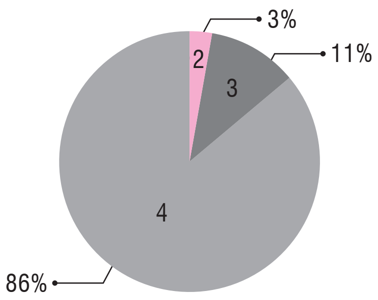
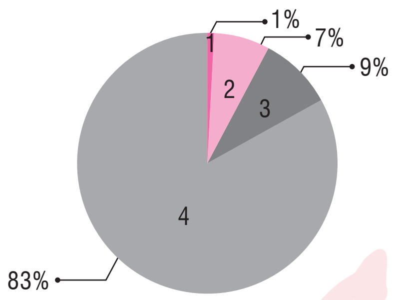
3 Classification of Elements and Periodicity in Properties

Topicwise Trend Analysis of Past 10 Years' Questions



NEET Analysis



JEE Analysis

Topics

1. Genesis of Periodic Classification
2. Modern Periodic Law and Long Form of Periodic Table
3. Electronic Configuration of Elements and Periodic Table
4. Periodic Trends in Properties of Elements

- It is very difficult to study individually the chemistry of all the elements and their compounds, hence to simplify and systematise the study of chemistry of the elements and their compounds, elements are classified into groups and periods.

Genesis of Periodic Classification

- **Dobereiner's Triads (1829):** Dobereiner classified elements into a group of three called *Triads*. In the triads of elements, the atomic weight of the middle element was the arithmetic mean of the atomic weights of the other two.

Triads			Mean of atomic masses of 1st and 3rd elements
Li 7	Na 23	K 39	$\frac{7+39}{2}=23$
Ca 40	Sr 88	Ba 137	$\frac{40+137}{2}=88.5$
Cl 35.5	Br 80	I 127	$\frac{35.5+127}{2}=81.25$

Limitation : He could not classify all the elements, known at that time, into Triads.

- **Newlands' Law of Octaves (1865) :** Newland arranged elements in increasing order of atomic masses where the properties of every eighth element are similar to the first one, like the musical scale.

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca					

This classification worked quite well for the lighter elements but it failed in case of heavier elements and therefore, discarded.

Limitations : He could not classify all the elements known at that time.

- **Lothar Meyer's Atomic Volume Curve :** Lothar Meyer, plotted a graph between atomic volumes (atomic mass/density) of the elements against their atomic masses and observed that the elements with similar properties occupied similar positions on the curve.
- **Mendeleev's Periodic Law :** Mendeleev's periodic law stated that, "the physical and chemical properties of the elements are periodic function of their atomic masses". He framed a periodic table with eight vertical columns called *groups* and six horizontal rows called *periods*.

Importance of Mendeleev's Periodic Table

(i) It made the study of elements systematic.
(ii) It helped in prediction of new elements and their properties.
(iii) It helped in correction of doubtful atomic masses.

Defects in Mendeleev's Periodic Table

(i) Position of hydrogen is controversial.
(ii) Anomalous position of some elements. In some cases the elements with higher atomic masses precessed the elements with lower atomic masses *e.g.*, Ar (39.9) precessed K (39.1); Co (58.9) precessed Ni (58.7).
(iii) Isotopes (atoms of same elements with different atomic masses) have been placed together.
(iv) There is no correlation of elements in sub-groups. The properties of elements in sub-group 'A' differ from those in sub-group 'B' of the same group.
(v) In some cases elements with similar properties have been placed in different groups *e.g.*, copper and mercury.
(vi) Lanthanides and actinides have not been provided a proper place.
(vii) No proper explanation was afforded for cause of periodicity.

In-text Examples

1 | X, Y and Z are three members of a Dobereiner's Triad. If the atomic mass of X is 7 and that of Z is 39, what is the atomic mass of Y?

Soln.: Atomic mass of Y

$$= \frac{\text{Atomic mass of } X + \text{Atomic mass of } Z}{2} = \frac{7+39}{2} = 23$$

2 | What was the basis of classifying the elements by Lothar Meyer?

Soln.: Lothar Meyer used the physical properties such as melting point, boiling point and atomic weight to construct a plot and obtained a periodically repeated pattern.

- **Modern Periodic Law and the present form of the periodic table :** According to Moseley "physical and chemical properties of elements are the periodic functions of their atomic numbers". If the elements are arranged in order of their increasing atomic number, after a regular interval, element with similar properties are repeated.
- Long form of periodic table consists of horizontal rows called as '*periods*' and vertical columns called as '*groups*'.
- The long or extended form of periodic table consists of seven periods and eighteen groups.
- The elements in a period have same number of energy shells, *i.e.*, same principal quantum number (n). These are numbered 1 to 7.

Period	Principal energy level being filled (n)	Orbitals being filled	No. of elements in the period
1st period	1	$1s$	2 elements
2nd period	2	$2s, 2p$	8 elements
3rd period	3	$3s, 3p$	8 elements
4th period	4	$4s, 3d, 4p$	18 elements
5th period	5	$5s, 4d, 5p$	18 elements
6th period	6	$6s, 4f, 5d, 6p$	32 elements
7th period	7	$7s, 5f, 6d, 7p$	32 elements
Total			118 elements

- In a group, the elements have similar valence shell electronic configuration, therefore, exhibit similar chemical properties.

Group 1 elements	Alkali metals
Group 2 elements	Alkaline earth metals
Group 11 elements	Coinage metals
Group 15 elements	Pnictogen
Group 16 elements	Chalcogens
Group 17 elements	Halogens
Group 18 elements	Noble gases or Aerogens

Classification of Elements and Their Electronic Configuration

- The elements can be classified into four blocks namely s-, p-, d- and f- depending on the type of atomic orbitals that are being filled with the valence electrons.

	Last e^- enters	Electronic configuration	Group no.	Position in periodic table	Metals/Non-metals
s-block	ns-orbital $n = 1$ to 7	ns^{1-2}	1 and 2	Left side	Metals
p-block	np-orbital $n = 2$ to 7	ns^2np^{1-6}	13 to 18	Right side	Few are metals and metalloids but most of them are non-metals.
d-block	$(n-1)d$-orbital $n = 4$ to 7	$(n-1)d^{1-10}ns^{0-2}$	3 to 12	In the middle	Metals
f-block	$(n-2)f$-orbital $n = 6$ and 7	$(n-2)f^{0-14}$ $(n-1)d^{0-2}ns^2$	Related to group 3	Placed separately below the main periodic table.	Metals

Predicting the Period and Group of an Element

- Period of an element = Principal quantum number of the valence shell.
- Block of an element = Type of orbital which receives the last electrons.
- Group of an element :
  - For s-block elements : Group number = number of valence electrons.
  - For p-block elements : Group number = 10 + number of valence electrons.
  - For d-block elements : Group number = number of electrons in $(n-1)d$ subshell + number of electrons in valence (n^{th}) shell.

Periodic Table

Metals | Metalloids | Non-metals

The zigzag line separates the metals from the non-metals.

s-block elements (Representative elements): groups 1–2
d-block elements (Transition elements): groups 3–12
p-block elements (Representative elements): groups 13–18

Periods (No. of elements)	IA 1	IIA 2	IIIB 3	IVB 4	VB 5	VIB 6	VIIB 7	VIII 8	9	10	IB 11	IIB 12	IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	Noble gases 0 18
1 (2) ($_1$H-$_2$He)	1 H Smallest atomic size																	2 He Highest ionisation potential
2 (8) ($_3$Li-$_{10}$Ne)	3 Li Lightest solid metal	4 Be											5 B Lightest solid	6 C Maximum catenation	7 N	8 O	9 F Highest electro negative	10 Ne
3 (8) ($_{11}$Na-$_{18}$Ar)	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4 (18) ($_{19}$K-$_{36}$Kr)	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br Liq. non-metal	36 Kr
5 (18) ($_{37}$Rb-$_{54}$Xe)	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te Most stable element	53 I Sublime on heating	54 Xe
6 (32) ($_{55}$Cs-$_{86}$Rn)	55 Cs Lowest electro-negative; lightest liq. metal	56 Ba	57 La	72 Hf	73 Ta	74 W Highest m.pt., b.pt. metal	75 Re	76 Os Heaviest solid metal	77 Ir	78 Pt (Noble Metals)	79 Au (Noble Metals)	80 Hg Lowest m.pt. metal	81 Tl	82 Pb	83 Bi	84 Po	85 At Heaviest solid	86 Rn Heaviest gaseous non-metal
7 (32) ($_{87}$Fr-$_{118}$Og)	87 Fr Liquid element of radio-active nature	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

f-block elements (Inner-transition elements)

Lanthanide series	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinide series	90 Th	91 Pa	92 U	93 Np	94 Pu Most poisonous element	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

In-text Examples

3 | In what manner is the long form of periodic table better than Mendeleev's periodic table? *(NCERT Exemplar)*

Soln.: Long form of periodic table is based on the modern periodic law which states that the physical and chemical properties of elements are periodic functions of their atomic numbers. This arrangement removed the defects of Mendeleev's periodic table and groups A and B are replaced by numbers 1-18.

4 | All transition elements are *d*-block elements, but all *d*-block elements are not transition elements. Explain. *(NCERT Exemplar)*

Soln.: According to the definition of transition metals the elements should have incomplete penultimate $(n - 1)$ shell. But few *d*-block elements have completely filled penultimate shell, they are not considered as transition metals.
e.g., Zn$(3d^{10}\,4s^2)$, Cd $(4d^{10}, 5s^2)$ and Hg $(5d^{10}\,6s^2)$.

5 | An element X with $Z = 112$ has been recently discovered. What is the electronic configuration of the element? To which group and period will it belong?

Soln.: (a) The electronic configuration of element X is $[Rn]^{86}\,5f^{14}\,6d^{10}\,7s^2$

(b) It belongs to *d*-block as last electron enters in *d*-subshell.

(c) As number of electrons in $(n - 1)d$ subshell and valence shell is equal to twelve *i.e.*, 10 + 2. So it belongs to group 12.

(d) It belongs to period 7 of the periodic table as principal quantum number of valence shell is 7 (*i.e.*, $7s^2$).

Practice Time 1

1. The electronic configuration of four elements are given below. Which element does not belong to the same family?
(a) [Xe] $4f^{14}5d^{10}6s^2$ (b) [Kr]$4d^{10}5s^2$
(c) [Ne] $3s^23p^5$ (d) [Ar] $3d^{10}4s^2$

2. Only one element of __________ forms hydride.
(a) group 6 (b) group 7 (c) group 8 (d) group 9

3. Mark out the incorrect match : (electronic configuration – nature of element)
(a) $3s^23p^6$: noble gas (b) $5s^24d^{10}5p^5$: non-metal
(c) $4s^23d^7$: metal (d) $3s^23p^2$: non-metal

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

5. Lothar Meyer attempt was based on plotting atomic mass *vs*
(a) atomic size (b) atomic volume
(c) density (d) all of these.

6. Eka-aluminium and Eka-silicon are known as
(a) gallium and germanium (b) aluminium and silicon
(c) iron and sulphur (d) boron and silicon.

7. Newlands' law of octave applies to which of the following set of elements?
(a) Be, Mg, Ca (b) As, K, Ca
(c) B, N, C (d) None of these

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

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

10. The element of the 17^{th} group of the periodic table is likely to be
(a) moderately non-metallic (b) moderately metallic
(c) strongly non-metallic (d) strongly metallic.

11. Predict the period, group number and block to which an element with $Z = 54$ belongs.
(a) 5^{th} period, group 18, p-block
(b) 5^{th} period, group 3, d-block
(c) 5^{th} period, group 14, p-block
(d) 6^{th} period, group 8, d-block

12. Of the following pairs, the one containing examples of metalloid elements in the periodic table is
(a) Na and K (b) F and Cl
(c) Cu and Hg (d) Si and Ge.

13. Which of the following sets of elements follows Dobereiner's Triads?
(a) Mg, Ca, Sr (b) Li, Na, K
(c) F, Cl, Br (d) Mn, Tc, Re

14. The elements of group 1, 2, 13, 14, 15, 16, 17, 18 are collectively called
(a) noble elements (b) typical elements
(c) transition elements (d) representative elements.

15. Atomic mass of which of the following was changed by Mendeleev?
(a) B (b) Ba (c) Be (d) Mg

Periodic Trends in Properties of Elements

- The recurrence of similar properties of the elements after certain definite intervals when the elements are arranged in order of increasing atomic numbers in the periodic table is called *periodicity.*
- The cause of periodicity is the repetition of similar electronic configurations of the atom in the valence shell after certain definite intervals. These definite intervals are 2, 8, 8, 18, 18 and 32.
- Periodicity is observed in a number of properties which are directly or indirectly linked with electronic configuration.

Atomic Radius

It is usually defined as the distance between the nucleus and outermost shell where electrons are present. Three types of radii are commonly used, *i.e.*, (a) covalent radii (b) crystal radii (c) van der Waals' radii.

- Atomic radius increases down the group as new valence shells are added.
- In general, atomic radius decreases with increase in atomic number as we move from left to right in a period (except for noble gases).
- **Covalent radius** : It is defined as half of the distance between the two nuclei of two like atoms bonded together by a single covalent bond.
  – For a homonuclear diatomic molecule (A_2) **:** $r_A = \dfrac{d_{A-A}}{2}$
  [$\because$ d_{A-A} = bond length]
  – For a heteronuclear diatomic molecule (AB) **:**
  $d_{A-B} = r_A + r_B$
  – If the covalent bond is formed between two elements of different electronegativity then the following relation is used :
  $d_{A-B} = r_A + r_B - 0.09\,(\chi_A - \chi_B)$
  where, χ_A and χ_B are electronegativity of A and B respectively.
- **Crystal radius :** It is defined as one half of the distance between the nuclei of two adjacent metal atoms in the metallic closed packed crystal lattice in which metal exhibits a coordination number of 12.

- **van der Waals' radius :** It is half of the distance between the nuclei of two non-bonded neighbouring atoms of two adjacent molecules in the solid state.

  $r_{\text{covalent}} < r_{\text{crystal}} < r_{\text{van der Waals'}}$

Ionic Radius

- It is defined as the distance between the nucleus and outermost shell of an ion or it is the distance between the nucleus and the point where the nucleus exerts its influence on the electron cloud.
- The ionic radii of elements exhibit the same trend as the atomic radii.
- The removal of an electron from an atom results in the formation of a cation.
- A cation is smaller than its parent atom because it has fewer electrons while its nuclear charge remains the same.
- More the electrons are removed, smaller the ion becomes. Thus, $Mg > Mg^+ > Mg^{2+}$
- Negative ion is formed by gain of one or more electrons in the neutral atom and thus, number of electrons increases but magnitude of nuclear charge remains the same.
- Due to decrease in nuclear charge per electron, there is an expansion of outer shell. Thus, size of anion is increased. $O^{2-} > O^- > O; I^- > I > I^+$
- These can be explained on the basis of Z/e ratio *i.e.*, $\left(\frac{\text{Nuclear charge}}{\text{No. of electrons}}\right)$. When Z/e ratio increases, the size decreases and when Z/e ratio decreases, the size increases.
- **Isoelectronic species :** The atoms and ions which contain the same number of electrons are known as *isoelectronic species e.g.*, O^{2-}, F^-, Na^+ and Mg^{2+}, etc.
- In case of isoelectronic ions, the size decreases with increase in the nuclear charge.

Ionisation Enthalpy

- It is the energy required to remove an electron from an isolated gaseous atom in its ground state.

