DAY THREE

Classification of Elements and Periodicity in Properties

Learning & Revision for the Day

- Modern Periodic Law and Long Form of Periodic Table
- Classification of Elements
- Periodic Trends in Properties of Elements

Earlier Attempts to Classify the Elements

Many attempts were made to classify the known elements from time to time. These are :

- Prout's hypothesis
- Dobereiner's triad law
- Newland's octave law
- Lother Meyer's curve
- **Mendeleev's Periodic Law** (1869) The properties of elements are periodic functions of their atomic weights.

Modern Periodic Law and Long Form of Periodic Table

- **Modern periodic law** (given by Moseley) The physical and chemical properties of elements are periodic functions of their atomic numbers.
- The long form of periodic table was prepared by **Rang**, developed by **Warner** and extended by **Bury**. This table is also known as Bohr table, since it contains both the **Moseley** and modern periodic laws.
- Following are the characteristic features of long form of periodic table:
  - (i) The modern periodic table is divided into two main categories known as,
    - (a) vertical columns–groups and (b) horizontal rows–periods.
  - (ii) There are 18 groups. In modern periodic table, these groups are further classified as A-subgroups and B-subgroups.
  - (iii) Members of the same group have same valence shell configurations.

(iv) Elements of group 1A to VIIA (i.e. 1, 2, 13, 14, 15, 16, 17, groups) are known as **representative elements**. Members of group 1B to VIII B. (i.e. 3, 4, 5, 6, 7, 8, 9, 10, 11 and 12) are known as **transition elements**.

(v) Elements of group-18 (also known as zero group) are **noble gases.**

(vi) There are seven horizontal rows in the periodic table, known as **periods.**

s-Block Elements

p-Block Elements

d-Block Elements

Group → New Notation: 1 … 17, 18

CAS Version: IA … VIIA, 0 (zero)

STATE

Gas G

Liquid L

Solid S

Not found in nature X

Key to Chart

G 8 – Atomic number

O – Symbol

Oxygen – Name

15.9994 – Atomic mass

Period ↓	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8	9 VIIIB	10	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 0 (zero)
1	1 G H Hydrogen 1.008																	2 G He Helium 4.003
2	3 S Li Lithium 6.941	4 S Be Beryllium 9.0121											5 S B Boron 10.811	6 S C Carbon 12.011	7 G N Nitrogen 14.007	8 G O Oxygen 15.999	9 G F Fluorine 18.998	10 G Ne Neon 20.180
3	11 S Na Sodium 22.999	12 S Mg Magnesium 24.305											13 S Al Aluminium 26.982	14 S Si Silicon 28.086	15 S P Phosphorus 30.914	16 S S Sulphur 32.066	17 G Cl Chlorine 35.453	18 G Ar Argon 39.948
4	19 S K Potassium 39.098	20 S Ca Calcium 40.078	21 S Sc Scandium 44.956	22 S Ti Titanium 47.867	23 S V Vanadium 50.942	24 S Cr Chromium 51.996	25 S Mn Manganese 54.938	26 S Fe Iron 55.845	27 S Co Cobalt 58.933	28 S Ni Nickel 58.693	29 S Cu Copper 63.546	30 S Zn Zinc 65.39	31 S Ga Gallium 69.723	32 S Ge Germanium 72.61	33 S As Arsenic 74.922	34 S Se Selenium 78.96	35 L Br Bromine 79.904	36 G Kr Krypton 83.30
5	37 S Rb Rubidium 85.468	38 S Sr Strontium 87. 62	39 S Y Ytterium 89.906	40 S Zr Zirconium 91.224	41 S Nb Niobium 92.906	42 S Mo Molybdenum 95.94	43 X Tc Technetium (98)	44 S Ru Ruthenium 101.07	45 S Rh Rhodium 102.906	46 S Pd Palladium 106.42	47 S Ag Silver 107.868	48 S Cd Cadmium 112.411	49 S In Indium 114.818	50 S Sn Tin 118.710	51 S Sb Antimony 121.760	52 S Te Tellurium 177.60	53 S I Iodine 126.904	54 G Xe Xenon 131.29
6	55 S Cs Caesium 132.905	56 S Ba Barium 137.327	57 S La Lanthanum 138.906	72 S Hf Hafnium 178.49	73 S Ta Tantalum 180.948	74 S W Tungsten 183.84	75 S Re Rhenium 186.207	76 S Os Osmium 190.23	77 S Ir Iridium 192.217	78 S Pt Platinum 195.078	79 S Au Gold 196.967	80 L Hg Mercury 200.59	81 S Tl Thallium 204.383	82 S Pb Lead 207.2	83 S Bi Bismuth 208.980	84 S Po Polonium (209)	85 S At Astatine (210)	86 G Rn Radon (222)
7	87 S Fr Francium (223)	88 S Ra Radium (226)	89 S Ac Actinium (227)	104 X Rf Rutherfordium (261)	105 X Db Dubnium (262)	106 X Sg Seaborgium (263)	107 X Bh Bohrium (262)	108 X Hs Hassium (265)	109 X Mt Meitnerium (266)	110 X Ds Darmstadtium (269)	111 X Rg Rontgenium (272)	112 X Cn Copernicium (227)	113 X Nh Nihonium (284)	114 X Fl Flevorium (289)	115 X Mc Moscovium (228)	116 X Lv Livermonium (292)	117 Ts Tennessive (292)	118 X Og Oganesson (294)

