configuration.

119. (a) Elements of group halogen are : F, Cl, Br I and At .

121. (d) N and P have 3 unpaired electrons in $2p$ and $3p$ respectively; V has 3 unpaired electrons in $3d$.

124. (b) Tungsten (W) having highest m.p.

125. (b) These atomic no. gives the configuration ns^2np^5 which are of halogen group or VIIth group.

126. (b) The atomic no. of an element is derived from the no. of proton because during chemical reaction no. of electron undergoes for change

127. (d) Due to identical ionic radii and polarising power

$$\left[\frac{\text{Charge}}{\text{Size}}\text{ ratio of pairs of these elements}\right]$$

Atomic and ionic radii

1. (b) Value of Z for hydrogen =1

Value of Z for helium = 2

Value of *n* for both is = 1

$$r_H = \frac{0.52\times 1^2}{1} \qquad r_{He^+} = \frac{0.52\times 1^2}{1}$$

$$\frac{r_H}{r_{He^+}} = 1:1 \text{ or } r_{He^+} : r_H = 1:1$$

2. (d) The size of an species decreases with increasing nuclear charge because the attraction for the electrons increases. Thus Al^{3+} is smaller in size.

3. (c) As the nuclear charge per electron is maximum in F^-. Therefore it is smallest in size.

4. (a) During the formation of cation the size decreases.

6. (d) Highest the nuclear charge smallest the atomic size as well as radius also.

7. (a) Atomic radius decreases on going from left to right in a period. Thus size of $O > F$. As O^{2-} and F^- are isoelectronic, therefore, size of $O^{2-} > F^-$.

8. (b) As the nuclear charge per e^- is maximum in Mg^{+2}, it has smallest size among Na^+, Mg^{+2}, Cl^- and F^-.

9. (b) S^{2-} and Cl^- both are isoelectronic but nuclear charge of Cl^- is more than S^{2-}. So it has largest size.

10. (d) In completely filled shell inter atomic repulsion is more so have greater size.

12. (d) I^- as it has the biggest size.

13. (d) *Mg*, as we move across the period atomic radius decreases.

14. (a) O^{-2} has the highest value of ionic radii as this can be explained on the basis of

$$Z/e\left\{\frac{\text{Nucleaus charge}}{\text{No. of electron}}\right\}$$

Whereas *Z* / *e* ration increases, the size decreases and when *Z* / *e* ration decreases the size increases.

15. (a) Continuous increase as no. of shells increases down the group.

16. (d) $Na^+ < F^- < O^{2-} < N^{3-}$

All are isoelectronic, effective number charge is highest for Na^+ so it has smallest size.

17. (d) $\underset{54}{I^-} > \underset{53}{I} > \underset{52}{I^+}$ atmoic number

19. (a) Continuously decreases as the effective nuclear charge increases.

20. (a) $Mg^{2+} < Na^+ < F^- < Al$

F^- has bigger size than Mg^{2+} and Na^+ due to small nuclear charge.

21. (b) More than F^- as K^+ has more no of shells in atomic state.

22. (d) All are isoelectronic but O^{2-} has lowest charge among them. So it is largest in size.

23. (a) As effective nuclear charge on Na^+ is maximum. It has smallest size.

25. (c) $Be > C > F > Ne$. Atomic size decreases across a period.

26. (d) As the nuclear charge per electron is maximum in P^{5+}. Therefore its size is smallest.

27. (c) $Na^+ - 10$ electron ; $Li^- -$ 4 electron

28. (b) Ionic radius of trivalent lanthanide's almost remains constant with increase in the atomic number.

30. (c) Halogens are most electronegative elements.

33. (d) On moving from left to right in a period value of radius decreases.