  $M_{(g)} + I.E. \longrightarrow M^+_{(g)} + e^-$
- If gaseous atom is to lose more than one electron, they can be removed only one after another *i.e.*, in succession and not simultaneously. This is known as *successive ionisation energy* or *potential* :

  $I.E_3 > I.E_2 > I.E_1$
- **The screening effect or shielding effect :** In a multielectron atom, valence shell electrons experience less attractive force due to repulsion between inner shell electrons and valence electrons. This is called *screening effect or shielding effect*.
- **Factors on which *I.E.* depends :**
  (i) **Atomic size :** The ionisation energy decreases with increase in atomic size.
  (ii) **Nuclear charge :** The ionisation energy increases with increase in magnitude of nuclear charge.
  (iii) **Screening effect :** The reduction in force of attraction by the electrons of shells present in between the nucleus and valence electrons is called *screening effect* or *shielding effect.* Greater the number of intervening electrons between valence electron and nucleus, the greater will be shielding or screening effect and hence lower is the *I.E.*
  (iv) **Penetration effect :** Penetration power of various subshells of a particular energy levels is in the order $s > p > d > f$. Therefore, for the same shell it is easier to remove an electron from p-subshell than from s-subshell. Thus greater the penetration power higher is the *I.E.*
  (v) **Electronic configuration :** The elements having relatively stable electronic configuration have relatively higher values of ionisation energy. *e.g.*
    (a) The noble gases having stable configuration (ns^2np^6) have highest ionisation energies within their respective periods.
    (b) The elements like N($1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$) and P($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^1$) having configurations in which orbitals belonging to the same subshell are exactly half-filled are quite stable and so they have relatively high ionisation energies.
    (c) The elements like Be($1s^22s^2$) and Mg($1s^22s^22p^63s^2$) have all electrons paired. Such configurations being stable also result in higher value of *I.E.*
- Ionisation energy decreases down the group due to increase in size of the atom and screening effect of intervening electrons.
- Generally, the first ionisation enthalpy increases along a period from left to right as the charge on nucleus increases while valence shell remains the same leading to increase in effective nuclear charge.

Electron Gain Enthalpy

- It may be defined as the amount of energy released when an electron is added to an isolated gaseous atom of the element.

  $A_{(g)} + e^- \longrightarrow A^-_{(g)};\ \Delta_{eg}H$
- **Factors affecting electron gain enthalpy :**
  – **Size of the atom :** $\Delta_{eg}H \propto \frac{1}{\text{Size of atom}}$
  – **Nuclear charge :** $\Delta_{eg}H \propto$ Effective nuclear charge
  – **Screening effect :** $\Delta_{eg}H \propto \frac{1}{\text{Screening effect}}$
- Electron gain enthalpy becomes less negative as we move down the group.
- Electron gain enthalpy becomes more and more negative from left to right in a period (except for noble gases).

Electronegativity

- It is the tendency of an atom to attract the shared pair of electrons towards itself in a covalent bond. Greater the electronegativity of an atom, greater will be its tendency to attract the shared pair of electrons towards itself.
- Electronegativity decreases on moving down the group and increases along a period.
- **Factors affecting electronegativity :**
  (i) **Effective nuclear charge :** Greater the effective nuclear charge greater is electronegativity.
  (ii) **Atomic radius :** Smaller the atomic radius greater is the electronegativity.

Valence or Oxidation State

The valence of representative elements is usually equal to the number of electrons in the outermost orbitals or equal to eight minus the number of outermost electrons.

- Oxidation state is frequently used for valence.
- The transition elements and actinoids exhibit variable valencies.
- Valency of the elements increases from 1 to 4 and then decreases from 4 to 0 in a period.
- When we move down the group, the number of valence electrons remains the same therefore, all the elements in a group exhibit the same valency.

Anomalous Properties of Second Period Elements

- The first element of each of the group 1 (lithium) and 2 (beryllium) and group 13-17 (boron to fluorine) differs in many respects from the other members of their respective groups.
- Behaviour of Li and Be is more similar with the second element of the following group *i.e.*, Mg and Al respectively. This sort of similarity is commonly referred to as *diagonal relationship* in the periodic properties.
- The anomalous behaviour is attributed to their small size, large charge/radius ratio and high electronegativity of the elements.

Periodic Trends and Chemical Reactivity

- Chemical reactivity is high at the two extremes of the periodic table and lowest in the centre.
- The high chemical reactivity of alkali metals on the extreme left is due to their ability to lose an electron while that of halogens on the right side is due to their high tendency to gain an electron.
- The metallic character of an element, which is highest at the extremely left decreases and the non-metallic character increases while moving from left to right across the period.
- The metallic character increases down the group and non-metallic character decreases down the group.

Nature of Oxides

- In a period, the nature of the oxides varies from basic to acidic.

Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Strongly basic	Basic	Amphoteric	Weakly acidic	Acidic	Strongly acidic	Strongly acidic

- When an element forms a number of oxides, the acidic nature increases as the percentage of oxygen increases.

N_2O	Neutral	
NO	Neutral	
N_2O_3		Acidic nature increases
NO_2		
N_2O_5		

P_2O_3, P_2O_4, P_2O_5 — Acidic nature increases

Acidic nature increases:

MnO	Basic
Mn_2O_3	Basic
MnO_2	Neutral
MnO_3	Acidic
Mn_2O_6	Acidic

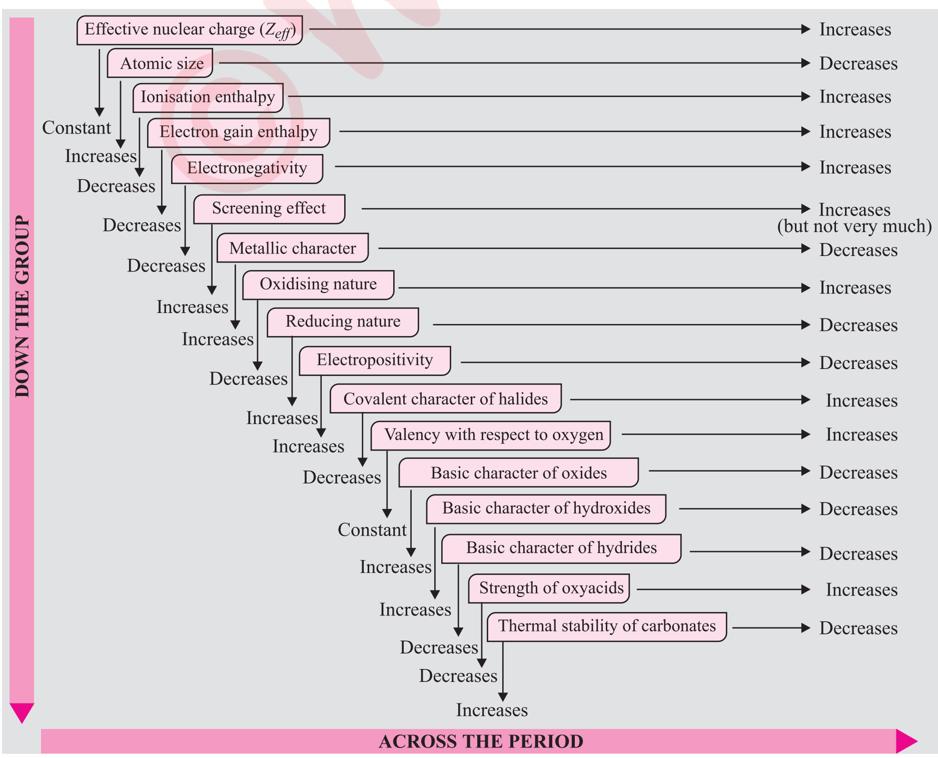
CO, N_2O, NO and H_2O are neutral oxides. The oxides CO_2, N_2O_5, P_2O_3, P_2O_5, SO_2, SO_3, Cl_2O_7, etc. are called *acid anhydrides* as these combine with water to form oxy-acids.

$CO_2 \longrightarrow H_2CO_3$: $P_2O_3 \longrightarrow H_3PO_3$
$SO_3 \longrightarrow H_2SO_4$: $N_2O_5 \longrightarrow HNO_3$
$P_2O_5 \longrightarrow H_3PO_4$: $Cl_2O_7 \longrightarrow HClO_4$
$N_2O_3 \longrightarrow HNO_2$: $SO_2 \longrightarrow H_2SO_3$

Nature of Hydrides

- The nature of the hydrides changes from basic to acidic in a period from left to right.

NH_3	H_2O	HF
Weak base	Neutral	Weak acid
PH_3	H_2S	HCl
Very weak base	Weak acid	Strong acid



In-text Examples

6 The radius of Na^+ cation is less than that of Na atom. Give reason. *[NCERT Exemplar]*

Soln.: The radius of Na^+ is less than Na atom because Na^+ is formed by losing one energy shell.

Na – $1s^2\,2s^2\,2p^6\,3s^1$; Na^+ – $1s^2\,2s^2\,2p^6$

7 Mg^{2+} is smaller than O^{2-} in size, though both have same electronic configuration. Explain.

Soln.: Mg^{2+} and O^{2-} both are isoelectronic *i.e.,* have same number of electrons. But Mg^{2+} having 12 protons in its nucleus exerts higher effective nuclear charge than O^{2-} having 8 protons and thus valence shell as well as inner shells electrons are more strongly attracted by the nucleus in Mg^{2+} resulting smaller size than O^{2-}.

8 Account for the large decrease in electron affinity between Li and Be despite the increase in nuclear charge.

Soln.: The electronic configurations of Li and Be are $[He]2s^1$ and $[He]2s^2$, respectively. The additional electron enters the 2*s* orbital of Li but the 2*p* orbital of Be and hence is much less tightly bound. In fact, the nuclear charge is so well shielded in Be that electron gain is endothermic.

Practice Time 2

16. Which of the following does not affect the periodicity of the elements?
(a) Bonding behaviour (b) Electronegativity
(c) Ionization energy (d) Neutron/proton ratio

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

18. The correct decreasing order of first ionisation enthalpies of five elements in the second period is
(a) Be > B > C > N > F (b) N > F > C > B > Be
(c) F > N > C > Be > B (d) N > F > B > C > Be.

19. Which of the following is associated with the biggest jump between the second and third ionization energies?
(a) $1s^2\,2s^2\,2p^3$ (b) $1s^2\,2s^2\,2p^6\,3s^1$
(c) $1s^2\,2s^2\,2p^6\,3s^2$ (d) $1s^2\,2s^2\,2p^1$

20. The oxidation state of nitrogen varies from
(a) –3 to +5 (b) 0 to +5
(c) –3 to 1 (d) +3 to +5

21. Which of the following have no unit?
(a) Electronegativity (b) Electron gain enthalpy
(c) Ionisation enthalpy (d) Atomic radius

22. Which of the following is most electronegative element?
(a) Li (b) Mg
(c) H (d) Na

23. Which of the following processes involves absorption of energy?
(a) $Cl + e^- \rightarrow Cl^-$ (b) $O^- + e^- \rightarrow O^{2-}$
(c) $O + e^- \rightarrow O^-$ (d) $S + e^- \rightarrow S^-$

24. The screening effect of *d*-electrons is
(a) equal to the *p*-electrons
(b) much more than *p*-electron
(c) same as *f*-electrons
(d) less than *p*-electrons.

25. The first (IE_1) and second (IE_2) ionization energies (kJ/mol) of a few elements designated by Roman numerals are given below. Which of these would be an alkali metal?

		IE_1	IE_2
(a)	I	2372	5251
(b)	II	520	7300
(c)	III	900	1760
(d)	IV	1680	3380

26. Electrons of which subshell do not participate in bonding due to inert pair effect?
(a) 6*s* (b) 6*p*
(c) 5*d* (d) 4*f*

27. Point out the correct statement, in a given period of the periodic table, the *s*-block elements has, in general, a higher value of
(a) electronegativity (b) atomic radius
(c) ionization energy (d) electron affinity.

28. What is the order of successive ionisation enthalpies?
(a) $IE_{III} > IE_{II} > IE_{I}$ (b) $IE_{I} > IE_{II} > IE_{III}$
(c) $IE_{II} > IE_{I} > IE_{III}$ (d) $IE_{III} > IE_{I} > IE_{II}$

29. Which of the following is the most polar bond?
(a) N–H (b) Cl–H
(c) O–H (d) Br–H

30. Which of the following remains unchanged on descending a group in the periodic table?
(a) Valence electrons (b) Atomic size
(c) Density (d) Metallic character

At a Glance

- **Factors affecting the value of ionisation potential**

	Properties	Effects
1.	Atomic Size	Larger the atomic size, smaller is the value of ionisation potential.
2.	Screening Effect	Higher the screening effect, lesser is the value of ionisation potential.
3.	Nuclear charge	Ionisation potential increases with the increase in nuclear charge.
4.	Penetration effect or Shape of orbital	Values of ionisation potential for *s, p, d* and *f* electrons are : $s > p > d > f$.

- **Factors affecting the value of electron affinity**

	Properties	Effects
1.	Nuclear charge	Electron affinity increases with the increase in nuclear charge.
2.	Atomic size	With the increase in atomic size, electron affinity decreases.
3.	Electronic configuration	Electron affinities are low or almost zero in cases of stable configurations *i.e.,* half filled or full-filled valence shell.

- **In short, periodic properties can be studied as follows :**

Properties	Along the period	Down the group
Ionisation potential	increases	decreases
Electron affinity	increases	decreases
Electro-negativity	increases	decreases
Atomic radii	decreases	increases
Ionic radii	iso-electronic ions decrease with increase in atomic number	increases
Atomic volume	decreases upto metals and then increases	increases
Melting point/ Boiling point	increases along the period for metals	decreases
Density	increases for metals	increases
Oxidant-Reductant nature	reducing nature decreases	reducing nature of metals increases, oxidising nature of non metals decreases
Metallic character	decreases	increases
Electro-positive character	decreases	increases
Oxide nature	basic character decreases	basic character increases
Hydride nature	basic character decreases or acidic character increases	basic character decreases
Valency	with respect to oxygen increases from 1-7 along the period, with respect to hydrogen increases from 1 to 4 and then decreases to 1.	remains the same

- Electronegativity increases with increase in +ve charge on cation.
  $M^{+2} < M^{+3} < M^{+4}$
- Electronegativity decreases with increase in –ve charge.
  $X^{+} > X > X^{-}$
- Nature of the bond between two atoms can be predicted from the electronegativity difference of the two atoms.
  (a) If the difference $\chi_A - \chi_B = 0$, *i.e.,* $\chi_A = \chi_B$, the bond is purely covalent. For example, H_2, Cl_2, O_2 and N_2 molecules are purely covalent and non-polar.
  (b) If the difference $\chi_A - \chi_B$ is small, *i.e.,* $\chi_A > \chi_B$, the bond is polar covalent.
  (c) If the difference $\chi_A - \chi_B$ is 1.9, the bond is 50% covalent and 50% ionic.
  (d) If the difference $\chi_A - \chi_B$ is very high, the bond is more ionic and less covalent. The molecule will be represented in such case as BA ($B^{+}A^{-}$).
- Elements having low values of electronegativity are metals while the elements having high values of electronegativity are non-metals.

Chemical Reactivity

- Chemical reactivity is lowest at the centre of periodic table and highest at left and right side of periodic table. The chemical reactivity of alkali metals at extreme left is due to their ability to lose an electron and form cation.

Extra Edge

- Atomic radius in the n^{th} orbit is given by

$$r_n = \frac{n^2 a_0}{Z_{eff}}$$

where, n is principal quantum number (*i.e.*, number of shell), a_0, the Bohr's radius of H-atom (= 0.529Å) and Z_{eff} the effective nuclear charge.

- Ionization potential (eV) $= \dfrac{\text{Ionization energy in Joules}}{\text{Charge of electron } (1.6 \times 10^{-19})}$

- Electronegativity (eV) $= \dfrac{IE + EA}{2}$

- $\chi_A - \chi_B = 0.208[E_{A-B} - \underbrace{(E_{A-A} \times E_{B-B})^{1/2}}_{\text{Geometric mean}}]^{1/2}$

- $\chi_A - \chi_B = 0.208\sqrt{E_{A-B} - \underbrace{1/2(E_{A-A} + E_{B-B})}_{\text{Arithmetic mean}}}$

- On Pauling scale, electronegativity of an atom $= \dfrac{IE + EA}{5.6}$

- Allred - Rochow electronegativity $= 0.744 + \dfrac{0.359 Z_{eff}}{r^2}$

- Electronegativity of an atom (in kcals per mole) $= \dfrac{IE + EA}{2 \times 64.5}$

- Slater's Rule for determination of Z_{eff} (effective nuclear charge) : $Z_{eff} = Z - \sigma^*$ (σ^* is Slater's screening constant)
σ^* = [0.35 × no. of other electrons in n^{th} shell + 0.85 × no. of electrons in $(n - 1)^{th}$ shell + 1.00 × total no. of electrons in the inner shells]

Atomic radii $\propto \dfrac{1}{Z_{eff}}$, $I.E. \propto Z_{eff}$

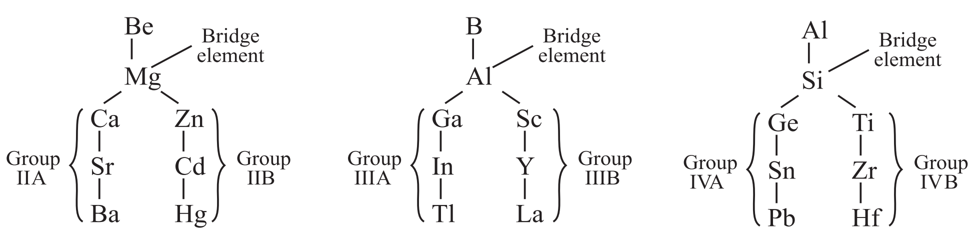
- Percentage ionic character $= 16(\chi_A - \chi_B) + 3.5(\chi_A - \chi_B)^2$
  - If $\chi_A - \chi_B = 1.7$, bond is 50% covalent and 50% ionic.
  - If $\chi_A - \chi_B > 1.7$, bond is predominately ionic.
  - If $\chi_A \approx \chi_B$, $A - B$ bond is purely covalent.