f-Block Elements

*Lanthanides (After La)	58 S Ce Cerium 140.116	59 S Pr Praseodymium 140.908	60 S Nd Neodymium 144.908	61 X Pm Promethium (145)	62 S Sm Samarium 150.36	63 S Eu Europium 151.964	64 S Gd Gadolinium 157.25	65 S Tb Terbium 158.925	66 S Dy Dysprosium 162.50	67 S Ho Holmium 164.930	68 S Er Erbium 167.26	69 S Tm Thulium 168.934	70 S Yb Ytterbium 173.04	71 S Lu Lutetium 174.967
**Actinides (After Ac)	90 S Th Thorium 232.038	91 S Pa Protactinium 231.036	92 S U Uranium 238.029	93 S Np Neptunium (237)	94 S Pu Plutonium (244)	95 X Am Americium (243)	96 X Cm Curium (247)	97 X Bk Berkelium (247)	98 X Cf Californium (251)	99 X Es Einsteinium (252)	100 X Fm Fermium (257)	101 X Md Mendelevium (258)	102 X No Nobelium (259)	103 X Lr Lawrencium (262)

☐ Metals

▨ Metalloids

☐ Non-metals

- The table contains 7 periods (representing 7 orbitals) and 18 groups (1-18). The concept of subgroup A and B is removed and groups are given no. 1 to18. The name of zero group is changed to group 18.

Classification of Elements

Depending upon the orbital, in which last electron enters, the elements are classified as

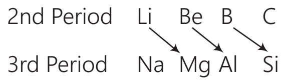
1. *s*-Block Elements

- I and II group elements belong to this block and the last electron enters in s-orbital.
- General electronic configuration is ns^{1-2}.

2. *p*-Block Elements

- Last electron enters in *p*-orbital.
- General electronic configuration is $ns^2, \ np^{1-6}$.
- Groups 13th to 18th excluding He, belongs to this block.

NOTE (i) **Diagonal relationship** The first three elements of second period (Li, Be, B) show similarity in properties with elements (Mg, Al, Si) of third period placed diagonally on the right hand side. This is called diagonal relationship.



(ii) *s* and *p*-block elements are known as representative elements.

3. *d*-Block or Transition Elements

- General electronic configuration is $ns^{1-2}\ (n-1)\ d^{1-10}$.
- Last electron enters in *d*-orbital.
- Comprises of 10 groups → from 3rd to 12th or four series $3d, 4d, 5d, 6d$.

4. *f*-Block or Inner-Transition Elements

- General electronic configuration is

$$ns^2(n-1)d^{0-1}(n-2)f^{1-14}$$

- Last electron enters in *f*-orbital.
- It constitue two series of element *viz.* 5*f* (lanthanoids) and 5*f* (actinoids).
- 4*f*-elements, also known as rare earth elements.

Nomenclature of Elements

The names are derived by using roots for the three digit in the atomic number of the element followed by adding ium at the end. The roots for the numbers are as

Notation for IUPAC Nomenclature of Elements

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

Periodic Trends in Properties of Elements

The elements in the periodic table are arranged in order of increasing atomic number. All of these elements display several other trends and we can use the periodic law and table formation to predict their chemical, physical and atomic properties.

1. Atomic Radii

As going down the group, the new shells are added and thus atomic radius increases whereas on moving along a period, the effective nuclear charge increases and thus the atomic radius decreases.

Atomic radii are of following types:

(i) **Covalent Radius** It is one-half of the internuclear distance between two identical or almost identical atoms bonded by a single covalent bond.

$$r_{\text{covalent}} = \frac{1}{2}\,[\text{Internuclear distance between two bonded atoms}]$$

(ii) **Metallic Radius** It is one-half of the internuclear distance between two adjacent metal atoms in the metallic closed packed crystal lattice.

$$r' = \frac{d}{2} \quad r' > \text{actual size of atom}$$

(iii) **van der Waals' Radius** It is one-half of the internuclear distance between two adjacent atoms belonging to two nearest neighbouring molecules of the substance in solid state.

van der Waals' radius

$$= \frac{1}{2} \times d \begin{bmatrix} \text{internuclear distance between two} \\ \text{non - bonded neighbouring atoms} \\ \text{of two covalently bonded molecules} \end{bmatrix}$$

NOTE Cs is largest in size. Its size is approximately 4 times to that of H-atom.