34. (a) $Sc^{3+} > Cr^{3+} > Fe^{3+} > Mn^{3+}$ the correct order is $Cr^{+3} > Mn^{+3} > Fe^{+3} > Sc^{+3}$

35. (d) $Na^+ > Mg^{2+} > Al^{3+} > Si^{4+}$. All are isoelectronic but nuclear charge per electron is greatest for Si^{4+}. So it has smallest size and nuclear charge per electron for Na^+ is smallest. So it has largest size.

36. (a) $N^{3-} > O^{2-} > F^-$. All are isoelectronic but nuclear charge per electron is highest for F^-, so it has smallest size.

38. (a) Cation has small size than parent atom and anion has greater size than parent atom.

39. (b) Ionic radii decreases significantly from left to right in a period among representative elements.

40. (d) H^- is most stable due to its full filled 1s-orbital.

43. (a) C^{4-} has largest radius due to least nuclear charge per electron.

44. (d) For ionic bond formation low I.E., high electron affinity and high lattice energy is needed.

45. (a) Ionic radii increases in a group.

46. (d) Size of elements decreases across a period.

47. (a) X^- ion larger in size than X atoms. Because of low effective nuclear charge on X^-, X has a bigger size.

48. (c) Fe, Co, Ni, Cu. Due to shielding of *d*-electrons, the effect of increased nuclear charge due to increase in atomic no. neutralised. Consequently atomic radius remains almost unchanged after chromium.

49. (d) $S^{2-} > Cl^- > K^+ > Ca^{2+}$

52. (a) Covalent radii decreases on going from left to right in periods. However among the transition elements the size do not changes much because the electrons add to the pneultimate *d*-subshells *i.e.* $(n-1)d$ - subshell.

59. (c) During the conversion of neutral atom to cation size decreases because after removal one e^- or more

(i) Nuclear charge per electron increases.

(ii) Outermost shell is completely removed.

60. (b) Atomic radius increases as no. of shells increases.

62. (d) Chloride ion and potassium ion are isoelectronic, isoelectronic ions are those ions having same number of electrons.

$K = 2, 8, 8, 1$ $\quad K^+ = 2, 8, 8$

$Cl = 2, 8, 7$ $\quad Cl = 2, 8, 8$

63. (c) Cs^+ has the largest ionic radius in the periodic table.

64. (a) Ionic radii increases down the group.

65. (d) Si^{4+} is smallest in size due to their greater +ve charge.

66. (d) Due to having three electrons atomic size increases.

$F^- = 9 + 1 = 10$ electrons ; $O^{-2} = 8 + 2 = 10$ electrons

$Al + 3 = 13 - 3 = 10$ electrons ; $N^{3-} = 7 + 3 = 10\, e^-$.

Because electrostatic force between nucleus and $\bar{e}$ cloud is least in nitrogen.

67. (d) The trivalent ion having largest size in lanthanide series is lanthanum. This is due to lanthanide contraction.

68. (c) As we know that hydration power decreases on moving down the group hence among alkali metals *Li* has excessive hydration & hence it has low mobility in aqeous solution.

69. (c) Ionic radius in the n^th orbit is given by $r_n = \frac{n^2 a}{z}$

or $r_n \propto \frac{1}{Z}$ where *n* is principal equation no., ao. bohr's radius of hydrogen atom and *Z* is the effective nuclear energy.

70. (c) Order of polarising power $Be^{++} > Li^+ > Na^+$

Hence order of covalent character $BeCl_2 > LiCl > NaCl$

71. (b) Higher the (*n*+1) value higher is the energy associated with orbitals.

72. (a) With the increase in size of cation the size of the hydrated ion decreases hence ionic conductance increases.

Ionisation energy

1. (b) I.E.(II) of Na is higher than that of Mg because in case of Na, the second e^- has to be remove from the noble gas core while in case of Mg removal of second e^- gives a noble gas core.

Mg has high first ionisation potential than Na because of its stable ns^2 configuration.

7. (c) Ionization potential decreases. Since, atomic size increases.

8. (d) Alkali metals, lower the no. of valence e^-, lower is the value of ionization potential.