Slater's Rule for Estimating Effective Nuclear Charges, Z*

- In a multielectron atom, valence shell electrons experience less attractive force due to repulsion between inner shell electrons and valence electrons. This is called screening effect or shielding effect.
- Effective nuclear charge $Z^* = (Z - \sigma)$
where σ is screening constant (slater's rule)
- Write out the electronic configuration of the element in the following order and grouping :
(1*s*), (2*s*, 2*p*), (3*s*, 3*p*), (3*d*), (4*s*, 4*p*), (4*d*), (4*f*), (5*s*, 5*p*) etc.
- Electrons in any group higher in the sequence show that the electron under consideration contributes nothing to the shielding σ.
- Then for an electron in an *ns* or *np*-orbital
  - all other electrons in (*ns*, *np*) group contribute $\sigma = 0.35$ each.
  - all electrons in the $(n - 1)$ shell contribute $\sigma = 0.85$ each.
  - all electrons in the $(n - 2)$ or lower shell contribute $\sigma = 1.00$ each.
- For electron in an *nd* or *nf*-orbital, all electrons in the same group contribute $\sigma = 0.35$ each; those in group lying lower in the sequence than the (*nd*) or (*nf*) group contribute $\sigma = 1.00$ each.

Bridge Elements

In the Mendeleev's periodic table, the elements of third period such as Mg, Al and Si are called *bridge elements* since division of each main groups into subgroups *A* and *B* starts from these elements.



SMART REVISION MAP

CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

Chemists would be overwhelmed with isolated pieces of information if they did not have some way of relating the facts they know about the more than 100 known elements. The periodic table provides a means of organizing information so that relationships among elements can be clearly seen and understood. Scientist William Prout, John Dobereiner, John Dalton and John Newlands have contributed towards the development of Modern Periodic Table.

Earlier Attempts

Döbereiner's Law of Triads

- Elements in a triad had similar properties.
- The atomic weight of the middle element was very close to the arithmetic mean of the other two elements.

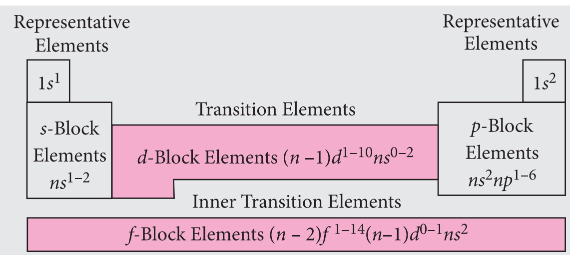
Newland's Law of Octaves

- When elements were arranged in increasing order of their atomic weights, properties of every eighth element were similar to those of the first one like the eighth note of a musical scale.

Mendeleev's Periodic Law

- The properties of the elements are periodic function of their atomic weights.
- Mendeleev's original periodic table contains 8 vertical columns called *groups* and 6 horizontal rows called *periods*.

Blocks



Modern Periodic Table

Modern Periodic Law

- The physical and chemical properties of the elements are a periodic function of their atomic numbers.

Long Form of Periodic Table

- Based on modern periodic law.
- Follows Bohr's scheme for the arrangement of various electrons around the nucleus.
- Contains 18 groups and 7 periods.

Periods

- **Periodic number :** highest principal quantum number (n) of the elements of the periodic table.
- Number of elements in each period is twice of the atomic orbitals available in the energy level that are being filled.
- Period 1 (n = 1) -2 elements
- Period 2 (n = 2) - 8 elements
- Period 3 (n = 3) - 8 elements
- Period 4 (n = 4) -18 elements
- Period 5 (n = 5) - 18 elements
- Period 6 (n = 6) - 32 elements
- Period 7 (n = 7) - 32 elements
  - Lanthanoids : 14 elements of period 6
  - Actinoids : 14 elements of period 7

  (Placed in the bottom of the periodic table separately.)

Arrangement of Elements

- **Metals :** > 78% of all known elements appear on the left hand side of the periodic table.
- **Non metals :** < 20, lie on the top right hand side of the periodic table.
- **Metalloids or semi-metals :** B, Si, Ge, As, Sb, Te, Po and At, run diagonally across the periodic table.

Groups

- Group number for
  - *s*-block : no. of valence electrons in *ns*-orbital.
  - *p*-block : 10 + no. of valence electrons in *np*-orbital.
  - *d*-block : no. of valence electrons in $(n-1)d$ and *ns*-orbitals.
- Group 1 - Alkali metals
- Group 2 - Alkaline earth metals
- Group 11 - Coinage metals
- Group 15 - Pnictogens
- Group 16 - Chalcogens
- Group 17 - Halogens
- Group 18 - Noble gases or Aerogens

Periodic Trends

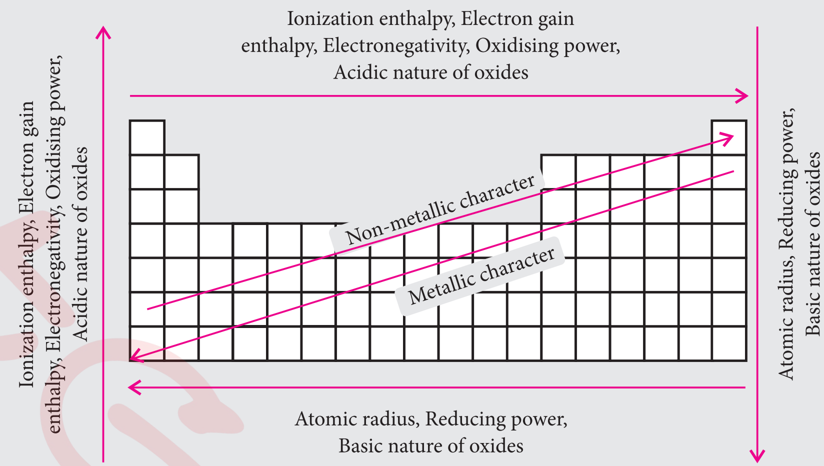
Periodic Properties

- Properties which show a regular gradation from left to right in a period and top to bottom in a group.

Exceptions

- $\Delta_{eg}H$ of Be, Mg, Ca, N and inert gases is positive.
- Ar and Kr have the same value of $\Delta_{eg}H$.
- $\Delta_{eg}H$ of F is less –ve than Cl.
- $\Delta_{eg}H$ of O is less –ve than S.
- $I.E._1$ of O < $I.E._1$ of N.

Properties



IN A NUTSHELL

- Mendeleev had predicted the properties of gallium (*eka-aluminium*) and germanium (*eka-silicon*) long before they were known.
- IUPAC names of elements with atomic numbers > 100 are derived directly from the atomic numbers using numerical roots for 0 and number from 1 – 9 and adding suffix *ium*.
- Valency first increases from 1 to 4 and then decreases to zero along a period and remains same within group.
- Some elements in second period show similarities with third period elements placed diagonally to each other.
- Boron (B) has the smallest atomic volume and highest tensile strength.
- Technetium (Tc) is the first synthetic element.
- Lanthanide which does not occur in nature is promethium (Pm).
- Francium (Fr) is liquid, radioactive element.
- Zn, Cd and Hg are volatile *d*-block elements.
- Mercury (Hg) is also called ‘liquid silver’.

EXERCISE

1 Topicwise Warm-up

Genesis of Periodic Classification

1. The places that were left empty by Mendeleev in his periodic table were for
   (a) aluminium and silicon
   (b) gallium and germanium
   (c) arsenic and antimony
   (d) molybdenum and tungsten.

2. The element cited as an example to prove the validity of Mendeleev's periodic law is
   (a) germanium (b) scandium
   (c) gallium (d) all of these.

3. As predicted by Johann Dobereiner's Triads, elements in triads are found with sequence of three similar elements, when the middle element has a mass equal to the average of the least and most massive.

Elements	Mass number
A	—
B	23
C	29

   What is mass number of A?
   (a) 7 (b) 8 (c) 31 (d) 17

4. In Lothar Meyer plot of atomic volume versus atomic mass, the peaks are occupied by
   (a) alkali metals (b) alkaline earth metals
   (c) halogens (d) noble gases.

5. Which is correct match?
   (a) Eka silicon-Ge (b) Eka aluminium-Ga
   (c) Both (a) and (b) (d) None of these

6. Which of the following statements is not correct about modified Mendeleev's periodic table?
   (a) It consists of nine groups and seven periods.
   (b) Group VIII like group I-VII has been divided into two sub-groups A and B.
   (c) The group of an element in the periodic table represents its valency.
   (d) The elements of group IB are called coinage metals.

7. In Lother Meyer's curve, the halogens occupy positions on the
   (a) descending portion of the curve
   (b) ascending portion of the curve
   (c) peak portion of the curve
   (d) no fixed position on the curve.

8. Which of the following is known as the bridge element of 2nd group in Mendeleev's table?
   (a) Zn (b) Sr (c) Mg (d) Hg

Modern Periodic Law and Long Form of Periodic Table

9. The number of elements in each period is
   (a) twice the number of atomic mass of the elements
   (b) twice the number of atomic orbitals available in the energy level being filled
   (c) same as the sum of atomic numbers of all elements
   (d) none of these.

10. If there were 10 periods in periodic table then maximum number of elements it can have is
   (a) 290 (b) 770
   (c) 204 (d) none of these.

11. The long form of periodic table has
   (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.

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

13. Which of the following expressions is correct?
   (a) $v = a(z - b)$ (b) $v^2 = a(z - b)^2$
   (c) $v = a^2(z - b)^2$ (d) $v^2 = a^2(z - b)^2$

Electronic Configuration of Elements and Periodic Table

14. The electronic configuration of an element is $1s^2$, $2s^2 2p^6$, $3s^2 3p^4$. The atomic number of the element present just below this element in the periodic table is
   (a) 36 (b) 34 (c) 33 (d) 32

15. Which one of the following is a metalloid?
   (a) Phosphorus (b) Antimony
   (c) Nitrogen (d) Bismuth

16. The element with the electronic configuration as [Ar] $3d^{10}4s^2 4p^3$ represents a
   (a) metal (b) non-metal
   (c) metalloid (d) transition element.

17. The electronic configuration $1s^2 2s^2 2p^6 3s^2 3p^5$ represents
   (a) electronegative element
   (b) an element from 18th group
   (c) 4th period and 5th group
   (d) none of these.

18. Which one pair of atoms or ions will have same configuration?
(a) F^+ and Ne (b) Li^+ and He^-
(c) Na and K (d) Cl^- and Ar

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

20. In the long form of the periodic table, the valence shell electronic configuration of $5s^2 5p^4$ corresponds to the element present in,

	Group	Period
(a)	16	6
(b)	17	5
(c)	16	5
(d)	17	6

21. Pd has exceptional valence shell electronic configuration of $4d^{10}5s^0$. It is a member of
(a) 5th Period, Group 10 (b) 4th Period, Group 12
(c) 6th Period, Group 10 (d) 5th Period, Group 14

22. An element has the electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$.
What will be its position in the periodic table?
(a) Period 4, Group 10 (b) Period 2, Group 2
(c) Period 4, Group 2 (d) Period 2, Group 8

23. The electronic configuration of the element which is just above the element with atomic number 43 in the same periodic group is
(a) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^5, 4s^2$
(b) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^{10}, 4s^2\ 4p^5$
(c) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^6, 4s^1$
(d) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^{10}, 4s^1\ 4p^6$

24. Lanthanoids are
(a) 14 elements in the seventh period (Atomic no. = 58 to 71) that are filling 4*f* subshell
(b) 14 elements in the sixth period (atomic no. = 90 to 103) that are filling 4*f* subshell
(c) 14 elements in the seventh period (atomic no. = 90 to 103) that are filling 5*f* subshell
(d) 14 elements in the sixth period (atomic no. = 58 to 71) that are filling 4*f* subshell.

25. The electronic configuration of four different elements is given below. Identify the group IVA element among these.
(a) [He] $2s^1$ (b) [Ne] $3s^2$
(c) [Ne] $3s^2\ 3p^2$ (d) [Ne] $3s^2\ 3p^5$

26. In the absence of Aufbau rule and also assume that each orbital can take maximum of three electrons, then number of elements in different periods are

Period	3	4	5
(a)	18	32	50
(b)	18	18	32
(c)	27	27	48
(d)	27	48	75

27. In the sixth period of the extended form of periodic table, the orbitals are filled as
(a) 6*s*, 5*f*, 6*d*, 6*p* (b) 5*s*, 5*p*, 5*d*, 6*p*
(c) 6*s*, 6*p*, 6*d*, 6*f* (d) 6*s*, 4*f*, 5*d*, 6*p*.

28. Which of the following have the same number of electrons in outermost shell?
(a) Elements with atomic numbers 30, 48, 80
(b) Elements with atomic numbers 14, 15, 16
(c) Elements with atomic numbers 20, 30, 50
(d) Elements with atomic numbers 10, 18, 26

29. Electronic configuration $[Xe]4f^{14}\ 5d^2\ 6s^2$ belongs to
(a) 5th group (b) 4th group
(c) 6th group (d) 3rd group

30. Which of the following groups contains metals, non-metals and metalloids?
(a) Group 17 (b) Group 14
(c) Group 13 (d) Group 12

31. Out of the following, which set does not show correct matching?
(a) Sc^{3+} : [Ne] $3s^2 3p^6$, zero group
(b) Fe^{2+} : [Ar] $3d^6$, 8th group
(c) Cr : [Ar] $3d^5 4s^1$, 6th group
(d) All of these

32. An element *X* belongs to fourth period and fifteenth group of the periodic table. Which one of the following is true regarding the outer electronic configuration of *X*? It has
(a) partially filled *d*-orbitals and completely filled *s*-orbital
(b) completely filled *s*-orbital and completely filled *p*-orbitals
(c) completely filled *s*-orbital and half filled *p*-orbitals
(d) completely filled *s*-, *p*- and *d*-orbitals.

33. The higher oxide of an element (*E*) has the formula EO_3. Its hydride contains 2.47% hydrogen, the element is
(a) Te (b) Se
(c) S (d) Si

34. The electronic configuration and the group number in the periodic table in which the element with atomic number 107 lies are
(a) [Rn] $5f^{14}\ 6d^1\ 7s^2\ 7p^4$, Group 3
(b) [Rn] $5f^{14}\ 6d^5\ 7s^2$, Group 7
(c) [Rn] $5f^{14}\ 7s^2\ 7p^5$, Group 7
(d) [Rn] $5f^{14}\ 6d^2\ 7s^2\ 7p^3$, Group 15

35. Which of the following is/are considered as metalloids?
(a) As, Sb (b) Po, Sb
(c) Te, Ge (d) All of these

36. Which of the following electronic configurations is of transition elements?
(a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
(b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^1$
(c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$
(d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$

37. Electronic configurations of few elements are given below. Mark the incorrect match.

(a) $1s^2 2s^2 2p^5$ – Most electronegative element

(b) $1s^2 2s^2 2p^3$ – An element belonging to 3rd period and 5th group

(c) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$ – A d-block element

(d) $1s^2 2s^2 2p^6 3s^2 3p^6$ – An element from 18th group

38. $ns^2 np^4$ (n-outermost orbit) represents the valence electrons. The corresponding group would be

(a) F, Cl, Br .... (b) N, P, As ....

(c) O, S, Se .... (d) C, Si, Ge ....