2. Ionic Radii

The distance from the centre of nucleus to the most distant electron cloud bound to it within an ion, is called ionic radius. In other words, "the distance between nucleus and the outermost shell containing an electron, in an ion is called ionic radius."

The ionic radii are of following two types

(i) **Cationic radii** The size of a cation is always smaller than that of its parent atom.

(ii) **Anionic radii** The size of anion is always larger than that of its parent atom.

NOTE For isoelectronic species : (In general)

(i) Ionic radii $\propto$ charge on anion $\propto \frac{1}{\text{charge on cation}}$

(ii) Ionic radii $\propto \frac{1}{Z_{\text{eff}}}$.

3. Ionisation Enthalpy [IE]

- It is the energy required when an electron is removed from an isolated gaseous atom. Generally left to right in periods IE increases; down the group it decreases but half-filled orbitals and fully filled orbitals are stable forms and have high IE.

  Various factors with which ionisation energy varies are:

  (i) Atomic size $\propto \frac{1}{\text{IE}}$ (ii) Screening effect $\propto \frac{1}{\text{IE}}$

  (iii) Nuclear charge $\propto$ IE
- Helium has the highest IE_1, while Cs has the lowest IE. Values of inert gases are exceptionally high due to stable configuration.
- First IE of group 2 elements is greater than the corresponding elements of group 13 due to fully filled *s*-orbitals.
- First IE of group 15 elements is greater than the corresponding elements of group 16 due to the presence of stable half-filled configuration.
- **Shielding** or **screening effect** is the repulsion of valence electrons by the electrons in penultimate shell to reduce effective nuclear charge.

$$\boxed{Z_{\text{eff}} = Z - \sigma}$$

  where, Z = atomic number,
  σ = shielding constant

  σ = [0.35 × number of electrons in nth shell excluding valence last electron] + [0.85 × number of electrons in $(n-1)$ th shell] + [1.0 × number of electrons in inner shell]. Take 0.30 instead of 0.35 for 1st orbit, i.e. $(n = 1)$
- **Effective nuclear charge** is the actual nuclear charge whose electrostatic force of attraction is being experienced by the outer electron. Greater the effective nuclear charge, more tightly the electrons hold with nucleus.

4. Electron Gain Enthalpy

- It is the energy released when an electron is added in an isolated gaseous atom.
- Left to right in a period electron gain enthalpy increases; down the group it decreases.
- Chlorine has the highest electron affinity but oxidising power of fluorine is larger than chlorine.
- Various factors affecting electron gain enthalpy are:

  (i) Atomic size $\propto \frac{1}{\text{electron gain enthalpy}}$

  (ii) Nuclear charge $\propto \frac{1}{\text{electron gain enthalpy}}$

  (i.e. ΔH is more (–) ve)

  (iii) Configuration : Half-filled orbitals and fully filled orbitals are stable form, therefore electron gain enthalpy will be low.

  Exception Cl > F (electron gain enthalpy)
  S > O

  ∵ F and O-atoms have small size and high charge density, therefore, extra electron experiences repulsive forces hence, they have lower electron gain enthalpy.

5. Electronegativity

- The tendency of an atom to attract the shared pair of electrons towards itself in a covalent bond is called electronegativity. F is the most electronegative element while Cs is the least.
- Difference in the electronegativities of two atoms (A and B) is $\boldsymbol{\chi_A - \chi_B = 0.208\sqrt{\Delta}}$

  where, Δ = actual bond energy – energy for 100% covalent bond

  $$\therefore \quad \Delta = E_{A-B} - \sqrt{E_{A-A} \times E_{B-B}}$$

  Here,
  E_{A-B} = dissociation enthalpy of $A - B$ (kcal mol^{-1})
  E_{A-A} = dissociation enthalpy of $A - A$ (kcal mol^{-1})
  E_{B-B} = dissociation enthalpy of $B - B$ (kcal mol^{-1})
- For noble gases, its value is taken as zero.
- Decreasing order of electronegativity,

  $$\text{F} > \text{O} > \text{Cl} \approx \text{N} > \text{Br} > \text{S} \approx \text{C} > \text{I} > \text{H}$$

  In periods : left to right electronegativity increases.

  In groups : down the group electronegativity decreases.

NOTE **% age s-character $\propto$ electronegativity**

∴ *sp* hybridised carbon is more electronegative than a sp^2 hybridised carbon, which in turn is more electronegative than a sp^3-hybridised carbon.

6. Valency

Valency of an element is usually determined by electrons present in valence shell (outer shell).

- **Variation of Valency in a Period** Valency of the elements with respect to hydrogen increases from 1 to 4 up to 14th group and then decreases to 1.

  e.g. NaH, CaH_2, AlH_3, SiH_4, PH_3, H_2S, HCl.