9. (a) The ionization energy of hydrogen is to high for group of alkali metals, but too low for halogen group.

13. (a) $E_1 < E_2$ because second I.E. is greater than first I.E.

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

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

18. (c) The energy required to remove an electron from outermost orbit of an isolated gaseous atom is called I.E. Now carbon has $4e^-$ in outermost shell. Thus it has 4 ionization energies.

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

21. (d) First I.P. of $Be > B$ because of stable ns^2 configuration.

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

24. (c) Since the IV, I.E. is very high. Thus electron is to be removed from stable configuration.

25. (b) Li and Cs belong to I^st group but Cs has larger size, hence low nuclear attraction force, thus low ionization energy.

26. (c) Li belongs to I^st group. There is $1e^-$ in outermost shell. Thus low I.E.

27. (b) Increases from left to right. Since, the size decreases.

28. (a) As the e^- is to be removed from stable configuration.

29. (c) Since e^- is to be removed from exactly half filled *p*-orbital.

31. (a) Ionisation potential increases across the period.

32. (b) $E = \frac{E_0}{n^2} = \frac{-54.4}{4} = -13.6\, eV$

34. (c) Due to stable half-filled orbitals.

35. (c) Greater than the first ionization energy because after removal one e^-, effective nuclear charge increases.

36. (c) Rare gases as the e^- is to removed from stable electron configuration.

37. (d) Since it is a noble gas.

38. (b) The first *I.P.* is maximum for hydrogen due to its small size.

41. (c) Due to his fullfilled configuration.

42. (b) 1st I.P. decreases down the group.

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

45. (b) First I.P. for C is 11.3, for N is 14.5 and for O is 13.6

47. (a) Li has least I.P about 5.4.

48. (b) I.E. increases across the period.

50. (a) He has highest ionisation energy due to it full fill 1s-orbital.

51. (a) s-electrons are strongly bonded to the nucleus. So large amount of energy is required to remove an e^-.

52. (d) $Mg > Al > Na$. This is due to the presence of fully filled s-orbital in Mg.

55. (c) The 1st I.P. for hydrogen is 13.6 *volts*

56. (b) Alkali metals are strong reducing agents

58. (a) Due to the large size of group IA elements, the outermost electron is far from the nucleus and can easily be removed. their ionisation energies or ionisation potentials are relatively low.

	Li	Na	K	Rb	Cs
Ionisation potential (eV)	5.4	5.1	4.3	4.2	3.9

60. (a) $N > O > Be > B$ Ist ionisation energy of $N > O$ because of half filled p-orbital.

61. (d) $M^{2+} \rightarrow M^{3+}$ After the removal of $2e^-$ the nuclear charge per e^- increases, due to which high energy is required to remove $3e^-$.

63. (a) I.E. increases from left to right in a period.

64. (b) More because of stable configuration of Mg.

65. (b) He and Xe belongs to same group but He has higher ionisation energy because of small size.

66. (c) In second transition electron is to be removed from half filled orbital.

68. (b) As it belongs to IA group and has maximum size.

69. (a) Since, they have larger size as compared to other.

70. (c) The second I.E. is greater than first I.E. similarly second E.A. is greater than first E.A. the energy is to be supplied to force the second e^- into the anion.

71. (a) Increases as the atomic size decreases and hence effective nuclear charge increases.

72. (b) B, Be, C, N as I.E. increases across the period.

73. (d) Ionization potential is least for alkali metals and it decreases down the group.

74. (b) It has maximum ionization energy due to half filled orbitals.

75. (d) It has maximum no. of e^- in outermost shell. So it has maximum I.E.

76. (b) Ionization potential increases as we go from left to right in a period, while it decreases as we come down a group.

Be	B	Li	Na
9.3	8.3	5.4	5.1

77. (c) Half filled p-orbitals possess extra stability.

78. (c) Ionization potential decreases down the group.