39. Which is the most non-metallic element among the following?

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

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

40. Electronic configurations of elements (ground state or excited states) are given

I. [Ar] $4s^1 3d^1$ II. $1s^2 2s^2 2p^5 3s^1$

III. [Ar] $3d^{10} 4s^1$ IV. [Xe] $4f^{14} 5d^1 6s^2$

These elements respectively belong to .... blocks.

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

(c) s, p, d, f (d) s, p, f, d

❑ Periodic Trends in Properties of Elements

41. Which of the following are the values of electron affinities (in kJ/mol) for the formation of O^- and O^{2-} from O?

(a) –142, –780 (b) –141, 780

(c) 142, 780 (d) –142, –142

42. The size of isoelectronic species F^-, Ne, Na^+ and Mg^{2+} is affected by

(a) nuclear charge (Z)

(b) interelectronic repulsion in the valence shell

(c) valence principal quantum number (n)

(d) none of these factors.

43. The maximum valency of an element having atomic number seven is

(a) 1 (b) 3

(c) 5 (d) 7

44. What would be the bond length of C–X bond, if C–C bond length is 1.54 Å, X–X bond length is 1.00 Å and electronegativity values of C and X are 2.0 and 3.0 respectively.

(a) 0.50 Å (b) 0.70 Å

(c) 1.18 Å (d) 0.27 Å

45. In any period, the valency of an element with respect to hydrogen

(a) increases one by one from IA to VIIA

(b) decreases one by one from IA to VIIA

(c) increases one by one from IA to IVA and then decreases from VA to VIIA one by one

(d) decreases one by one from IA to IVA and then increases from VA to VIIA one by one.

46. An example of amphoteric oxide is

(a) Ti_2O_2 (b) MgO

(c) Cl_2O_7 (d) Al_2O_3

47. Why is the electron gain enthalpy of O or F less than that of S or Cl?

(a) O and F are more electronegative than S and Cl.

(b) When an electron is added to O or F, it goes to a smaller (n = 2) level and suffers more repulsion than the electron in S or Cl in larger level (n = 3).

(c) Adding an electron to $3p$-orbital leads to more repulsion than $2p$-orbital.

(d) Electron gain enthalpy depends upon the electron affinity of the atom.

48. Which of the following should be correct for Z_1 and Z_2 in the following two processes :

$$M^+ + Z_1 \rightarrow M^{2+} + e^-$$

$$M^{2+} + Z_2 \rightarrow M^{3+} + e^-$$

(a) $\frac{1}{2}Z_1 = Z_2$ (b) $Z_1 = Z_2$

(c) $Z_1 = \frac{1}{2}Z_2$ (d) $Z_1 < Z_2$

49. Fluorine is the most reactive among all the halogens, because of it's

(a) small size

(b) low electronegativity

(c) large size

(d) high dissociation energy of F – F bond.

50. $O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)} - E, E = +780$ kJ/mol

What does value of E indicates

(a) endothermic reaction

(b) exothermic reaction

(c) both (a) and (b)

(d) none of these.

51. Which of the following oxides is most basic?

(a) Bi_2O_3 (b) SeO_2

(c) Al_2O_3 (d) Sb_2O_3

52. Which oxide of N is isoelectronic with CO_2?

(a) NO_2 (b) NO

(c) N_2O (d) N_2O_3

53. From the following electronic configuration which will have the highest electron affinity?

(a) $1s^2 2s^2 2p^3$ (b) $1s^2 2s^2 2p^5$

(c) $1s^2 2s^2 2p^6 3s^2 3p^5$ (d) $1s^2 2s^2 2p^6 3s^2 3p^3$.

54. The first ionization potential will be maximum for

(a) uranium (b) iron

(c) hydrogen (d) lithium.

55. The ionic radii of N^{3-} and F^- are 1.76 Å and 1.36 Å respectively. The ionic radius of O^{2-} will be

(a) 1.40 Å (b) 1.96 Å

(c) 1.20 Å (d) 2.06 Å

56. Which of the following relation is correct?

(a) 2 $I.P.$ – $E.A.$ – $E.N.$ = 0 (b) 2 $I.P.$ – $E.N.$ + $E.A.$ = 0

(c) 2 $E.N.$ – $I.P.$ – $E.A.$ = 0 (d) $E.N.$ – $I.P.$ – $E.A.$ = 0

57. The atomic radius decreases in a period due to
(a) increase in nuclear attraction
(b) decrease in nuclear attraction
(c) increase in number of electrons
(d) decrease in number of electrons.

58. Calculate the energy needed to convert three moles of sodium atoms in the gaseous state to sodium ions. The ionization energy of sodium is 495 kJ mol^{-1}.
(a) 1485 kJ (b) 495 kJ
(c) 148.5 kJ (d) None of these

59. Sodium forms Na^+ ion but it does not form Na^{2+} because of
(a) very low value of $(IE)_1$, and $(IE)_2$
(b) very high value of $(IE)_1$, and $(IE)_2$
(c) low value of $(IE)_1$ and low value of $(IE)_2$
(d) low value of $(IE)_1$ and high value of $(IE)_2$.

60. Enthalpy change in the following process is $M + e^- \longrightarrow M^-$, $\Delta H = -X$ kJ $mole^{-1}$.
Which of the following process have enthalpy change $= X$ kJ $mole^{-1}$?
(a) $M^- \longrightarrow M + e^-$ (b) $M^+ + e^- \longrightarrow M$
(c) $M^{2+} + e^- \longrightarrow M^+$ (d) $M + e^- \longrightarrow M^-$

61. The *E.N.* of H, *X*, O are 2.1, 0.8 and 3.5 respectively comment on the nature of the compound H–O–*X*, that is
(a) basic (b) acidic
(c) amphoteric (d) can't be predicted.

62. Diagonal relationship is shown by
(a) elements of first period
(b) elements of second period
(c) elements of third period
(d) (b) and (c) both.

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

64. The Z_{eff} for
3*d* electron of Cr
4*s* electron of Cr
3*d* electron of Cr^{3+}
are in the order respectively
(a) 4.6, 2.95, 4.95 (b) 4.95, 2.95, 4.6
(c) 4.6, 2.95, 5.3 (d) none of these.

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

66. Valence electrons in the element *A* are 3 and that in element *B* are 6. Most probable compound formed from *A* and *B* is
(a) A_2B (b) AB_2
(c) A_6B_3 (d) A_2B_3

67. The first element of a group in many ways differs from the other heavier members of the group. This is due to
(a) the small size
(b) the high electronegativity and high ionization potential
(c) the unavailability of *d*-orbitals
(d) all of the above.

68. Chloride ion and potassium ion are isoelectronic. Then
(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.

69. Which among the following species has the same number of electrons in its outermost as well as penultimate shell?
(a) Mg^{2+} (b) O^{2-} (c) F^- (d) Ca^{2+}

70. Property of alkaline earth metals that increases with their atomic number is
(a) ionization energy
(b) solubility of their hydroxides
(c) solubility of their sulphates
(d) electronegativity.

71. Among the elements *W*, *X*, *Y* and *Z* having atomic numbers 9, 10, 11 and 12 respectively, the correct order of ionization energies is
(a) $W > Y > X > Z$ (b) $X > W > Z > Y$
(c) $X > Z > Y > W$ (d) $Z > Y > X > W$

72. Atomic radius decreases in a period, but after halogens, the atomic radius suddenly increases. Thus, inert gases have almost highest radius in a period. The explanation for such an increase is
(a) inert gases have most stable configuration
(b) inert gases do not take part in bonding
(c) van der Waal's radius is reported in case of inert gases
(d) none of these.

73. 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

74. Let electronegativity, ionization energy and electronic affinity be represented as *EN*, *IP* and *EA* respectively. Which one of the following equation is correct according to Mulliken?
(a) $EN = IP \times EA$ (b) $EN = IP/EA$
(c) $EN = \dfrac{IP + EA}{2}$ (d) $EN = IP - EA$

75. The correct order of electronegativity for O, O^+ and O^- is
(a) $O^- > O > O^+$ (b) $O > O^+ > O^-$
(c) $O^+ > O^- > O$ (d) $O^+ > O > O^-$

76. In the periodic table, the maximum chemical reactivity is at the extreme left (alkali metals) and extreme right (halogens). Which properties of these two groups are responsible for this?
(a) Least ionisation enthalpy on the left and highest negative electron gain enthalpy on the right.
(b) Non-metallic character on the left and metallic character on the right.
(c) Low atomic radii on the left and large atomic radii on the right.
(d) Highest electronegativity on the left and least electronegativity on the right.

77. Few values are given in the table in the direction from left to right and top to bottom. Predict the property which could be depicted in the table.

152						
186	160	143	117	110	104	99
231						
244						
262						

(a) Atomic number (b) Ionisation enthalpy
(c) Atomic radius (d) Electron gain enthalpy

78. Screening effect is not observed in
(a) He^+ (b) Li^{2+}
(c) Be^{3+} (d) all of the above.

79. Elements P, Q, R and S belong to the same group. The oxide of P is acidic, oxide of Q and R are amphoteric while the oxide of S is basic. Which of the following elements is the most electropositive?
(a) P (b) Q (c) R (d) S

80. In which of the following, the order is not in accordance with the property mentioned.
(a) Li < Na < K < Rb – Atomic radius
(b) F > N > O > C – Ionisation enthalpy
(c) Si < P < S < Cl – Electronegativity
(d) F < Cl < Br < I – Electronegativity

81. Which of the following statements is true about effective nuclear charge?
(a) Z_{eff} decreases from top to bottom.
(b) Z_{eff} increases from top to bottom.
(c) Z_{eff} increases as we move from left to right in periodic table.
(d) $Z_{eff} = Z \times \sigma$ (here σ is screening constant).

82. The radius of iso-electronic species
(a) increases with increase in nuclear charge
(b) decreases with increase in nuclear charge
(c) same for all
(d) first increases and then decreases.

83. For the process
$X_{(g)} + e^- \longrightarrow X^-_{(g)}$, $\Delta H = x$ and
$X^-_{(g)} \longrightarrow X_{(g)} + e^-$, $\Delta H = y$
select correct alternatives.
(a) Ionization energy of $X^-_{(g)}$ is y.
(b) Electron affinity of $X_{(g)}$ is x.
(c) Electron affinity of $X_{(g)}$ is $-y$.
(d) All are correct statements.

84. For an element, IE values are
$(IE)_1 = 738$ kJ mol^{-1} $(IE)_2 = 1450$ kJ mol^{-1}
$(IE)_3 = 7.7 \times 10^3$ kJ mol^{-1} $(IE)_4 = 1.1 \times 10^4$ kJ mol^{-1}
this element belongs to
(a) alkali metals (b) alkaline earth metals
(c) chalcogens (d) halogens.

85. Anything that influences the valence electrons will affect the chemistry of the element. Which one of the following factors does not affect the valence shell?
(a) Principal quantum number (n)
(b) Nuclear charge (Z)
(c) Nuclear mass
(d) Number of core electrons

86. As we move from left to right, the electronegativity increases. An atom which is highly electronegative has
(a) large size (b) low electron affinity
(c) high ionisation enthalpy
(d) low chemical reactivity.

87. Bond distance C–F in (CF_4) and Si–F in (SiF_4) are respectively 1.33 Å and 1.54 Å. C–Si bond is 1.87 Å. Calculate the covalent radius of F atom, ignoring the electronegativity differences.
(a) 0.64 Å (b) $\frac{1.33+1.54+1.8}{3}$ Å
(c) 0.5 Å (d) $\frac{1.54}{2}$ Å

88. The shielding effect of d-electrons is
(a) more than s-electrons (b) more than p-electrons
(c) less than s-electrons (d) same as f-electrons.

89. Which of the following transitions involve maximum amount of energy?
(a) $M^-_{(g)} \longrightarrow M_{(g)}$ (b) $M^-_{(g)} \longrightarrow M^+_{(g)}$
(c) $M^+_{(g)} \longrightarrow M^{2+}_{(g)}$ (d) $M^{2+}_{(g)} \longrightarrow M^{3+}_{(g)}$

90. (A), (B) and (C) are elements in the second short period. Oxide of (A) is ionic, that of (B) is amphoteric and of (C) a gaint molecule. (A), (B) and (C) have atomic number in the order-
(a) (A) < (B) < (C) (b) (C) < (B) < (A)
(c) (A) < (C) < (B) (d) (B) < (A) < (C).

91. Which one of the following group of atoms or ions is not isoelectronic?
(a) He, H^-, Li^+ (b) Na^+, Mg^{2+}, Al^{3+}
(c) F^-, O^{2-}, N^{3-} (d) K^+, Ca^{2+}, Ne

92. Second and successive electron affinity of an element
(a) is always negative (energy is released)
(b) is always positive (energy is absorbed)
(c) can be positive or negative
(d) is always zero.

93. Which of the following ionisation energy values for calcium show a sudden increase?
(a) Third (b) Second
(c) First (d) Fourth

94. Which is true about the electronegativity order of the following elements?
(a) P > Si (b) C > N (c) Br > Cl (d) Sr > Ca

95. The ionisation potential of nitrogen is more than that of oxygen molecules because of

(a) greater attraction of electrons by the nucleus
(b) extra stability of the half-filled p-orbitals
(c) smaller size of nitrogen
(d) more penetrating effect.

96. Downwards in a group, the electropositive character of elements
(a) increases (b) decreases
(c) remains same (d) none of these.

97. Which of the following oxides is most basic?
(a) Na_2O (b) MgO (c) Al_2O_3 (d) CuO

98. IE_1 and IE_2 of Mg are 178 and 348 kcal mol^{-1}. The energy required for the reaction $Mg \longrightarrow Mg^{2+} + 2e^-$ is
(a) + 170 kcal (b) +526 kcal
(c) – 170 kcal (d) – 526 kcal

99. The *EA* for inert gases is likely to be
(a) high (b) small (c) zero (d) positive.

100. Ionisation of energy F^- is 320 kJ mol^{-1}. The electron gain enthalpy of fluorine would be
(a) – 320 kJ mol^{-1} (b) –160 kJ mol^{-1}
(c) + 320 kJ mol^{-1} (d) + 160 kJ mol^{-1}

2 Check Your Understanding

1. Elements, P, Q, R and S have the following electronic configurations
P : $[Ar]3d^{10}4s^1$
Q : $[Ar]3d^{10}4s^24p^65s^1$
R : $[Ar]3d^{10}4s^24p^64d^{10}5s^1$
S : $[Ar]3d^{10}4s^24p^64d^55s^1$
Which two elements fall into the same group?
(a) Q and R (b) P and R
(c) Q and S (d) P and Q

2. Which of the following statements with regard to the modern periodic table is incorrect?
(a) Each block contains the same number of columns as the number of electrons that can occupy the subshell.
(b) A block suggests the value of the azimuthal quantum number (l) of the last subshell that receives electrons to build up the electronic configuration of the atoms.
(c) The p-block has 6 columns.
(d) The d-block has 8 columns.

3. Which of the following sets gives the correct order of atomic radius?
(a) B < Al < Mg < Ca (b) Al < B < Ca < Mg
(c) Mg < Al < B < Ca (d) Ca < Mg < Al < B

4. Why Sc (Z = 21) is not considered as a transition element?
(a) Properties of Sc are similar to alkali metals.
(b) 3d-orbitals are empty in its stable compound.
(c) Stable oxidation number of Sc is +2.
(d) Atomic volume of Sc is very large.

5. The chemistry of lithium is very similar to that of magnesium even though they are placed in different groups. Its reason is
(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 and size (*i.e.*, charge density) is nearly the same.

6. Which is correct increasing order of their tendency of the given elements to form M^{3-} ion?
(a) Bi > Sb > As > P > N (b) Bi < Sb < As < P < N
(c) N < P < Sb < Bi < As (d) Bi > Sb ~ N ~ P > As

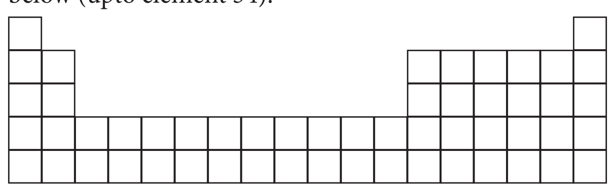
7. The electronegativity difference between N and F is greater than that between N and H yet the dipole moment of NH_3 (1.5 D) is larger than that of NF_3 (0.2 D). This is because
(a) in NH_3 as well as NF_3 the atomic dipole and bond dipole are in opposite directions
(b) in NH_3 the atomic dipole and bond dipole are in the opposite directions whereas in NF_3 , these are in the same direction
(c) in NH_3 as well as in NF_3 the atomic dipole and bond dipole are in the same direction
(d) in NH_3, the atomic dipole and bond dipole are in the same direction whereas in NF_3 these are in opposite direction.