But valency of the elements with respect to oxygen increases from one to seven along a period.

e.g. Na_2O, CaO, Al_2O_3, SiO_2, P_4O_{10}, SO_3, Cl_2O_7

- **Variation of Valency in a Group** All elements in a group have the same valency. *p*-block elements show variable valency on account of inert pair effect.
- **Variation of Valency in Transition Elements** Transition metals and inner-transition metals show variable valency of 1, 2 or 3 as they use electrons from outer as well as penultimate or inner-penultimate shell.

NOTE
- Term oxidation state indicates the apparent charge present on atom in that particular molecule. It follows the same trend along period or group as valency.
- Os has the highest oxidation state, i.e. (+8).

7. Chemical Reactivity

- Reactivity of metals increases with decrease in ionisation energy, electronegativity and increase in atomic radii and electropositive character.
- Reactivity of non-metals increases with increase in electronegativity and electron gain enthalpy and decreases with increase in atomic radii.
- **Variation in a Group** On moving down the group, reactivity of metals increases while for non-metals it decreases.
- **Variation in a Period** Reactivity of metals decreases, while reactivity of non-metals increases across the period.

8. Reducing and Oxidising Character

- The reducing character of the elements increases down the group and decreases along the period. The oxidising character of the element decreases down the group and increases along the period
- Alkali metals are most reactive, strongest reducing agent and have lowest ionisation enthalpy.
- Halogens are most reactive, strongest oxidising agent and have higher ionisation enthalpy.

9. Nature of Oxides

- Basic strength of the oxides or hydroxides increases down the group and decreases along the period. Acidic nature of oxides increases across the period and decreases down the group.
- The acidic character of oxides of a non-metal increases with increase in oxygen content, e.g. $Cl_2O < Cl_2O_3 < Cl_2O_7$
- Almost all metallic oxides are basic but ZnO and Al_2O_3 are amphoteric. Similarly non-metallic oxides are acidic but CO, N_2O and NO are neutral. Cl_2O_7 is the most acidic oxide.

DAY PRACTICE SESSION 1

FOUNDATION QUESTIONS EXERCISE

1 The modern periodic table has been divide into
(a) seven periods : three short and four long
(b) nine periods : six short and three long
(c) eighteen periods : four short and four long
(d) nine periods from zero to VIII

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

3 The elements having the electronic configuration [Kr] $4d^{10} f^{14}, 5s^2p^6d^2, 6s^2$ belongs to
(a) *s*-block (b) *p* -block
(c) *d*-block (d) *f*-block

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

5 An atom has electronic configuration $1s^2, 2s^2, 2p^6, 3s^2\, 3p^6\, 3d^3\, 4s^2$, it would be placed in
(a) fifth group (d) fifteen group
(c) second group (d) third group

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

7 Chemical properties of lithium and magnesium are similar because

(a) these belong to same group
(b) they shown diagonal relationship
(c) both has same ionisation potential
(d) both has same electron affinity

8 The correct order of atomic radii in group 13 elements is **→ NEET 2018**

(a) B < Ga < Al < Tl < In (b) B < Al < Ga < In < Tl
(c) B < Al < In < Ga < Tl (d) B < Ga < Al < In < Tl

9 Which of the following pairs has almost same atomic radii?

(a) Al, Ga (b) Be, Mg (c) Mg, Al (d) B, Be

10 A trend common to both for group I and VII elements in the periodic table as atomic number increases its

(a) atomic radius increases
(b) oxidising power increases
(c) reactivity with water increases
(d) maximum valency increases

11 Indetify the correct order of the size of the following.

(a) $Ca^{2+} < K^+ < Ar < S^{2-} < Cl^-$
(b) $Ca^{2+} < K^+ < Ar < Cl^- < S^{2-}$
(c) $Ar < Ca^{2+} < K^+ < Cl^- < S^{2-}$
(d) $Ca^{2+} < Ar < K^+ < Cl^- < S^{2-}$

12 In the periodic table, with the increase in atomic number the metallic character of the element

(a) decreases in a period and increases in a group
(b) increases in period and decreases in a group
(c) increases in a period as well as in the group
(d) decreases in a period and also in the group

13 Which of the following order of ionic radii is correctly represented?

(a) $H^- > H > H^+$ (b) $Na^+ > F^- > O^{2-}$
(c) $F^- > O^{2-} > Na^+$ (d) $Al^{3+} > Mg^{2+} > N^{3-}$

14 Identify the wrong statement in the following. **→ CBSE-AIPMT 2012**

(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 we move down the first group of the periodic table
(d) Atomic radius of the elements decreases as we move across from left to right in the second period of the periodic table

15 The correct order of the decrease in ionic radii among the followin isoelectronic species is **→ CBSE-AIPMT 2010**

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

16 Which one of the following is the correct order of size of the iodine species?

(a) $I > I^+ > I^-$ (b) $I > I^- > I^+$
(c) $I^+ > I^- > I$ (d) $I^- > I > I^+$

17 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

18 The species Ar, K^+ and Ca^{2+} contain the same number of electrons. In which order do their radii increases. **→ CBSE-AIPMT 2015**

(a) $Ar < K^+ < Ca^{2+}$ (b) $Ca^{2+} < Ar < K^+$
(c) $Ca^{2+} < K^+ < Ar$ (d) $K^+ < Ar < Ca^{2+}$

19 In which of the following crystals of ionic compounds would you expect maximum distance between the centres of cations and anions?