79. (d) Li^+ and Mg^{+2} ions have similar polarising power or ionic potential and therefore have similar properties. This type of relationship of the first element of a group with the second element of the next group is known as diagonal relationship.

80. (a) The addition of second electron in an atom or ion is always endothermic.

81. (d) We know that ionisation potential gradually decreases of moving down the group while atomic size increases as we move down the group. Hence larger the atomic size, smaller is ionisation potential.

82. (b) Fluorine has highest E^o red {Equal to +2.9 V} due to which it can easily accept an electron & hence it is the best oxidising agent.

83. (b) The ionisation energy of tin {Sn} is less than that of lead (Pb). It is due to the poor shielding of d- and f-electron in Pb due to which it feels greater attraction from nucleus.

84. (a) The order of screening effect in a given shell are in order $s > p > d > f$.

85. (d) The ionisation energy of Li, Be, B and C is 520, 899, 801, 1086 $kJ\ /\ mol$ respectively hence, carbon has highest IE_1.

86. (c) Isoelectronic species are those which have same no. of electrons.

$K^+ = 19 - 1 = 18$; $Ca^{+2} = 20 - 2 = 18$

$Sc^{+3} = 21 - 3 = 18$; $Cl^- = 17 + 1 = 18$

87. (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 reactivity of halogens is $F > Cl > Br > I$

88. (b) Ionisation potential generally increases when we more in a period from left to right but IE_1 of N_2 is greater than that of O_2. It is due to the more stable (half-filled orbitals) configurations of N.

89. (c) Nitrogen has more ionisation potential than carbon & oxygen because, if outermost orbit is half filled so it is more stable & order is $C < N > O$

Electron affinity

3. (c) $\underset{140.9}{O} > \underset{122.3}{C} > \underset{83}{B} > \underset{0}{N}$ Value of electron affinity increases on going from left to right in periods but the value of electron affinity of Vth A elements is less than that of IVth A element, this is due to half filled *p*-orbitals presence.

4. (d) Halogens have maximum electron affinity due to their smaller size.

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

8. (b) Energy released when an electron is added to an isolated atom in gaseous state.

9. (a) Electron affinity value of *Cl* is greater the *F* and then decreases down the group.

10. (b) Electron affinity increases across the period.

13. (c) Electron affinity of chorine is maximum.

14. (b) The formation of ionic bond depends upon easy formation of cation and anion. therefore the ionisation energy value of the metal atom should be low, so that it can easily form cation. on the other hand, the electron affinity value of the non-metal atom should be high so that it can easily form anion.

15. (a) Because it can easily accept an e^-.

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

19. (b) In IB group all elements are metals.

22. (b) Flourine although have highest electronegativity due to its very small size, effective inter electronic repulsions are observed which brings down its electron affinity.

23. (d) The bond dissociation energy of *F-F* bond is very low. The weak *F-F* bond makes fluorine the strongest oxidising halogen.

24. (c) Atomic radius increase from top to bottom in a group while decrease from left to right in a period on the other hand electron affinity shows severe trends i.e. decrease from top to bottom in a group and increase from left to right in a period.

25. (d) It is a fact.

26. (a) Electron affinity of *Cl* is greater than fluorine so the order are as $F < Cl > Br > I$

27. (b) Halogens have very high electron affinity. It may be rated that the electron affinity of fluorine is unexpectedly low $(< Cl)$. This may perhaps be due to small size of *F* atom. The value of electron gain enthalpies for Cl, F, S and *O* are respectively 349, 333, 200 & 142 *KJ / mol* hence correct order is $Cl > F > S > O$

Electronegativity

2. (b) Decrease as atomic size increases.

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

5. (b) An atom with high electronegativity has high I.P.

6. (a) If electronegativity difference is greater than 1.7 bond is ionic, if less than 1.7, the bond is covalent.