8. Among the following, the correct order of radius is
(a) Na < Be < B (b) $F^- < O^{2-} < N^{3-}$
(c) Na < Li < K (d) $Fe^{3+} < Fe^{2+} < Fe^{4+}$.

9. In a given energy level, the order of penetration effect of different orbitals is
(a) $f < d < p < s$ (b) $s = p = d = f$
(c) $s < p < d < f$ (d) $p > s > d > f$.

10. Sum of first three ionization energies of Al is 53.0 eV atom^{-1} and the sum of first two ionization energies of Na is 52.2 eV atom^{-1}. Out of Al (III) and Na(II),
(a) Al (III) is more stable than Na(II)
(b) Na (II) is more stable than Al (III)
(c) both are equally unstable
(d) both are equally stable.

11. The overall layout of the empty periodic table is shown below (upto element 54).



Answer the following question:
If X represents an element of atomic number 9 and Y an element of atomic number 20, the compound formed by these elements would be

(a) ionic with formula YX_2
(b) covalent with formula YX_2
(c) ionic with formula YX
(d) covalent with formula YX.

12. The oxide of which of the following elements will be acidic in character?
(a) Mg (b) Rb (c) Li (d) Cl

13. Which of the following is an incorrect statement?
(a) The first ionization enthalpy of Al is smaller than the first ionization enthalpy of Mg.
(b) The second ionization enthalpy of Mg is greater than the second ionization enthalpy of Na.
(c) The first ionization enthalpy of Na is smaller than the first ionization enthalpy of Mg.
(d) The third ionization enthalpy of Mg is greater than the third ionization enthalpy of Al.

14. The hydrides of the first elements in groups 15-17, namely NH_3, H_2O and HF respectively show abnormally high values for melting and boiling points. This is due to
(a) small size of N, O and F
(b) the ability to form extensive intermolecular H-bonding
(c) the ability to form extensive intramolecular H-bonding
(d) effective van der Waals' interaction.

15. Which of the following statements are correct?
(1) HF is a stronger acid than HCl.
(2) Among halide ions, iodide is the most powerful reducing agent.
(3) Fluorine is the only halogen that does not show a variable oxidation state.
(4) HOCl is a stronger acid than HOBr.
(a) 2 and 4 (b) 2 and 3
(c) 1, 2 and 3 (d) 2, 3 and 4

16. Which of the following statements concerning ionisation energy is not correct?
(a) The IE_2 is always more than the first.
(b) Within a group, there is a gradual increase in ionisation energy because nuclear charge increases.
(c) Ionisation energies of Be is more than B.
(d) Ionisation energies of noble gases are high.

17. Which of the following represents the correct order of the property indicated?
(a) $Sc^{3+} > Cr^{3+} > Fe^{3+} > Mn^{3+}$ – Ionic radii
(b) $Sc < Ti < Cr < Mn$ – Density
(c) $Mn^{2+} > Co^{2+} > Ni^{2+} < Fe^{2+}$ – Ionic radii
(d) $FeO < CaO < MnO < CuO$ – Basic nature

18. The first, second and third ionization enthalpies of an element are 737, 1450 and 7731 kJ mol^{-1} respectively. What will be the formulae of its oxide and chloride?
(a) M_2O, MCl (b) MO, MCl_2
(c) M_2O_3, MCl_3 (d) MO_2, MCl_4

19. Which is the correct order of atomic sizes?
(a) Ce > Sn > Yb > Lu (b) Sn > Ce > Lu > Yb
(c) Lu > Yb > Sn > Ce (d) Sn > Yb > Ce > Lu.
(At. Nos. : Ce = 58, Sn = 50, Yb = 70 and Lu = 71)

20. The electronic configuration of elements A, B and C are [He] $2s^1$, [Ne] $3s^1$ and [Ar] $4s^1$ respectively. Which one of the following order is correct for the first ionization potentials (in kJ mol^{-}) of A, B and C?
(a) $A > B > C$ (b) $C > B > A$
(c) $B > C > A$ (d) $C > A > B$

21. The ions O^{2-}, F^-, Na^+, Mg^{2+}, and Al^{3+} are isoelectronic. Their ionic radii show
(a) an increase from O^{2-} to F^- and then a decrease from Na^+ to Al^{3+}
(b) a decrease from O^{2-} to F^- and then an increase from Na^+ to Al^{3+}
(c) a significant increase from O^{2-} to Al^{3+}
(d) a significant decrease from O^{2-} to Al^{3+}.

22. In which of the following pairs of species, the size of the first species is not more than the second species?
(a) Na^+, F^- (b) Fe^{2+}, Fe^{3+} (c) Li, F (d) S, O

23. The correct order of basic character is
(a) $Na_2O < MgO < Al_2O_3$ (b) $CaO < SrO < BaO$
(c) $MnO < Mn_2O_3 < MnO_2$ (d) $Cr_2O_3 < CrO_3$

24. Generally, the first ionization energy increases along a period. But there are some exceptions. The one which is NOT an exception is
(a) Na and Mg (b) Be and B
(c) N and O (d) Mg and Al

25. For one of the element various successive ionisation enthalpies (in kJ mol^{-1}) are given below :

I.E.	1st	2nd	3rd	4th	5th
	577.5	1810	2750	11,580	14,820

The element is
(a) P (b) Mg (c) Si (d) Al

26. Arrange the following oxides in order of increasing acidic character. Na_2O, Cl_2O_7, As_2O_3, N_2O.
(a) $Na_2O < As_2O_3 < N_2O < Cl_2O_7$
(b) $Na_2O < N_2O < As_2O_3 < Cl_2O_7$
(c) $N_2O < Cl_2O_7 < Na_2O < As_2O_3$
(d) $Cl_2O_7 < N_2O < As_2O_3 < Na_2O$

27. The electronic configurations of four elements are given below. Arrange these elements in the correct order of magnitude (without sign) of their electron affinity
(i) $2s^2\,2p^5$ (ii) $3s^2\,3p^5$ (iii) $2s^2\,2p^4$ (iv) $3s^2\,3p^4$
Select the correct answer using the codes given below :
(a) (i) < (ii) < (iii) < (iv) (b) (ii) > (i) > (iv) > (iii)
(c) (i) < (iii) < (iv) < (ii) (d) (iii) < (iv) < (ii) < (i)

28. Which of the following statements is true for the elements of period 2?
(a) They have the lowest negative electron-gain enthalpies in their groups.
(b) They have the lowest ionization enthalpies in their groups.
(c) They have the lowest electronegativities in their groups.
(d) Being small in size and devoid of d-orbitals, they differ from the other elements of their groups.

29. For the gaseous reaction, $K + F \longrightarrow K^+ + F^-$, ΔH was calculated to be 19 kcal under conditions where the cations and anions were prevented by electrostatic separation from combining with each other. The ionization potential of K is 4.3 eV. What is the electron affinity of F?
(a) 3.21 (b) 4.28 (c) 3.48 (d) 1.48

30. Which of the following represents increasing order of ionic conductance?
(a) $F^- < Cl^- < Br^- < I^-$ (b) $I^- < Br^- < Cl^- < F^-$
(c) $F^- < Cl^- < I^- < Br^-$ (d) $F^- < I^- < Cl^- < Br^-$

31. Consider the iso-electronic series, K^+, S^{2-}, Cl^-, Ca^{2+}, the radii of the ions decrease as
(a) $Ca^{2+} > K^+ > Cl^- > S^{2-}$ (b) $Cl^- > S^{2-} > K^+ > Ca^{2+}$
(c) $S^{2-} > Cl^- > K^+ > Ca^{2+}$ (d) $K^+ > Ca^{2+} > S^{2-} > Cl^-$

32. Two p-block elements x (outer configuration $ns^2\ np^3$) and z (outer configuration $ns^2\ np^4$) occupy neighbouring positions in a period. Using this information which of the following is correct with respect to their ionisation potential I_x and I_z?
(a) $I_x > I_z$ (b) $I_z > I_x$
(c) $I_z = I_x$
(d) Relation between I_x and I_z is uncertain

33. Which element can exhibit more than one oxidation state in compounds?
1. Cr, 2. Pb, 3. Sr
(a) 1 only (b) 1 and 2 only
(c) 2 and 3 only (d) 1, 2 and 3

34. $N_0/2$ atoms of $X(g)$ are converted into $X^+(g)$ by energy E_1. $N_0/2$ atoms of $X(g)$ are converted into $X^-(g)$ by energy E_2. Hence, ionisation potential and electron affinity of $X(g)$ are
(a) $\frac{2E_1}{N_0}, \frac{2(E_1 - E_2)}{N_0}$ (b) $\frac{2E_1}{N_0}, \frac{2E_2}{N_0}$
(c) $\frac{(E_1 - E_2)}{N_0}, \frac{2E_2}{N_0}$ (d) none is correct.

35. The formation of the oxide ion $O^{2-}_{(g)}$ require first an exothermic and then an endothermic step as shown below :
$$O_{(g)} + e^- \longrightarrow O^-_{(g)}; \quad \Delta H° = -142 \text{ kJ mol}^{-1}$$
$$O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)}; \quad \Delta H° = 780 \text{ kJ mol}^{-1}$$
This is because
(a) oxygen is more electronegative
(b) oxygen has high electron affinity
(c) O^- ion has comparatively larger size than oxygen atom
(d) O^- ion will tend to resist the addition of another electron.

36. Two elements X and Y contain only one electron in the outer level. Element X is reactive and loses electron easily while element Y is relatively unreactive and non-corrosive. The elements X and Y respectively are
(a) Cs and Li (b) Rb and Na
(c) Li and Cu (d) Ag and Au

37. Considering the elements B, Al, Mg and K, the correct order of their metallic character is
(a) B > Al > Mg > K (b) Al > Mg > B > K
(c) Mg > Al > K > B (d) K > Mg > Al > B

38. Which of the oxides behaves both as suboxide and as neutral oxide?
(a) CO (b) CO_2 (c) C_3O_2 (d) N_2O

39. The statement that is not true for the long form of the periodic table is
(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 bond between any two elements.

40. Select equation(s) having exothermic step.
(I) $S^-_{(g)} \rightarrow S^{2-}_{(g)}$
(II) $Na^+_{(g)} + Cl^-_{(g)} \rightarrow NaCl_{(s)}$
(III) $N_{(g)} \rightarrow N^-_{(g)}$
(IV) $Al^{2+}_{(g)} \rightarrow Al^{3+}_{(g)}$
Choose the correct code.
(a) II (b) I and II
(c) III and IV (d) II and III

41. Aqueous solutions of two compounds M—O—H and M'—O—H have been prepared in two different beakers. If the electronegativity of $M = 3.5$, $M' = 1.72$, O = 3.0 and H = 2.1, then the solutions respectively are
(a) acidic, acidic (b) acidic, basic
(c) basic, basic (d) basic, acidic.

42. Which of the oxides is acidic as well as neutral oxide?
(a) N_2O (b) Na_2O
(c) NO (d) H_2O

43. If the ionization enthalpy and electron gain enthalpy of an element are 275 and 86 kcal mol^{-1} respectively, then the electronegativity of the element on the Pauling scale is
(a) 2.8 (b) 0.0
(c) 4.0 (d) 2.6

44. Considering the elements B, C, N, Si and F the correct order of their non-metallic character is?
(a) B > C > Si > N > F (b) Si > C > B > N > F
(c) F > N > C > B > Si (d) F > N > C > Si > B

45. Which one of the following statements is incorrect?
(a) Greater is the nuclear charge, greater is the electron affinity.
(b) Neon has zero electron affinity.
(c) Electron affinity decreases from fluorine to iodine in the group.
(d) Electron affinity decreases in going down a group and increases across period from the left to the right.

46. Given below are the names of few elements based on their position in the periodic table. Identify the element which is not correctly placed.

(a) An element which tends to lose three electrons – Aluminium
(b) An element which tends to gain two electrons – Iodine
(c) An element with valency four – Silicon
(d) A transuranium element – Plutonium

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

48. The electronegativity values of C, N, O and F
(a) increases from carbon to fluorine
(b) decreases from carbon to fluorine
(c) increases up to oxygen and is minimum at fluorine
(d) is minimum at nitrogen and then increases continuously.

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

50. Electropositive or metallic character
(a) increases in a period
(b) decreases in a group
(c) decreases in a period and increases in a group
(d) of an element is reflected in its tendency to form covalent compounds.

3 Mixed Category

1. Which of the following statements is not correct for modern Mendeleev's periodic table?
(a) It consists of nine groups.
(b) It consists of seven periods.
(c) Each group has equal number of elements.
(d) Each period starts with the element of alkali group and ends with an element of zero group.

2. Which of the following statements is not correct about Lother Meyer's classification in the form of curve?
(a) The elements present at the peaks are chemically very active.
(b) Alkaline earth metals are present at the descending portions of the curve.
(c) Halogens occupy ascending portions of the curve.
(d) The elements present in the troughs are chemically very reactive and are known as representative elements.

3. In the long form (Modern form) of Periodic Table, following are some of the facts :
I. Periods are horizontal rows and groups are vertical columns.
II. Transition elements have been placed in 4^{th} to 7^{th} periods.
III. Inner-transition elements have been placed in 6^{th} and 7^{th} periods and grouped at the bottom of the periodic table in 3^{rd} group.
IV. Elements (58 - 71) are called actinides and elements (90 - 103) are called lanthanides.
Select the incorrect fact(s).
(a) Only III and IV (b) Only I and II
(c) Only III (d) Only IV

4. Which of the following statements is not correct?
(a) 'Law of triads' was enunciated by Dobereiner.
(b) 'Law of octave' was presented by Newland.
(c) The periodic law presented by Mendeleev was based on atomic masses of the elements.
(d) The attempt for classifying elements by plotting the atomic masses of elements against the volumes was made by Moseley.

5. Following are some of the defects of Mendeleev's Periodic Table :
I. Position of hydrogen was not defined.
II. Positions of lanthanides and actinides were not defined.
III. There was no discussion about noble gases.
IV. Isotopes were placed in the same position.
Select correct defects in it.
(a) I, III and IV (b) II and III
(c) I, II and III (d) All of these

6. Consider the following statements :
(i) Total inner-transition elements are 28.
(ii) Total *d*-block elements are 40.
(iii) Except hydrogen, all non-metals belong to *p*-block series.
(iv) All the members of the actinide series are man-made (synthetic).
The correct statements are
(a) (i) and (iv) (b) (i), (ii) and (iii)
(c) (iii) and (iv) (d) all of the above.

7. Consider the following statements :
I. The radius of an anion is larger than that of parent atom.
II. The *I.E.* increases from left to right in a period generally.
III. The electronegativity of an element is the tendency of an isolated atom to attract an electron.
The correct statements are
(a) I only (b) II only
(c) I and II only (d) II and III only.