(a) LiF (b) CsF
(c) CsI (d) LiI

20 Amongst the elements with following electronic configuration which one of them may have the highest ionisation energy? **→ CBSE-AIPMT 2009**

(a) [Ne] $3s^2, 3p^3$ (b) [Ne] $3s^2, 3p^2$
(c) [Ar] $3d^{10}, 4s^2, 4p^3$ (d) [Ne] $3s^2, 3p^1$

21 The first ionisation enthalpies of Na, Mg, Al and Si are

(a) Na < Mg > Al < Si (b) Na > Mg > Al > Si
(c) Na < Mg < Al < Si (d) Na > Mg > Al < Si

22 The order of screening effect of electrons of *s*, *p*, *d* and *f*-orbitals of a given shell of an atom on its outer shell electrons is

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

23 Which one of the following arrangements represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species? **→ [CBSE-AIPMT 2010]**

(a) Cl < F < S < O
(b) O < S < F < Cl
(c) S < O < Cl < F
(d) F < Cl < O < S

24 The electron affinity of halogens are F = 322, Cl = 349, Br = 324, I = 259 kJ mol^{-1}. The higher value of Cl as compared to that of F is due to

(a) weaker electron-electron repulsion in Cl
(b) higher atomic radius of F
(c) smaller electronegativity of F
(d) more vacant *p*-subshell in Cl

25 In which of the following options the order of arrangement does not agree with the variation of property indicated against it ? **→ NEET 2016, Phase (I)**

(a) B < C < N < O (increasing first ionisation enthalpy)
(b) I < Br < Cl < F (increasing electron gain enthalpy)
(c) Li < Na < K < Rb (increasing metallic radius)
(d) $Al^{3+} < Mg^{2+} < Na^+ < F^-$ (increasing ionic size)

26 The electronegativity of the following elements increases in the order **→ CBSE-AIPMT 2012**
(a) C, N, Si, P (b) N, Si, C, P
(c) Si, P, C, N (d) P, Si, N, C

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

28 Which one of the elements with the following outer orbital configurations may exhibit the largest number of oxidation states? **→ CBSE-AIPMT 2009**
(a) $3d^3, 4s^2$ (b) $3d^5, 4s^1$ (c) $3d^5, 4s^2$ (d) $3d^2, 4s^2$

29 Which of the following has the correct order of basic character?
(a) $MgO > Na_2O > CuO > Al_2O_3$
(b) $Na_2O > MgO > Al_2O_3 > CuO$
(c) $CuO > MgO > Na_2O > Al_2O_3$
(d) $Al_2O_3 > CuO > MgO > Na_2O$

30 Which of the following is not expected to react with sodium hydroxide? **→ CBSE-AIPMT 2011**
(a) B_2O_3 (b) CaO (c) SiO_2 (d) BaO

31 In which of the following option, the order of arrangements does not agree with the variation of property indicated against it?
(a) Li < Be < B < C → Ist ionisation potential
(b) $Li^+ < Na^+ < K^+ < Cs^+ \rightarrow$ Ionic radius
(c) $NH_3 < PH_3 < AsH_3 \rightarrow$ Acidic character
(d) $Al_2O_3 < MgO < Na_2O < K_2O \rightarrow$ Basic character

Direction (Q.Nos. 32-34) *Each of these questions contains two statements : Assertion and Reason. Each of these questions also has four alternative choices, only one of which is the correct answer. You have to select one of the code (a), (b), (c) and (d) given below.*
(a) Assertion is true, Reason is true; Reason is a correct explanation for Assertion
(b) Assertion is true, Reason is true; Reason is not a correct explanation for Assertion
(c) Assertion is true, Reason is false
(d) Assertion is false, Reason is true

32 **Assertion** Electron gain enthalpy becomes less negative as we go down the group.

Reason Size of the atom increases on going down the group and the added electron would be farther from the nucleus.

33 **Assertion** Noble gases have large positive electron gain enthalpy.

Reason Electron has to enter the next higher principal quantum level.

34 **Assertion** Removal of *s*-electron is relatively difficult than removal of *p*-electron of the same main shell.

Reason *s*-electrons are closer to the nucleus than *p*-electrons of the same shell and hence, are more strongly attracted by nucleus.