7. (b) Due to decrease in hydration energy of cation and lattice energy remains almost unchanged.

8. (a) *F*, because of its smallest size.

9. (c) Because of small size and high nuclear charge.

10. (a) Electronegativity decreases down the group.

11. (c) Halogens are most electronegative.

12. (b) Electronegativity decreases down the group.

13. (d) Because of smallest size.

14. (a) Electronegativity decreases down the group.

16. (a) Electronegativity increases since the size decreases.

17. (b) Electropositive character decreases across the period as metallic character decreases.

18. (c) *Si, P, S*. As across the period electronegativity increases.

19. (a) Both electronegativity and electron affinity increases. This is because decrease in the size and increase in the nuclear charge. But electronegativity increases continuously.

20. (a) Electropositive nature increases down the group.

21. (d) Electropositive nature increases down the group.

23. (d) The electronegative character increases as the size decreases.

24. (b) Electronegativity increases across a period.

25. (a) $Li_3 - 1s^2 2s^1$ donates $1e^-$ easily.

28. (b) Electronegativity decreases down the group as atomic radius increases.

30. (a) Electronegativity increases across the period because size decreases.

31. (b) Alkali metals are most electropositive and moreover, electropositive character increases down the group.

32. (b) Electronegativity increases when moves towards period & decrease when moves toward group.

33. (a) Electronegativity is the property of a bonded atom. The relative tendency on an atom to attract the shared pair of electron toward itself is called electronegativity.

34. (a) Due to Raving small in size and electron defficient in nature it has highest polarising ability we can use Fazan's rule to understand it further.

35. (d) With decrease in size from *Al* to *S* the basic nature of oxide decrease and acidic nature increases.

$Al_2O_3 < SiO_2 < P_2O_3 < SO_2$

Al_2O_3 is amphoteric, SiO_2 is slightly acidic whereas P_2O_3 and SO_2 are the anhydrides of acids H_3PO_3 and H_2SO_3.

Valency and Oxidation state

1. (a) Examples of neutral oxides are CO, H_2O, N_2O. These oxides are neutral towards litmus paper.

5. (a) Basic or metallic character of pentaoxides in VA group increases down the group. Hence acidity decreases.

7. (a) Na_2O, MgO, Al_2O_3, CuO. More the metallic character higher the e^- donating tendency. Therefore lower the I.E. more the basic nature of oxide.

8. (d) As it can donate e^- easily due to low comparative attraction by the nucleus to the valence e^-.

9. (b) Because of the non-metallic character increases.

10. (d) Oxidizing power increases in a group.

14. (c) HF is least acidic due to the small size of fluorine.

16. (d) $Co - [Ar]3d^7 4s^2$, it has 3 unpaired e^- so it is a paramagnetic.

17. (a) Transition elements due to presence of vacant *d*-orbitals.

18. (c) Its valency is 2. So it will form *MO* type compound.

19. (a) Oxides of alkali metals are most basic.

21. (a) Fluorine is the most easily reduced in halogens.

22. (b) Across the period non-metallic character increases. Hence basic nature of oxide decreases.

23. (b) Fluorine is more reactive than chlorine, bromine and iodine.

24. (b) Both are coinage metals

$3d^{10} 4s^1 - Cu$; $4d^{10} 5s^1 - Ag$

26. (c) *Li*, *Na*, *K*, contains only one e^- in outer most orbit.

27. (b) Valency is according to valence shall configuration which here is $1s^2, 2s^2, 2p^3$, *ie.* 5

28. (c) *Fe* belongs to first transition series.

29. (d) Reactivity of alkaline earth metals increases down the group.

31. (d) Tendency to gain e^- and oxidising power are related. Among halogens F is the directly most powerful oxidising agent.

32. (d) Electronic configuration of outermost shell of group-17 or halogens are $ns^2\ np^5$.

33. (b) On passing from left to right in a period acidic character of the normal oxides of the element goes on increasing with increase in electronegativity.