8. Match the elements given in column-II with their type given in column-I and select the correct option.

	Column-I		Column-II
(A)	Metalloid	(p)	Sulphur
(B)	Radioactive	(q)	Gold
(C)	Transition metal	(r)	Arsenic
(D)	Chalcogen	(s)	Uranium

(a) A → r; B → s; C → q; D → p
(b) A → q; B → s; C → p; D → r
(c) A → s; B → r; C → q; D → p
(d) A → s; B → p; C → r; D → q

9. Match column-I with column-II and select the correct option.

	Column-I (Element)		Column-II (Electronic configuration)
(A)	Gallium	(p)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
(B)	Vanadium	(q)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$
(C)	Zinc	(r)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$
(D)	Scandium	(s)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

(a) A → q; B → r; C → p; D → s
(b) A → r; B → s; C → p; D → q
(c) A → r; B → s; C → q; D → p
(d) A → p; B → q; C → s; D → r

10. Match the column I with column II and select the correct option.

	Column I		Column II
(A)	Li < Na < K < Rb	(p)	increasing order of ionisation energy
(B)	Li < Be > B < C	(q)	decreasing order of non-metallic nature
(C)	$F_2 > Cl_2 > Br_2 > I_2$	(r)	increasing order of size
(D)	O < S < F < Cl	(s)	increasing order of electron affinity

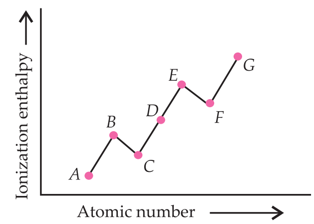
(a) (A) → (r); (B) → (p); (C) → (s); (D) → (q)
(b) (A) → (p); (B) → (q); (C) → (q); (D) → (p)
(c) (A) → (r); (B) → (p); (C) → (q); (D) → (s)
(d) (A) → (s); (B) → (r); (C) → (p); (D) → (q)

11. Match the column I with column II and select the correct option.

	Column I Compound (oxide)		Column II Significant properties
(A)	P_4O_{10}	(p)	Ionic
(B)	SiO_2	(q)	Covalent
(C)	Al_2O_3	(r)	Basic
(D)	MgO	(s)	Amphoteric

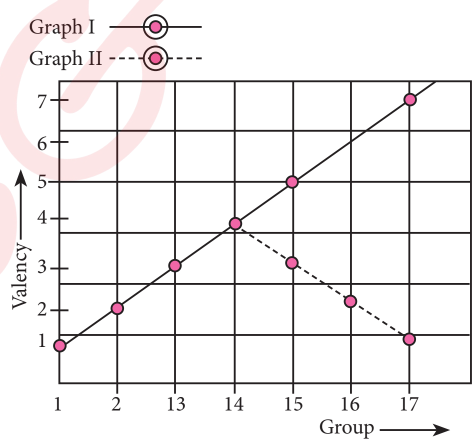
(a) (A) → (r); (B) → (p); (C) → (s); (D) → (q)
(b) (A) → (q); (B) → (q); (C) → (p, s); (D) → (p, r)
(c) (A) → (r); (B) → (p); (C) → (q); (D) → (s, p)
(d) (A) → (p); (B) → (r, s); (C) → (p); (D) → (q)

12. The ionization enthalpy of second period elements vary with atomic number as



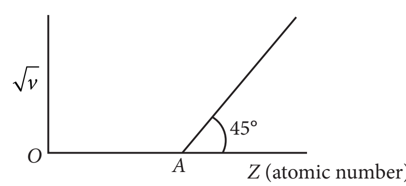
The elements present at points *B* and *E* are
(a) Be, C
(b) B, N
(c) Be, O
(d) Be, N

13. Variation of valency while moving along a period left to right is shown below



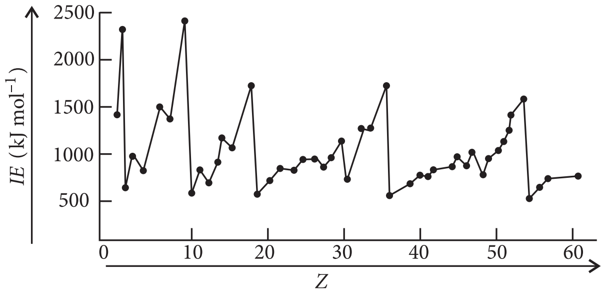
Select the correct statement(s).
(a) Graph I is for valency w.r.t. oxygen
(b) Graph II is for valency w.r.t. hydrogen
(c) Both (a) and (b)
(d) None of these

14. Modern periodic law (1913) was based on the *X*-ray spectra of Henry Moseley and equation.



$\sqrt{v} = a(Z - b)$ where, a and b are constants. If $OA = 1$, then element showing frequency of 225 Hz, then incorrect answer is
(a) *p*-block element
(b) *d*-block element
(c) non-metal
(d) of group 16 and 3rd period.

15. The graph of IE_1 versus atomic number (Z) is given below :



Which of the following statement is correct?

(a) Alkali metals are at the maxima and noble gases at the minima.

(b) Noble gases are at the maxima and alkali metals at the minima.

(c) Transition elements are at the maxima.

(d) Minima and maxima do not show any regular behaviour.

4 Assertion & Reason

Directions : In the following questions, a statement of assertion is followed by a statement of reason. Mark the correct choice as :

(a) If both assertion and reason are true and reason is the correct explanation of assertion.

(b) If both assertion and reason are true but reason is not the correct explanation of assertion.

(c) If assertion is true but reason is false.

(d) If both assertion and reason are false.

1. **Assertion :** Fe_2O_3 is more acidic than FeO.
Reason : Higher the oxidation state, higher the electronegativity, thus non-metallic characteristic is higher.

2. **Assertion :** Ionization potential across the period is $Na < Al < Mg < Si$.
Reason : Ionization potential decreases with decrease in atomic size.

3. **Assertion :** F atom has less negative electron gain enthalpy than Cl atom.
Reason : Additional electrons are repelled more effectively by $3p$-electrons in Cl than by $2p$-electrons in F atom.

4. **Assertion :** Electron gain enthalpy can be exothermic or endothermic.
Reason : Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form anion.

5. **Assertion :** The radii derived for noble gases are van der Waals' radii rather than covalent radii.
Reason : Noble gases are monoatomic.

6. **Assertion :** In the present form of periodic table, the period number corresponds to the highest principal quantum number of the elements in the period.
Reason : Elements having similar outer electronic configurations in their atoms belong to same period.

7. **Assertion :** Atomic number of the element copernicium is 112.
Reason : IUPAC name of this element is ununbium in which un - and bi-are used for 1 and 2 respectively in latin words.

8. **Assertion :** Boron can only form $[BF_4]^-$, whereas aluminium forms $[AlF_6]^{3-}$.
Reason : The first member of a group of elements in the s- and p- blocks shows anomalous behaviour.

9. **Assertion :** F^- ion is larger in size compared to F.
Reason : Electron repulsion increases because of addition of electron which results in decrease in effective nuclear charge.

10. **Assertion :** If the electron affinity of bromine is 3.48 eV, 8.02 kcal energy is released when 8 g of bromine is completely converted to Br^- ions in the gaseous state.
Reason : 8 g of bromine is equal to 0.1 mole of bromine.

5 Numerical / Integer Value Based

1. IE and EA values of an element are 13.0 eV and 3.8 eV respectively. Its electronegativity on Pauling's scale is ____.

2. From the given compounds : CaO, SO_2, SO_3, Fe_2O_3, Cl_2O_7, CO_2, Na_2O. If X is the number of compounds are acidic in water. The value of X is ______.

3. The electronic configuration of the element is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$. What is the atomic number of element which is just below the given element in the periodic table?

4. How many of the following properties show in general an increasing trend when moving down in a group in the modern periodic table?
Z_{eff}, atomic radius, ionization energy, electron affinity, electronegativity, acidic nature of oxides, acidic nature of hydracids, basic nature of hydroxides.

5. How many pairs are, in which first species has lower ionisation energy than second species?
(i) N and O (ii) Br and K

(iii) Be and B (iv) I and I^-
(v) Li and Li^+ (vi) O and S
(vii) Ba and Sr

6. The ionisation enthalpy of lithium is 521.7 kJ/mol. The amount of energy required to convert 80.5 mg of Li in gaseous state into Li^+ ion is ________ kJ.

7. How many ions are isoelectronic with argon.

Ba^{2+}, Sr^{2+}, K^+, S^{2-}, O^{2-}, F^-, Cl^-, Sc^{3+}, Mg^{2+}, Ca^{2+}, Li^+

8. If all the elements of period seven of periodic table have been discovered then atomic number of the alkaline earth metal of the 8th period would be _______.

9. The successive ionization energy values for an element X are given below.
1st ionisation energy = 410 kJ mol^{-1}
2nd ionisation energy = 820 kJ mol^{-1}
3rd ionisation energy = 1100 kJ mol^{-1}
4th ionisation energy = 1500 kJ mol^{-1}
5th ionisation energy = 3200 kJ mol^{-1}
The number of valence electrons for the atom X is

10. $I.E._1$ and $I.E._2$ of Mg are 178 and 348 kcal mol^{-1} respectively. The energy required for the reaction : $Mg \longrightarrow Mg^{2+} + 2e^-$ is ________ kcal mol^{-1}.

6 Exam Archive

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

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

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

4. The increasing order of the first ionisation enthalpies of the elements B, P, S and F (lowest first) is
(a) F < S < P < B (b) P < S < B < F
(c) B < P < S < F (d) B < S < P < F ***(AIEEE)***

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

6. Which one of the following sets of ions represents the collection of isoelectronic species?
(a) $K^+, Ca^{2+}, Sc^{3+}, Cl^-$ (b) $Na^+, Ca^{2+}, Sc^{3+}, F^-$
(c) $K^+, Cl^-, Mg^{2+}, Sc^{3+}$ (d) $Na^+, Mg^{2+}, Al^{3+}, Cl^-$
(Atomic nos. : F = 9, Cl = 17, Na = 11, Mg = 12, Al = 13, K = 19, Ca = 20, Sc = 21) ***(AIEEE)***

7. Which of the following order is wrong?
(a) $NH_3 < PH_3 < AsH_3$ – acidic
(b) Li < Be < B < C – 1st IP
(c) $Al_2O_3 < MgO < Na_2O < K_2O$ – basic
(d) $Li^+ < Na^+ < K^+ < Cs^+$ – ionic radius. ***(AIPMT)***

8. The following statements concern elements in the periodic table. Which of the following is true?
(a) For group 15 elements, the stability of +5 oxidation state increases down the group.
(b) Elements of group 16 have lower ionization enthalpy values compared to those of group 15 in the corresponding periods.
(c) The group 13 elements are all metals.
(d) All the elements in group 17 are gases. ***(JEE Main)***

9. An atom has electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3\ 4s^2$, you will place it in
(a) fifth group (b) fifteenth group
(c) second group (d) third group. ***(AIPMT)***

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

11. Ionic radii are
(a) inversely proportional to effective nuclear charge
(b) inversely proportional to square of effective nuclear charge
(c) directly proportional to effective nuclear charge
(d) directly proportional to square of effective nuclear charge. ***(AIPMT)***

12. In the periodic table, with the increase in atomic number, the metallic character of an element
(a) decreases in a period and increases in a group
(b) increases in a period and decreases in a group
(c) increases both in a period and the group
(d) decreases in a period and the group. *(AIPMT)*

13. Which one of the following sets of ions represents a collection of isoelectronic species?
(a) K^+, Cl^-, Ca^{2+}, Sc^{3+} (b) Ba^{2+}, Sr^{2+}, K^+, S^{2-}
(c) N^{3-}, O^{2-}, F^-, S^{2-} (d) Li^+, Na^+, Mg^{2+}, Ca^{2+}
(AIEEE)

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

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

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

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

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

19. The ionic radii (in Å) of N^{3-}, O^{2-} and F^- are respectively
(a) 1.71, 1.40 and 1.36 (b) 1.71, 1.36 and 1.40
(c) 1.36, 1.40 and 1.71 (d) 1.36, 1.71 and 1.40
(JEE Main)

20. The first ionization potentials (eV) of Be and B respectively are
(a) 8.29, 8.29 (b) 9.32, 9.32
(c) 8.29, 9.32 (d) 9.32, 8.29 *(AIPMT)*

ANSWER KEYS

Practice Time 1

1. (c)	2. (a)	3. (d)	4. (c)	5. (b)	6. (a)	7. (a)	8. (b)	9. (c)	10. (c)
11. (a)	12. (d)	13. (b)	14. (d)	15. (c)					

Practice Time 2

16. (d)	17. (c)	18. (c)	19. (c)	20. (a)	21. (a)	22. (c)	23. (b)	24. (d)	25. (b)
26. (a)	27. (b)	28. (a)	29. (c)	30. (a)					

1 Topicwise Warm-up

1. (b)	2. (d)	3. (d)	4. (a)	5. (c)	6. (b)	7. (b)	8. (c)	9. (b)	10. (a)
11. (b)	12. (b)	13. (c)	14. (b)	15. (b)	16. (c)	17. (a)	18. (d)	19. (a)	20. (c)
21. (a)	22. (a)	23. (a)	24. (d)	25. (c)	26. (d)	27. (d)	28. (a)	29. (b)	30. (b)
31. (a)	32. (c)	33. (b)	34. (b)	35. (d)	36. (d)	37. (b)	38. (c)	39. (b)	40. (c)
41. (b)	42. (a)	43. (c)	44. (c)	45. (c)	46. (d)	47. (b)	48. (d)	49. (a)	50. (a)
51. (a)	52. (c)	53. (c)	54. (c)	55. (a)	56. (c)	57. (a)	58. (a)	59. (d)	60. (a)
61. (a)	62. (d)	63. (b)	64. (c)	65. (c)	66. (d)	67. (d)	68. (d)	69. (d)	70. (b)
71. (b)	72. (c)	73. (c)	74. (c)	75. (d)	76. (a)	77. (c)	78. (d)	79. (d)	80. (d)
81. (c)	82. (b)	83. (d)	84. (b)	85. (c)	86. (c)	87. (c)	88. (c)	89. (d)	90. (a)
91. (d)	92. (b)	93. (a)	94. (a)	95. (b)	96. (a)	97. (a)	98. (b)	99. (c)	100. (a)

2 Check Your Understanding

1. (b)	2. (d)	3. (a)	4. (b)	5. (d)	6. (b)	7. (d)	8. (b)	9. (a)	10. (a)
11. (a)	12. (d)	13. (b)	14. (b)	15. (d)	16. (b)	17. (b)	18. (b)	19. (a)	20. (a)
21. (d)	22. (a)	23. (b)	24. (a)	25. (d)	26. (a)	27. (b)	28. (d)	29. (c)	30. (c)
31. (c)	32. (a)	33. (b)	34. (b)	35. (d)	36. (c)	37. (d)	38. (d)	39. (b)	40. (a)
41. (b)	42. (c)	43. (a)	44. (c)	45. (c)	46. (b)	47. (a)	48. (a)	49. (a)	50. (c)

3 Mixed Category

1. (c)	2. (d)	3. (d)	4. (d)	5. (d)	6. (b)	7. (c)	8. (a)	9. (c)	10. (c)
11. (b)	12 (d)	13. (c)	14. (b)	15. (b)					

4 Assertion & Reason

1. (a)	2. (c)	3. (c)	4. (b)	5. (b)	6. (c)	7. (b)	8. (b)	9. (a)	10. (b)

5 Numerical / Integer Value Based

1. (3)	2. (4)	3. (35)	4. (3)	5. (2)	6. (6)	7. (5)	8. (120)	9. (4)
10. (526)								

6 Exam Archive

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

3 Classification of Elements and Periodicity in Properties

Practice Time 1

1. (c) : $[Ne]3s^2 3p^5$ belongs to *p*-block element. Remaining electronic configurations of elements belong to the *s*-block element.

2. (a) : Only one element chromium from group 6 forms hydride, (CrH). The other elements do not form hydride due to low affinity towards hydrogen.

3. (d) : $3s^2\ 3p^2$: (Si); Metalloid

4. (c) : The period number in the long form of the periodic table is equal to maximum principal quantum number of any element of the period.

5. (b) : Lothar Meyer plotted the physical properties such as atomic volume, melting point and boiling point against atomic weight and obtained a periodically repeated pattern.

6. (a) : Eka- aluminium and Eka-silicon are known as gallium and germanium.

7. (a) : Among the given elements, Newlands' law of octave applies to Be, Mg, Ca set of elements.

8. (b) : $[Xe]\ 6s^1$: Cs is the most electropositive element.

9. (c) : The elements in which electrons are progressively filled in 4*f*-orbital are called lanthanoids.

10. (c) : The element of the 17^{th} group (Halogens) of the periodic table is likely to be strongly non-metallic.

11. (a)

12. (d) : Si and Ge.

13. (b) : Set of elements follows Dobereiner's triad is Li, Na, K.

14. (d) : The elements of group 1, 2,13, 14, 15, 16, 17, 18 are collectively called representative elements.

15. (c) : Among the given elements, atomic mass of Be was changed by Mendeleev. He corrected the atomic mass of beryllium from 13.5 to 9 by using the formula, Atomic weight = Equivalent weight × valency.

Practice Time 2

16. (d) : Neutron/proton ratio does not affect the periodicity of the elements.

17. (c) : $Na \longrightarrow Na^+ + e^-\ ;\ IE = 5.1$ eV

$Na^+ + e^- \longrightarrow Na$

Second reaction is inverse of first reaction, therefore electron gain enthalpy of Na^+ is reverse of ionisation energy.