DAY PRACTICE SESSION 2

PROGRESSIVE QUESTIONS EXERCISE

1 In which of the following options the order of arrangement does not agree with the variation of property indicated aginst it.
(a) Ba > Sr > Mg ; atomic radius
(b) F > O > N ; first ionisation enthalpy
(c) Cl > F > I ; electron affinity
(d) O > Se > Te; electronegativity

2 The ionisation energy of nitrogen is more than that of oxygen because
(a) nitrogen has half-filled *p*-orbitals
(b) nitrogen is left to the oxygen in the same period of the periodic table
(c) nitrogen contains less number of electrons
(d) nitrogen is less electronegative

3 Which of the following statements regarding the chemical reactivity of the alkali metals and the halogens is correct?
(a) In alkali metals the chemical reactivity increases whereas in the halogens, it decreases with increase in atomic number down the group
(b) With increase atomic number down the group, the chemical reactivity increases in both alkali metals and halogens
(c) With increase in atomic number down the group, the chemical reactivity decreases in both alkali metals and halogens
(d) The chemical reactivity decreases in the alkali metals but increases in the halogens with increase in atomic number

4 The group whose elements have positive as well as negative oxidation states are
(a) Na, Mg, Al
(b) He, Li, Be
(c) H, F, O
(d) H, Cl, Br

5 Three elements *A*, *B* and *C* have atomic numbers 20, 36 and 35 respectively. Then the correct statement is
(a) B would have highest ionisation potential
(b) C would have highest ionisation potential
(c) B would have an ionisation potential between those of *A* and *C*.
(d) *A* would have highest ionisation potential

6 Which of the following option is correct regarding the ionic radii?
(a) $^{35}Cl^- > {}^{37}Cl^-$
(b) $K^+ > Cl^-$
(c) $Ti^{4+} > Mn^{7+}$
(d) $P^{5+} > P^{3+}$

7 The ionic radii of C^{4-} and O^{2-} respectively are 2.60 and 1.4 (Å). The ionic radius of the isoelectronic ion N^{3-} would be
(a) 2.6
(b) 1.71
(c) 1.4
(d) 0.95

8 Which of the following electronic configuration exhibits a sudden large jump between the value of IE_2 and IE_3?
(a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
(b) $1s^2\ 2s^2\ 2p^6\ 3s^1\ 3p^2$
(c) $1s^2\ 2s^2\ 2p^6\ 3s^1$
(d) $1s^2\ 2s^2\ 2p^6\ 3s^2$

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

10 Choose the correct option with respect to first (I) and second (II) ionisation potentials of sodium and magnesium?
(a) $I_{Na} > I_{Mg}$
(b) $II_{Na} > II_{Mg}$
(c) $II_{Mg} < II_{Na}$
(d) $II_{Mg} = II_{Na}$

11 In group 17, fluorine, chlorine, bromine and iodine are placed. This is due to
(a) their atoms are generally univalent
(b) they are non-metals
(c) they have 7 electrons in the outermost shell of their atom
(d) they are electronegative

12 Which is incorrect statement?
(a) In solid O^{2-} is stabilised by neighbouring cations
(b) Formation of O^{2-} from O^- is unfavourable in the gas phase
(c) Electron affinity of O > S
(d) All of the above are incorrect statements

13 For the following process,

$$A(g) + e^- \longrightarrow A^-(g)\ ;\ \Delta H = x$$
$$A^-(g) \longrightarrow A(g) + e^-\ ;\ \Delta H = y$$

Select the correct alternate.
(a) Ionisation energy of $A^-(g)$ is y
(b) Electron affinity of $A\,(g)$ is $-y$
(c) Electron affinity of $A(g)$ is x
(d) All of the above are correct

14 Among the following oxide, the least acidic is
(a) P_4O_{10} (b) As_4O_6 (c) As_4O_{10} (d) P_4O_6

15 Arrange the following four elements in the correct order of magnitude of their electron affinity.
(1) $2s^2\ 2p^5$
(2) $3s^2 3p^4$
(3) $3s^2\ 3p^5$
(4) $2s^2\ 2p^4$
(a) 4 < 2 < 1 < 3
(b) 1 < 4< 3 <2
(c) 1 < 3 < 4< 2
(d) 4 < 3 < 1 < 2

16 Which of the following is not arranged in the correct sequence?
(a) MO, M_2O_3, MO_2, M_2O_5 - decreasing basic strength
(b) Sc, V, Cr, Mn - increasing number of oxidation states
(c) $Co^{+2}, Fe^{+3}, Cr^{+3}, Sc^{+3}$- increasing stability
(d) d^5, d^3, d^1, d^4-increasing magnetic moment

ANSWERS

SESSION 1	**1** (a)	**2** (c)	**3** (c)	**4** (c)	**5** (a)	**6** (b)	**7** (b)	**8** (d)	**9** (a)	**10** (a)
	11 (b)	**12** (a)	**13** (a)	**14** (a)	**15** (c)	**16** (d)	**17** (a)	**18** (c)	**19** (c)	**20** (a)
	21 (a)	**22** (a)	**23** (b)	**24** (a)	**25** (a,b)	**26** (c)	**27** (a)	**28** (c)	**29** (b)	**30** (b)
	31 (a)	**32** (a)	**33** (a)	**34** (a)						
SESSION 2	**1** (b)	**2** (a)	**3** (a)	**4** (d)	**5** (a)	**6** (c)	**7** (b)	**8** (d)	**9** (b)	**10** (b)
	11 (c)	**12** (c)	**13** (d)	**14** (b)	**15** (a)	**16** (d)				

Hints and Explanations

SESSION 1

1 The modern periodic table has been divided into seven horizontal rows called periods and eighteen vertical columns called groups.