36. (b) Gold is found in native state.

37. (d) The elements which having same number of electrons in the valence shell are placed in the same group of periodic table.

38. (c) Alkali metals have the configuration $(n-1)s^2 p^6, ns^1$

41. (a) As going down the group size increases, an liberation of H^+ ion becomes easy. So the order of acidity is : $HI > HBr > HCl > HF$

44. (d) Valence shell configuration for IIA group elements is : ns^2

45. (b) A_2B_3

$A \xrightarrow{-3e^-} A^{+3}$; $B \xrightarrow{+2e^-} B^{-2}$

47. (b) Lower the value of I.P. of an element, the greater will be the basic character of the element.

48. (c) *N*, *O* and *F* have strong tendency to attract the shared pair of electrons i.e. by gaining electrons to form anions.

49. (c) B_2O_3, Al_2O_3 are amphoteric oxides.

50. (d) He has the atomic number 2 so it does not have octet.

51. (c) Beryllium has the valency of +2 while aluminum exhibits its valency as +3

Critical Thinking Questions

1. (d) As we go down the group inertness of ns^2 pair increase hence tendency to exhibit +2 oxidation state increases and that of +4 oxidation state decreases.

2. (d) Both $Be(OH)_2$ and $Zn(OH)_2$ are amphoteric in nature.

3. (d) CaO is basic; CO_2 is acidic; SiO_2 is weakly acidic. SnO_2 is amphoteric.

4. (b) In $BeCl_2$ has the highest melting point due to ionic bond.

5. (a) Correct order of electron affinity is *Se* < *S* < *O*. In a group electron affinity decreases with increase in atomic number.

6. (b) The correct increasing order of I.E. is,

$Cu < Ag < Au$.

7. (b) $B < C < N < O$; When we move from *B* to *O* in a periodic table the first ionisation enthalpy increase due to the attraction of nucleous towords the outer most of electron.

8. (c) Both Fe and Be are metal but Be has stable configuration so it is difficult to release e^- from it. So it has less metallic character than Fe.

9. (a) The basic nature of oxide decreases across the period as metallic character decreases. Therefore acidic nature of oxide increases.

10. (b) Halogens have very high electron affinities. It may be noted that the electron affinity of fluorine is unexpectedly low $(< Cl)$. This may perhaps be due to the small size of the F atom. The values of electron gain enthalpies for *Cl,F, S* and *O* are respectively 349, 333, 200 and 142 *kJ/mole* hence correct order is *Cl>F>S>O*.

11. (a) Increasing order of electronegativity is *Bi<P<S<Cl*.

12. (a) In a group, the ionisation potential decreases from top to bottom. In the alkali group, the ionisation potential decreases from *Li* to *Cs*.

	Li	*Na*	*K*	*Rb*	*Cs*
eV	5.3	5.1	4.3	4.2	3.9

13. (c) Because for removing second electron, it has to be taken out from stable configuration that needs a large amount of energy.

14. (a) First I.E. of $N >$ First I.E. of *O*.

15. (d) All the nobal gases occupy the peaks of I.E. curve.

16. (b) Correct order of ionic size is *Sn* > *Ce* > *Lu* > *Yb*.

17. (d) $1s^2 2s^2 p^6 3s^2$ - In III transition e^- is to be removed from stable configuration.

18. (c) $1s^2 2s^2 2p^6 3s^1$. It belongs to IA group which has least ionization potential and it decreases down the group

19. (d) Electronegativity increases on going from left to right in a period. Thus electronegativity of $F > O > N \simeq Cl$.

20. (d) Nuclear charge per electron is greater in P^{5+}. Therefore, its size is smaller.

21. (b) The electron affinities of some of the elements of second period (i.e., N, O, F etc.) are however, lower than the corresponding elements (i.e., *P, S, Cl*, etc.) of the third period. This is due to the reason that the elements of second period have the smallest atomic size amongst the elements in their respective groups. As a result, there are considerable electron-electron repulsion within the atom itself and hence the additional electron is not accepted with the same ease as is the case with the remaining elements in the same group.