$\therefore\ \Delta H_{eg} = -5.1$ eV

18. (c) : In general ionisation energy increases as we move from left to right in a period. It is due to the increase in effective nuclear charge. *I.E.* of Be and N is high due to stable configuration. Hence, the correct order is as follows F > N > C > Be >B.

19. (c) : From the given options as we notice that the greatest energy will be involved for option (c) as it has attained the noble gas configuration after the loss of 2 electrons and removal of a 3^{rd} electron will require a lot of energy.

20. (a) : Nitrogen compounds encompass oxidation states of nitrogen ranging from – 3 (as in ammonia and amines) to +5 (as in nitric acid).

21. (a) : Electronegativity has no unit.

22. (c) : Among the given elements, H is the most electronegative element.

23. (b) : The process, $O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)}$, involves absorption of energy. When a neutral atom gains an electron, energy is released. When an anion (having unit negative charge) gains an electron, energy is absorbed. This is due to the repulsion between the additional electron and negatively charged anion.

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

25. (b) : Alkali metals have lowest *IE*, values in their respective period. Also there is a large jump between IE_1, and IE_2 due to the removal of the second electron from the noble gas core. Therefore, out of the given IE_1 and IE_2 values, (II) belongs to alkali metals.

26. (a) : Electrons of 6*s* subshell do not participate in bonding due to inert pair effect.

27. (b) : For *s*-block elements, in general the electronegativity, ionisation energy and electron affinity has lower value, while atomic radius decreases along the period. So, *s*-block element will have high atomic radius.

28. (a) : The ionisation energy of a multivalent atom increases with a consecutive removal of an electron. This is due to an increase in effective nuclear charge on the valence electron that makes it difficult to ionize a cation. Hence, $IE_{III} > IE_{II} > IE_I$.

29. (c) : O—H has maximum electronegativity difference, so it is the most polar bond.

30. (a) : Valence electrons remain unchanged on descending a group in the periodic table.

1 Topicwise Warm-up

1. (b) : He described these elements as Eka-aluminium and Eka-silicon, but the elements were actually Ga and Ge.

2. (d) : The discovery of these elements were made later on. However, Mendeleev left blank spaces in periodic table and predicted their properties.

3. (d) : Mass number of B is an average of A and C

$$m_B = 23 = \frac{m_A + m_C}{2} = \frac{m_A + 29}{2}$$

$$m_A = 46 - 29 = 17$$

4. (a)

5. (c) : According to Mendeleev periodic table,
Eka silicon = Germanium and Eka aluminium = Gallium

6. (b) : Group VIII consists of three triads such as Fe, Co, Ni; Ru, Rh, Pd and Os, Ir, Pt arranged in 4^{th}, 5^{th} and 6^{th} periods respectively.

7. (b) : Lother Meyer calculated atomic volumes and plotted them against atomic masses. Halogens occupied positions on the ascending positions of the curve before noble gases.

8. (c) **9. (b)**

10. (a) : Magic no $(2n^2)$: 2, 8, 8, 18, 18, 32, 32, 50, 50, 72, 72
Period : 1, 2, 3, 4, 5, 6, 7, 8, 9, 10
Element : 2 + 8 + 8 + 18 + 18 + 32 + 32 + 50 + 50 + 72 = 290

11. (b) **12. (b)**

13. (c) : Moseley gave the relation, $\sqrt{v} = a(z - b)$
or, $v = a^2(z - b)^2$

14. (b) **15. (b)**

16. (c) : p-orbital is being filled, thus it is a p-block element.
[Ar] $3d^{10}\ 4s^2\ 4p^3$ represent as, it is a metalloid.

17. (a)

18. (d) : Cl = $3s^2\ 3p^5$, Cl^- = $3s^2\ 3p^6$, Ar = $3s^2\ 3p^6$

19. (a)

20. (c) : $5s^2\ 5p^4$
Valence electrons = 6
Thus, group = 10 + 6 = 16
n = 5. Thus, period = 5

21. (a) : $4d^{10}5s^0$ is the exceptional configuration of Pd. Its electronic configuration should be $[Kr]^{36}4d^8, 5s^2$. Thus, its
Period = 5^{th}
Group = $ns + (n - 1)d$ electrons = 2 + 8 = 10

22. (a) : Element : $1s^22s^22p^63s^23p^63d^84s^2$ (n = 4)
Period : As n = 4, the element belong to 4^{th} period
Group : $ns + (n - 1)d + np$ = 2 + 8 + 0 = 10

23. (a) **24. (d)** **25. (c)**

26. (d) :

Period	Suborbit	Orbitals	Elements	Total
3	$3s$ $3p$ $3d$	1 3 5	3 9 15	27
4	$4s$ $4p$ $4d$ $4f$	1 3 5 7	3 9 15 21	48
5	$5s$ $5p$ $5d$ $5f$ $5g$	1 3 5 7 9	3 9 15 21 27	75

27. (d) **28. (a)** **29. (b)**

30. (b) : Group – 14

C	Si	Ge	Sn	Pb	Fl
Non-metal	Metalloid	Metalloid	Metal	Metal	Metal

31. (a) : Group is based on number of electrons in metal atom (*i.e.*, number of protons)

Sc^{3+}[Ne] $3s^23p^6$

Sc : [Ne] $3s^23p^64s^23d^1$

Valence electrons = 3

Group = 3 (d-block element)

32. (c) **33. (b)**

34. (b) : In accordance with, *Aufbau principle*, the most stable electronic configuration is (b) and its group number is 7.

35. (d)

36. (d) : Since last electron enter in d-orbital, it is a transition element.

37. (b) : Since n = 2 it belongs to 2^{nd} period.

38. (c) : ns^2np^4 represent 16 group *i.e.* O, S, Se, Te and Po.

39. (b) : It contains maximum number of electrons in the outermost shell. An element with more than 3 electrons in outermost shell is non-metallic.

40. (c) : I. It is in first excited state. It represents element of atomic number 20 (Ca). Its ground state electronic configuration is [Ar] $4s^2$. Thus, s-block element.
II. It is in first excited state. It represents element of atomic number 10 (Ne). Its ground state electronic configuration is $1s^2\ 2s^2\ 2p^6$. Thus, p-block element.
III. It represents ground state of element having atomic number 29 (Cu). Due to more stability of paired electrons, $4s$ electron goes to $3d$. Thus d-block element.
IV. It represents group state of element with atomic number 71(Lu) (lanthanide). Thus, f-block element.

41. (b) : The second electron gain enthalpy is always positive.

$O \rightarrow O^-$; $\Delta H_{eg} = -141$ kJ/mol

$O \rightarrow O^{2-}$; $\Delta H_{eg} = 780$ kJ/mol

42. (a)

43. (c) : As maximum valency equals to group number.

44. (c) : C–C bond length = 1.54 Å

$$r_C = \frac{1.54}{2} = 0.77\ \text{Å}$$

$$r_X = \frac{1.00}{2} = 0.50\ \text{Å}$$

C–X bond length

$d_{C-X} = r_C + r_X - 0.09|\chi_X - \chi_C|$

$= 0.77 + 0.50 - 0.09|3 - 2| = 0.77 + 0.50 - 0.09 \times 1$

$= 1.27 - 0.09 = 1.18$ Å

Thus, C–X bond length is 1.18 Å.

45. (c) : Increases one by one from IA to IVA and then decreases from VA to VIIA one by one.

46. (d)

47. (b) : There is more repulsion for the incoming electron when the size of atom is smaller.

48. (d) : Z_1 = Second ionization potential and Z_2 = Third ionisation potential. Second ionisation potential is always less than the third ionisation potential.

49. (a)

50. (a) : When electron is brought near O^- there will be repulsion between them, and therefore the energy will be positive *i.e.*, there will be absorption of energy during the process.

51. (a) : Here Bi_2O_3 is the most basic oxide SeO_2 is acidic oxide while rest are amphoteric oxides.

52. (c) : N_2O

Total electrons ⇒ 14 + 8 = 22

CO_2

Total electrons ⇒ 6 + 16 = 22

53. (c) **54. (c)**

55. (a) : The Z/e for N^{3-}, O^{2-} and F^- is $\frac{7}{10}, \frac{8}{10}, \frac{9}{10}$ so the order of ionic radii is $N^{3-} > O^{2-} > F^-$. Hence, the radius of O^{2-} is in between 1.76 Å and 1.36 Å.

56. (c) : *E.N.* stands for electronegativity, *E.A.* stands for electron affinity and *I.P.* stands for Ionization potential.

It is observed that for an element, *E.A.*, *E.N.* and *I.P.* usually vary in the same direction. Hence, when *E.A.* and *E.N.* increase the *I.P.* also increases. *E.N.* has the mean value of *I.P.* and *E.A.*

$$E.N. = \frac{I.P. + E.A.}{2}$$

$\therefore\ 2E.N. = I.P. + E.A.$ or $2E.N. - I.P. - E.A. = 0$

57. (a)

58. (a) : Energy needed to convert 1 mole of $Na_{(g)}$ to Na^+ ions = 495 kJ

∴ Energy needed to convert 3 mole of $Na_{(g)}$ to Na^+ ions = 495 × 3 = 1485 kJ

59. (d)

60. (a) : In $M^- \longrightarrow M + e^-$, an electron is removed and for it energy is required. This makes it an endothermic process, therefore, enthalpy is positive (X).

61. (a) : H–O–X

∵ O–H bond will break

∴ basic

$\underbrace{2.1 \quad 3.5}_{1.4}\underbrace{\quad 0.8}_{2.7}$

More electronegativity difference

∵ O—X bond would break first

Hence, it is basic in nature.

62. (d) : Element of second and third period.

Diagonal relationship

II	Li	Be	B	C	N	O	F
III	Na	Mg	Al	Si	P	S	Cl

63. (b) : Be ($1s^2\,2s^2$) due to its fully filled 2s-sub shell has least tendency to take up an electron. As such Be^- is least stable.

64. (c) : Cr = $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1\,3d^5$

Z_{eff} = Atomic no – σ

σ for 4s electron = (valence shell e^- –1) × 0.35 + Penultimate shell e^- × 0.85 + remaining e^-

= (1 – 1) × 0.35 + 13 × 0.85 + 10

= 0 + 11.05 + 10 = 21.05

Z_{eff} = 24 – 21.05 = 2.95

Z_{eff} for 3d electron :

σ for 3d e^- = (valence shell e^- –1) × 0.35 + remaining e^-

= (5 – 1) × 0.35 + 18 = 1.4 + 18 = 19.4

= 24 – 19.4 = 4.6

For 3d e^- of Cr^{3+}

σ = (3 – 1) × 0.35 + 18 = 18.70

∴ Z_{eff} = 24 – 18.7 = 5.3

65. (c)

66. (d) : $2\dot{A}: + 3\ddot{\underset{\cdot\cdot}{B}}:$

$(A^{3+})_2(:\ddot{\underset{\cdot\cdot}{B}}:^{2-})_3 \equiv A_2B_3$

67. (d) : Consider the case of alkali metals,

$_3Li = 1s^2\,2s^1$ (No d-orbital)

Size ↓ (EN) ↑ (IP) ↑

$Na_{11} = 1s^2\,2s^2\,2p^6\,3s^1$

K_{19} = [Ar] $4s^1$ (Vacant d-orbitals present)

Rb_{37} = [Ar] $4s^2\,3d^{10}\,4p^6\,5s^1$

Li lacks d-orbital.

Li is smallest in size.

Li has maximum (EN) and(IP).

If differs from other alkali metals.

68. (d) : Chloride ion and potassium ion are iso-electronic, *iso*-electronic ions are 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

69. (d) : Ca^{2+} = 2, 8, 8

70. (b) : Solubility of hydroxides of group II increases down the group.

71. (b) $\quad$ **72. (c)**

73. (c) : Since the fourth *I.E.* is very high, thus, electron is to be removed from stable configuration.

74. (c) $\quad$ **75. (d)**

76. (a) : Elements on the left side have lowest ionisation enthalpy due to which they can very easily lose electrons while the elements on the right can accept electrons easily as they show highest negative electron gain enthalpy.

77. (c) : Atomic radius decreases in a period from left to right and increases in a group from top to bottom.

78. (d) : All the three ions, *i.e.*, He^+ $(1s^1)$, Li^{2+} $(1s^1)$, Be^{3+} $(1s^1)$ have no electrons in the inner shells and hence screening effect is not observed.

79. (d)

80. (d) : Electronegativity decreases down the group.

81. (c) $\quad$ **82. (b)** $\quad$ **83. (d)**

84. (b) : $(IE)_1 < (IE)_2 << (IE)_3 < (IE)_4$

$(IE)_3$ is very high thus, M^{2+} is stable with inert gas configuration. Thus, element belongs to alkaline earth metals.

85. (c)

86. (c) : It is difficult to remove an electron from a highly electronegative element.

87. (c) : $r_C + r_F = 1.33$ Å ...(i)

$r_{Si} + r_F = 1.54$ Å ...(ii)

$r_C + r_{Si} + 2r_F = 2.87$Å ...(ii)

$r_C + r_{Si} = 1.87$ Å (given)

1.87 Å $+ 2r_F = 2.87$ Å

$2r_F = 1.00$ Å

$r_F = 0.5$ Å

88. (c) : Screening effect order is $s > p > d > f$.

89. (d) : Since, in (d) it is IE_3 and $IE_3 >> IE_2 > IE_1$

90. (a)

91. (d) :

$$\left.\begin{array}{l} K^+, (Z=19), \text{no. of } e^-\text{s} = 19-1 = 18 \\ Ca^{2+}, (Z=20), \text{no. of } e^-\text{s} = 20-2 = 18 \\ Ne, (Z=10), \text{no. of } e^-\text{s} = 10 \end{array}\right] \begin{array}{l}\text{Hence, not} \\ \text{iso electronic}\end{array}$$

92. (b) : $O_{(g)} + e^- \xrightarrow[-EA_1]{} O^-_{(g)} + e^- \xrightarrow[-EA_2]{} O^{2-}_{(g)}$

Energy is required to add an electron to the negatively charged species due to electron-electron repulsion.

93. (a) : Ca : [Ne] $3s^2\,3p^6\,4s^2$

Ca^{2+} : [Ne] $3s^2\,3p^6\,4s^0$ (Stable noble gas configuration)

Hence, there would be sudden increase in IE_3 value.

94. (a) : P and Si belong to group 4 and group 5 respectively and both of them are in the 3rd period.

Electronegativity increases along the period.

$\therefore$ *EN* of P(2.1) > Si(1.8)

95. (b)

96. (a) : Trend of *EP* (electropositive) character is reverse that of *EN* character of elements.

Thus, *EP* character increases down the group (↓) but decreases along the period (→).

97. (a) : Applying Fajans' rule, lower the charge; the more is ionic and the more is basic and vice versa.

Moreover, small cation, high charge and presence of *d*-electrons mean more covalent and more acidic.

Therefore, decreasing order of basic nature is

	$\overset{+1}{Na_2O}$	>	$\overset{+2}{MgO}$	>	$\overset{+3}{Al_2O_3}$	>	$\overset{+2}{CuO}$
Group ⇒	1		2		13		11
Period ⇒	3rd		3rd		3rd		4th

Size of $Al^{3+} > Cu^{2+}$ and presence of *d*-electron in Cu^{2+} makes it less basic (or more acidic) than Al_2O_3.

98. (b) : (i) $Mg_{(g)} \longrightarrow Mg^+_{(g)} + e^-$, $\quad IE_1 = 178$ kcal mol^{-1}

(ii) $Mg^+_{(g)} \longrightarrow Mg^{2+}_{(g)} + e^-$, $\quad IE_2 = 348$ kcal mol^{-1}

Adding (i) and (ii), we have

$Mg_{(g)} \longrightarrow Mg^{2+}_{(g)} + 2e^-$

$\therefore$ $E = IE_1 + IE_2 = 178 + 348 = 526$ kcal

99. (c) : Due to stable full-filled orbitals of noble gases, it is very-very difficult to add electron. Hence, their *EA*s is almost zero.

100. (a) : *IE* and *EA* of an element are equal in magnitude but opposite in sign.