2 As the element '*X*' belongs to fourth period and fifteenth group of periodic table, its outer configuration is $4s^2\ 4p^3$.

Its complete electronic configuration is

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

Hence, its *s*-orbital is completely filled, while *p*-orbital is half-filled.

3 Since, the last electron enters in *d*-subshell, therefore element belongs to *d*-block

4 In a family, all elements have same outermost electronic configuration. Since, [Ne] $3s^2\ 3p^5$, chlorine belongs to halogen family, while the remaining three are in same group, i.e. group 12.

$_{80}\text{Hg} = [\text{Xe}]\ 4f^{14}\ 5d^{10}\ 6s^2$

$_{48}\text{Cd} = [\text{Kr}]\ 4d^{10}\ 5s^2$

$_{30}\text{Zn} = [\text{Ar}]\ 3d^{10}\ 4s^2$

5 Electronic configuration of an atom is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3\ 4s^2$ enters into the *d*-orbital, thus it belongs to fifth group.

6 The element with atomic number $Z = 114$ is flerovium (Fl). It is a super heavy artificial chemical element. In the periodic table of the elements, it is a transactinide element in the *p*-block. It is a member of the 7th period and is the heaviest known member of carbon family.

7 Chemical properties of lithium and magnesium are similar because both exhibits diagonal relationship due to same electronegativity and charge 1 radius ratio.

8 The atomic radii as well as ionic radii increases on moving down the group 13 elements because of the successive addition of one extra shell of electrons.

However, there is an anomaly at gallium in case of atomic radii. Atomic radii of Ga is lesser as compared to Al. Gallium (Ga) with electronic configuration, $[\text{Ar}]_{18}\ 3d^{10}\ 4s^2\ 4p^1$ has an extra *d*-electrons which do not screen the nucleus effectively. Consequently, electrons of Ga are more attracted by nucleus.

Thus, the increasing order of atomic radii of the group 13 elements is B (85 pm) < Ga (135 pm) < Al (143 pm) < In (167 pm) < Tl (170 pm)

9 $\text{Al} = [\text{Ne}]\ 3s^2\ 3p^1$

$\text{Ga} = [\text{Ar}]\ 3d^{10}\ 4s^2\ 4p^1$

Electronic configuration of Ga exhibits that it contains 10 electrons in *d*-orbitals which do not screen the nucleus effectively because of the large size of *d*-orbitals. The electrons are attracted by a greater force toward nucleus, that's why atomic radii of Ga and Al are almost same.

10 Atomic radius increases as atomic number increases due to increase in the number of shells.

11 A cation has always the lesser ionic size than a metal atom due to loss of electrons and an anion has always the greater size than metal atom due to gain of electrons. The given species are isoelectronic species as they contain same number of electrons.

For isoelectronic species,

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

Ion :	Ca^{2+}	K^+	Ar	S^{2-}	Cl^-
Atomic number :	20	19	18	16	17

So, the correct order of size is $Ca^{2+} < K^+ < Ar < Cl^- < S^{2-}$

12 In periodic table, the metallic character increases down the group because the ionisation enthalpy decreases down the group and metallic character decreases from left to right because the ionisation enthalpy increases from left to right.

13 (a) $H^- > H > H^+$

It is known that radius of a cation is always smaller than that of neutral atom due to decrease in the number of orbits.

Whereas, the radius of anion is always greater than a cation due to decrease in effective nuclear charge.

Hence, the correct order is $H^- > H > H^+$.

14 Amongst isoelectronic species, ionic radius increases with increase in negative charge or decrease with the increase in positive charge.

15 Ionic radii ∝ charge on anion

$$\propto \frac{1}{\text{charge on cation}}$$

During the formation of a cation, the electrons are lost from the outer shell and the remaining electrons experience a great force of attraction by the nucleus, i.e. attracted more towards the nucleus. In other words, nucleus hold the remaining electrons more tightly and this results in decreased radii.

However, in case of anion formation, the addition of electron(s) takes place in the same outer shell, thus the hold of nucleus on the electrons of outer shell decreases and this results in increased ionic radii.

Thus, the correct order of ionic radii is

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

16 $I^- > I > I^+$

Size of anion is always larger than its neutral atom, while size of cation is always smaller than its neutral atom.