22. (c) Element belongs to *d*-block is unnilhexium $(Unh)_{106}$.

Assertion and Reason

1. (d) Positive ions will be smaller than parent atoms.

3. (c) Calcium has a higher nuclear charge than sodium.

4. (c) 2*s* orbital has lower energy than 2*p*.

5. (c) *Cl* is more electronegative than *Li*. Although the difference is not much. Therefore the electron pair moves equally to both an thus forming a covalent compound.

6. (c) The lower value of electron affinity of *F* is due to electron-electron repulsion in 2-*p* orbitals of *F*-atom is stronger.

7. (d) All noble gases have stable configuration. Therefore, they can not take any electron means that they have no affinity for electrons. High electron affinity shows that electron is strongly bonded to the atom. Here both assertion and reason are false.

8. (a) The first ionization energy of *Be* is greater than Boron because it is difficult to remove electron from *Be* in comparison to boron. It is also true that the 2P orbitals have lower energy than 2 s-orbitals. Both assertion and reason are true and reason is correct explanation.

9. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

10. (b) Both assertion and reason are true but reason is not the correct explanation of assertion.

Actinoids are more complicated due to the opssibility of large number oxidation states.

11. (d) Both assertion and reason are false.

Ionization enthalpies are always positive. Energy is always absorbed when electrons are removed from an atom.

12. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

The phenomenon is which the Penultimate Shell $(n-1)$ electrons act as screen or shield in between nucleus and valence shell electrons thereby reducing the effective nuclear charge is known as shielding effect.

13. (c) Assertion is true but reason is false.

Ionisation potential decreases with increase in atomic size and also for a given a shell. I.E. is in given order.

$s > p > d > f$

14. (e) Assertion is false but reason is true.

More is the electron affinity, greater is the Oxidising character.

15. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

I.E. of N is more than that of $_8O$ as well as $_6C$.

16. (c) Assertion is true but reason is false.

N is half-filled $(1s^2 2s^2 2p^3)$ and therefore more stable and hence energy required to lose electron is greater.

17. (e) Assertion is false but reason is true.

$NO^- = 7 + 8 + 1 = 16\, e^-$ whereas

$CN^- = 6 + 7 + 1 = 14\, e^-$. So both are not isoelectronic.

18. (e) Assertion is false but reason is true.

Outermost electronic configuration of most electropositive elements is ns^1

19. (b) First ionization energy for nitrogen is lower than oxygen due to decrease nuclear charge in nitrogen comparison than oxygen.

Chemical Periodicity

SET Self Evaluation Test -15

1. If the difference in electronegativities of two elements is very large, then

(a) The bond is 50% ionic

(b) The bond is 100% covalent

(c) The bond is more covalent than ionic

(d) The bond is more ionic than covalent

2. Which of the following elements will have the lowest electron affinity

(a) Nitrogen (b) Flourine

(c) Chlorine (d) Phosphorus

3. The correct order of second ionization potential of carbon, nitrogen, oxygen and fluorine is

[IIT-JEE 1981; CBSE PMT 1991; MADT Bihar 1995; MP PMT 2003]

(a) $C > N > O > F$ (b) $O > N > F > C$

(c) $O > F > N > C$ (d) $F > O > N > C$

4. Which of the following species has the highest ionisation potential **[EAMCET 1998]**

(a) Li^+ (b) Mg^+

(c) Al^+ (d) Ne

5. Which of the following elements are analogous to the lanthanides **[AIIMS 1998]**

(a) Actinides (b) Borides

(c) Carbides (d) Hydrides

6. Which of the order for ionisation energy is correct

[CPMT 1999; CBSE PMT 2001]