$\therefore$ $IE = -EA$

2 Check Your Understanding

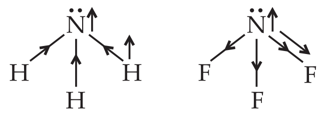
1. (b) : *P* : [Ar] $3d^{10}4s^1$ = Cu

R : [Ar] $3d^{10}4s^24p^64d^{10}5s^1$; [Kr] $4d^{10}5s^1$: Ag

2. (d) : The *d*-block has 10 columns corresponding to d^1 to d^{10} configurations.

3. (a) $\quad$ **4. (b)** $\quad$ **5. (d)**

6. (b) : On moving down the group, the stability of –3 oxidation state decreases. This is due to the following reasons (i) On descending a group the size of the atom or ion increases. As a result, attraction of the nucleus per newly added electron decreases (ii) A large anion cannot fit easily into lattice of a small cation. (iii) As the negative charge on the ion increases, it becomes more and more susceptible to polarisation.

7. (d) : 

8. (b)

9. (a) : Penetration effect order is $s > p > d > f$.

10. (a) **11. (a)** **12. (d)**

13. (b) : The removal of an electron from Na^+, *i.e.*, from an [Ne] core, is more difficult than from Mg^+, *i.e.*, from [Ne] $3s^1$.

14. (b) :

Atom	Electronegativity in Pauling scale
F	4.0
O	3.5
N	3.0

Thus, NH_3, H_2O and HF can form extensive intermolecular H-bonding.

15. (d) : HF is not a stronger acid than HCl because fluorine is more electronegative than chlorine, therefore it does not donate hydrogen easily as is done in the case of HCl.

16. (b) **17. (b)**

18. (b) : The element belongs to group 2, and so it forms MO and MCl_2.

19. (a) : Generally as we move from left to right in a period, there is regular decrease in atomic radii and in a group as the atomic number increases the atomic radii also increase. Thus the atomic radius of Sn should be less than lanthanides. La > Sn. But due to lanthanide contraction, in case of lanthanides there is a continuous decrease in size with increase in atomic number. Hence the atomic radius follow the given trend :

Ce > Sn > Yb > Lu.

20. (a) : Ionization energy decreases with the increase in number of orbitals or down the group.

21. (d)

22. (a) : Na^+ and F^- are isoelectronic ions. Since amongst isoelectronic ions, the size of the anion is larger than that of the cation, therefore, $Na^+ < F^-$.

23. (b) : The electropositive character increases down a group and so does the basic nature of the oxide.

24. (a) : Na and Mg is not an exception because there is no half-filled or completely filled orbital in them.

25. (d) : Large jump between $I.E._3$ and $I.E._4$ suggests that the element has three valence electrons.

26. (a) **27. (b)** **28. (d)**

29. (c) : $K \longrightarrow K^+ + e^-$ $\quad \Delta E_1 = 4.3$ eV

$F + e^- \longrightarrow F^-$ $\quad \Delta E_2 = -E$ eV

$$\frac{19.0}{23.06} = \Delta E_1 + \Delta E_2 = 4.3 - E$$

$0.82 = 4.3 - E$

$E = 3.48$

30. (c) **31. (c)**

32. (a) : Stable configuration of X (due to half-filled orbitals) than that of Z results greater ionisation potential.

$\therefore \quad I_x > I_z$

33. (b)

34. (b) : $X(g) \longrightarrow X^+(g) + e^-$

If I is ionisation energy then

$$\therefore \quad \frac{N_0}{2}(I) = E_1 \quad \therefore \quad I = \frac{2E_1}{N_0}$$

If E is electron-affinity then

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

$$\frac{N_0}{2}(E) = E_2 \quad \therefore \quad E = \frac{2E_2}{N_0}$$

35. (d) : It is because of electronic repulsion.

36. (c) : Alkali metals and coinage elements, both have valence electron one, elements of same group I (A and B). Alkali metals are highly reactive while coinage elements are very less reactive.

37. (d) : Metallic character increases down the group (↓) and decreases along the period (⟶).

Therefore, metallic character order is

K > Mg > Al > B

38. (d)

39. (b) : It does not help to predict the stable valency states of the elements.

40. (a)

41. (b) : In the compound M—O—H, if *IE* or *E.N.* of M is low, the compound will act as a base and if the *IE* or *E.N.* of M is high, the compound will behave as an acid. Therefore, M—O—H will act as an acid as *E.N.* of M is high (3.5) and M'—O—H will act as a base as *E.N.* of M' is low (1.72).

42. (c)

43. (a) : $IE + EA = 275 + 86 = 361$ kcal mol^{-1}

$$\therefore \quad \text{Electronegativity} = \frac{IE + EA}{2 \times 62.5} = \frac{361}{125} = 2.8$$

44. (c) : Non-metallic character decreases down the group (↓) and increases along the period (⟶).

Therefore, non-metallic character order is

F > N > C > B > Si.

45. (c) **46. (b)**

47. (a) : Electronegativity values of given elements are as follows :

Be – 1.5 (I) Mg – 1.2 (IV)

O – 3.5 (II) N – 3.0 (III)

i.e., II > III > I > IV

48. (a) : *EN* order of period 2 elements is C < N < O < F.

49. (a) : As we move across the period, atomic size decreases and electronegativity increases and acidic strength increases.

Hence, the acidic strength order is $Cl_2O_7 > SO_2 > P_4O_{10}$.

50. (c)

3 Mixed Category

1. (c) **2. (d)**

3. (d) : Elements (58 - 71) are called lanthanides and elements (90 - 103) are called actinides.

4. (d)

5. (d) : I. H resembles alkali metals as it forms monovalent cation as Na^+. H resembles halogen family as it forms H^-. Thus, its position is not defined. Thus, correct.

II. Elements (58 - 71) called lanthanides and (90 - 103) called actinides do not find proper place in the table.

Thus, correct.

III. Members of zero group (noble gases) were not discovered at the time of Mendeleev hence, their position was not given. Thus, correct.

IV. Isotopes have same atomic number but different atomic mass. Since, classification was based on atomic mass, isotopes would be placed at different positions. But they were placed at the same position. Thus, correct.

6. (b) **7. (c)** **8. (a)** **9. (c)**

10. (c) **11. (b)**

12. (d) : The electronic configuration of Be = $1s^2, 2s^2$ (Fully-filled).

The electronic configuration of N = $1s^2, 2s^2, 2p^3$ (Half-filled).

Due to stable electronic configuration of Be and N the *IE* of these elements is more than elements of the next group.

13. (c) : As given in graph I, valency increases from 1 to 7 while moving along a group in *s* and *p*-block. It is w.r.t. oxygen. From Graph II, valency increases from 1 to 4 and then decreases to 1. It is w.r.t. hydrogen.

	s-block		*p*-block				
Group	1	2	13	14	15	16	17
Elements	Na	Mg	Al	Si	P	S	Cl
Compounds with oxygen	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Valency w.r.t. oxygen	1	2	3	4	5	6	7
Compounds with hydrogen	NaH	MgH_2	AlH_3	SiH_4	PH_3	H_2S	HCl
Valency w.r.t. hydrogen	1	2	3	4	3	2	1

14. (b) : $\sqrt{v} = aZ - ab$

Thus, represents a straight line, $y = mx - c$

When $\sqrt{v} = 0$, $aZ = ab$

$Z = b = OA = 1$

When $Z = 0$, $\sqrt{v} = -ab = -a$ (since, $b = 1$)

$= \tan 45°$ (slope) $= 1$

Thus, $a = 1 = Z - 1$

$\sqrt{225} = Z - 1 \Rightarrow Z = 16$

EC of atomic number 16 = $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^4$

Thus, it is *p*-block element thus, option (a) is correct.

Group = 16, valence electrons = 6

Period = 3, thus, option (d) is correct

$(P + 1) = 4 > 6$ (valence electrons)

It is non-metal, hence option (c) is correct, and option (b) is incorrect.

15. (b) : Noble gases are at the maxima which have closed electron shells and very stable electronic configuration. On the other hand, minima occurs at the alkali metals and their low IE_1 can be correlated with their high reactivity.

4 Assertion & Reason

1. (a)

2. (c) : Ionization potential decreases with increase in atomic size and also for a given shell, *I.E.* is in order $s > p > d > f$.

3. (c) : Additional electron is repelled more effectively by 2*p*-electrons in F than 3*p*-electrons in Cl.

4. (b) : Depending on the element, the process of adding an electron to the atom can be either endothermic or exothermic.

5. (b)

6. (c) : Elements having similar outer electronic configurations in their atoms are arranged in vertical columns, called groups or families.

7. (b) : IUPAC name of element copernicium having atomic number 112 is ununbium, for 1, suffix 'un' and for 2 suffix 'bi' is used which are latin words.

8. (b) : The maximum covalency of first member of each group of *s*- and *p*-block elements is 4, (due to non-availability of *d*-orbitals) whereas the other members can expand their valence shell to accommodate more than four pairs of electrons (due to availability of *d*-orbitals).

9. (a)

10. (b) : No. of moles of bromine = $\frac{8}{80} = 0.1$

$\therefore$ Energy required = $0.1 \times 3.48 \times 23.06 = 8.02$ kcal

5 Numerical / Integer Value Based

1. (3) : On Mulliken scale, $\text{E.N.} = \frac{13+13.8}{2} = 8.4$

Mulliken values are nearly 2.8 times larger than Pauling's values.

$\therefore$ On Pauling's scale, $\text{E.N.} = \frac{8.4}{2.8} = 3.$

2. (4)

3. (35) : $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$, $Z = 17$
Atomic number of element below = 17 + 18 = 35

4. (3)

5. (2) : $Li < Li^+$ and $Ba < Sr$.

6. (6) : Mass of Li = 80.5 mg = 80.5×10^{-3} g = 8.05×10^{-2} g

No. of moles of Li = $\dfrac{8.05\times10^{-2}}{7} = 0.0115$ moles

Energy required to convert 0.0115 moles atoms of Li into Li^+ is $521.7 \times 0.0115 = 6$ kJ.

7. (5) : $K^+ = 19 - 1 = 18e^-$
$S^{2-} = 16 + 2 = 18e^-$
$Cl^- = 17 + 1 = 18e^-$
$Ca^{2+} = 20 - 2 = 18e^-$
$Sc^{3+} = 21 - 3 = 18e^-$
These five ions are isoelectronic species.

8. (120) : Oganesson on with atomic number 118 is the last element of period seven of the periodic table. Therefore, the atomic number of next element of period eight of alkali metal will be 119 and of alkaline earth metal will be 120.

9. (4) : The ionisation energies provide an indication about the number of valence electrons in an atom.
For example, Na has $IE_2 >> IE_1$ because it has one valence electron,while Mg has $IE_3 >> IE_2 > IE_1$, because it has two valence electrons.
In this case X has $IE_5 >> IE_4 > IE_3 > IE_2 > IE_1$.
Therefore it has four valence electrons.

10. (526) : $Mg \longrightarrow Mg^+ + e^-$; $I.E._1 = 178$ kcal mol^{-1} ...(i)
$Mg^+ \longrightarrow Mg^{2+} + e^-$; $I.E._2 = 348$ kcal mol^{-1} ...(ii)
On adding (i) and (ii), we get
$Mg \longrightarrow Mg^{2+} + 2e^-$
$\therefore\quad E = I.E._1 + I.E._2 = 178 + 348 = 526$ kcal mol^{-1}

6 Exam Archive

1. (c) : This can be explained on the basis of z/e whereas $\left\{\dfrac{\text{nuclear charge}}{\text{no. of electrons}}\right\}$, z/e ratio increases, the size decreases and when z/e ratio decreases the size increases.

$$\text{For } Li^+, \frac{z}{e} = \frac{3}{2} = 1.5\ ;\ \text{For } B^{3+}, \frac{z}{e} = \frac{5}{2} = 2.5$$

$$\text{For } O^{2-}, \frac{z}{e} = \frac{8}{10} = 0.8\ ;\ \text{For } F^-, \frac{z}{e} = \frac{9}{10} = 0.9$$

Hence, O^{2-} has highest value of ionic radius.

2. (d) : Abnormally high difference between 2nd and 3rd ionisation energy means that the element has two valence electrons, which is a case in configuration (d).

3. (d) : All the alkali metals are highly reactive elements since they have a strong tendency to lose the single valence s-electron to form unipositive ions having inert gas configuration. This reactivity arises due to their low ionisation enthalpies and high negative values of their standard electrode potentials.
However, the reactivity of halogens decreases with increase in atomic number due to following reasons:
(i) As the size increases, the attraction for an additional electron by the nucleus becomes less.
(ii) Due to decrease in electronegativity from F to I, the bond between halogen and other elements becomes weaker and weaker.

4. (d) :

Element :	B	S	P	F
***I.E.* (eV) :**	8.3	10.4	11.0	17.4

In general as we move from left to right in a period, the ionisation enthalpy increases with increasing atomic number. The ionisation enthalpy decreases as we move down a group. P($1s^2\,2s^2\,2p^6\,3s^2\,3p^3$) has a stable half filled electronic configuration than S($1s^2\,2s^2\,2p^6\,3s^2\,3p^4$). For this reason, ionisation enthalpy of P is higher than S.

5. (a) : The correct order of increasing negative electron gain enthalpy is : I < Br < F < Cl due to electron-electron repulsion in small sized F atom.

6. (a) : Isoelectronic species are those which have same number of electrons.
$K^+ = 19 - 1 = 18$; $Ca^{2+} = 20 - 2 = 18$
$Sc^{3+} = 21 - 3 = 18$; $Cl^- = 17 + 1 = 18$
Thus all these ions have 18 electrons in them.

7. (b) : Li, Be, B, C - these elements belong to the same period. Generally the value of 1st ionisation potential increases on moving from left to right in a period, since the nuclear charge of the elements also increase in the same direction. But the ionisation potential of boron ($B \rightarrow 2s^2\,2p^1$) is lower than that of beryllium ($Be \rightarrow 2s^2$), since in case of boron, $2p^1$ electron has to be removed to get B^+ while in case of Be ($2s^2$), s-electron has to be removed to get Be^+ ($2s^1$). p-electron can be removed more easily than s-electron so the energy required to remove electron will be less in case of boron. The order will be
Li < B < Be < C.

8. (b) : Group 15 elements have stable half-filled (ns^2np^3) configurations hence, their ionization enthalpy is higher than that of group 16 elements.

9. (a) : The electronic configuration of an atom :
$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^3\,4s^2$
In the configuration, the last electron of the atom is filled in d-subshell as $3d^3$. Thus, this element belongs to d-block of the periodic table with group no. VB or 5.

10. (d) : Ionization enthalpy decreases from top to bottom in a group while it increases from left to right in a period.

11. (a)

12. (a) : Metallic character decreases in a period and increases in a group.

13. (a) : $K^+ = 19 - 1 = 18\,e^-$; $Cl^- = 17 + 1 = 18\,e^-$
$Ca^{2+} = 20 - 2 = 18\,e^-$; $Sc^{3+} = 21 - 3 = 18\,e^-$
Thus, all the species are isoelectronic.

14. (c) : Atomic number of the given element is 15 and it belongs to group 15. Therefore atomic number of the element below the above element $= 15 + 18 = 33$.

15. (a) : All the given species are isoelectronic. Among isoelectronic species, anions generally have greater size than cations.
Also greater, the nuclear charge (Z) of the ion, smaller the size.
Thus the order is : $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$

16. (c) : N, O and F are highly electronegative non-metals and will have the strongest tendency to form anions by gaining electrons from metal atoms.

17. (b) : As we move from left to right across a period, ionisation enthalpy increases with increasing atomic number. So the order of increasing ionisation enthalpy should be $B < C < N < O$.
But $N(1s^2\,2s^2\,2p^3)$ has a stable half filled electronic configuration.
So, ionization enthalpy of nitrogen is greater than oxygen.
So, the correct order of increasing the first ionization enthalpy is $B < C < O < N$.

18. (a) : As positive charge on the cation increases, effective nuclear charge increases. Thus, atomic size decreases.

19. (a) : The ionic radii of isoelectronic ions increase with the decrease in magnitude of the nuclear charge.

F^-	<	O^{2-}	<	N^{3-}
1.36 Å		1.40 Å		1.71 Å

20. (d) : ${}_4Be \rightarrow 1s^2\,2s^2$, ${}_5B \rightarrow 1s^2\,2s^2\,2p^1$
Due to stable fully-filled 's'-orbital arrangement of electrons in 'Be' atom, more energy is required to remove an electron from the valence shell than 'B'-atom. Therefore 'Be' has higher ionisation potential than 'B'.