(a) $Be > B > C > N > O$ (b) $B < Be < C < O < N$

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

7. Modern periodic table is based on the atomic number of the elements. The experiment which proved the significance of the atomic number was **[CBSE PMT 1989]**

(a) Millikan's oil drop experiment

(b) Moseley's work on X-ray spectra

(c) Bragg's work on X-ray diffraction

(d) Discovery of X-rays by Rontgen

8. Which one of the elements is most metallic

[MP PMT 2002]

(a) P (b) As

(c) Sb (d) Bi

9. For a *p* - block element, its 3*d*, 3*s*, 3*p* and 4*s* orbitals are completely filled and the differentiating electron goes to the 4*p* orbital. The element should have its atomic number in the range

(a) 13 - 18 (b) 21 - 26

(c) 31 - 36 (d) 49 - 54

10. The most common lanthanide is **[AFMC 1995]**

(a) Lanthanum (b) Cerium

(c) Samarium (d) Plutonium

11. In a period, elements are arranged in strict sequence of

[CPMT 1989]

(a) Decreasing charges in the nucleus

(b) Increasing charges in the nucleus

(c) Constant charges in the nucleus

(d) Equal charges in the nucleus

12. Some of the polar crystal when heated produce electric current. This phenomena is termed as **[AMU 2001]**

(a) Ferroelectric effect (b) Phyroelectric effect

(c) Antiferroelectric effect (d) Piezoelectric effect

13. Which of the following pairs has elements containing same number of electrons in the outermost orbit

[CPMT 1985]

(a) $N - O$ (b) $Na - Cl$

(c) $Ca - Cl$ (d) $Cl - Br$

14. Coinage metals are present in **[DCE 2000]**

(a) *s*-block (b) *d*-block

(c) *p*-block (d) *f*-block

15. In which of the following metal carbonate which metal carbonate is decomposed on heating **[UPSEAT 1999]**

(a) $MgCO_3$ (b) Na_2CO_3

(c) K_2CO_3 (d) Pb_2CO_3

16. Which one of the following is the correct decreasing order of boiling point **[AMU 2000]**

(a) $H_2O > H_2S > H_2Se > H_2Te$

(b) $H_2Te > H_2Se > H_2S > H_2O$

(c) $H_2O > H_2Te > H_2Se > H_2S$

(d) $H_2Te > H_2O > H_2Se > H_2O$

Answers and Solutions

(SET -15)

1. (d) If the difference in electronegativities of two elements is very high then the bond is more ionic than covalent.

2. (d) Phosphorus have the lowest electron affinity due to half filled p orbital, but in nitrogen electron affinity is greater than phosphorus because of large nuclear attraction in comparison with phosphorus.

3. (c) The ionization potential increases across the period but the second ionization potential of oxygen is highest among them because after the removal of $1e^-$ the $2e^-$ is to be removed from half filled orbital which is difficult.

4. (d) As, now the e^- is to be removed from stable configuration. Li^+ has the highest ionisation potential due to its stability.

5. (a) Actinides are homologous of Lanthanides.

6. (b) Ionisation energy increases across the period but due to stable half filled configuration of VA group, its I.E. is more than VI-A group.

7. (b) Moseley's work on X-ray spectra was proved the significance of the atomic number.

8. (d) The metallic property of an element increases from top to bottom in group.

$P < As < Sb < Bi$

9. (c) 31– 36 $\Rightarrow$ Ga to Kr.

10. (b) The most common lanthanide is cerium.

11. (b) Increasing charges in the nucleus as atomic number increases across a period.

12. (d) This phenomena is called piezoelectric effect.

13. (d) $Cl - Br$. Both belong to VII-A group having $7e^-$ in valence shell.

14. (b) Copper, Silver and Gold are coinage metals

15. (a) $MgCO_3 \rightarrow MgO + CO_2$

16. (c) Correct decreasing order of boiling point is,

$H_2O > H_2Te > H_2Se > H_2S$.

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