17 $\text{Ionic radii} \propto \dfrac{1}{Z_{\text{eff}}}$

where, Z_{eff} = effective nuclear charge

$= Z -$ screening constant (σ)

The value of screening constant is based upon the number of electrons in valence shell as well as in penultimate shells.

18 $Ca^{2+} < K^+ < Ar$

Ar, K^+ and Ca^{2+} are isoelectronic, i.e. with same number of electrons, 18. For isoelectronic species, ionic radii decreases with increase in effective (relative) positive charge. Also Ar, K and Ca belong to the same period (3rd period).

19 Ionic radii is the distance between the nucleus of an ion and a point up to which the nucleus has its influence on its electron cloud. The size of ions increases on moving from top to bottom in a group. Hence, the maximum distance between the centres of cations and anions is in CsI because Cs is the largest cation and I is the largest anion.

20 Ionisation energy decreases as we move down the group and is more for half-filled outermost sub-shell.

21 First ionisation enthalpies of Na, Mg, Al and Si are

$\text{Na} = 496\ \text{kJ mol}^{-1}$,

$\text{Mg} = 737.6\ \text{kJ mol}^{-1}$

$\text{Al} = 577\ \text{kJ mol}^{-1}$,

$\text{Si} = 786\ \text{kJ mol}^{-1}$

Therefore, order of first ionisation enthalpies is $\text{Na} < \text{Mg} > \text{Al} < \text{Si}$

22 For the same shell, screening effect decreases in the order : $s > p > d > f$

23 The correct order of electron gain enthalpy (electron affinity) is

$$\underset{(1.48)}{\text{O}} < \underset{2.07}{\text{S}} < \underset{3.45}{\text{F}} < \underset{3.6\ (\text{eV})}{\text{Cl}}$$

24 Electron affinity value of Cl is higher than that of F as, Cl belongs to third period (have larger size), while F belongs to second period (have small size). In Cl, electron-electron repulsive forces are weaker than that of F.

25 (*a, b*) For option (*a*),

First ionisation energy is the energy required to remove an electron from outermost shell.

Hence, correct order is

$$\text{B} < \text{C} < \text{O} < \text{N}.$$

For option (b),

Electron gain enthalpy is the energy change to gain an electron in the outermost shell.

Hence, the correct order is

$$\text{I} < \text{Br} < \text{F} < \text{Cl}.$$

For option (c),

As we move down the group in alkali metal, metallic radius increases Li < Na < K < Rb.

For option (d),

In case of isoelectronic species, as positive charge decreases or negative charge increases the ionic size of the species increases and *vice-versa*

$\text{Al}^{3+} < \text{Mg}^{2+} < \text{Na}^{+} < \text{F}^{-}$

26 On moving along a period from left to right in periodic table, electronegativity increases (due to decrease in size). While on moving downward in a group, electronegativity decreases. Thus, the correct order of electronegativity is

$$\underset{(1.8)}{\text{Si}} < \underset{(2.1)}{\text{P}} < \underset{(2.5)}{\text{C}} < \underset{(3.0)}{\text{N}}$$

27 Electronegativity increases along a period and decreases in a group. Thus, the order is

II > III > I > IV

28 The sum of number of electrons (unpaired) in *d*-orbitals and number of electrons in *s*-orbital gives the number of oxidation states (os)exhibited by a *d*-block elements. Therefore,

(a) $3d^3, 4s^2 \Rightarrow \text{OS} = 3 + 2 = 5$

(b) $3d^5, 4s^1 \Rightarrow \text{OS} = 5 + 1 = 6$

(c) $3d^5, 4s^2 \Rightarrow \text{OS} = 5 + 2 = 7$

(d) $3d^2, 4s^2 \Rightarrow \text{OS} = 2 + 2 = 4$

Hence, element with $3d^5, 4s^2$ configuration exhibits largest number of oxidation states.

29 Basic character depends upon the metallic character, i.e. greater the metallic character, more is the basic character. Since the metallic character decreases from Na to Cu. Thus, basic character also decreases.

30 Sodium hydroxide, NaOH, being a strong alkali never react with a basic oxide (compound). Among the given options, B_2O_3 and BeO are amphoteric oxides, SiO_2 is an acidic oxide and CaO is a basic oxide. Therefore, NaOH does not react with CaO.

31 Li, Be, B and C are present in IInd period. In a period from left to right ionisation potential increases.

$$\xrightarrow[\text{Li Be B C}]{\text{Ionisation potential increases}}$$

But in case of Be and B, Be has higher ionisation potential than B due to stable configuration of Be.

$${}_4\text{Be} = 1s^2, 2s^2$$

$\underset{2s^2}{\boxed{\uparrow\downarrow}}$ Stable configuration (due to fully-filled orbital)

$${}_5\text{B} = 1s^2, 2s^2\ 2p^1$$

$\underset{2p^1}{\boxed{\uparrow}\boxed{}\boxed{}}$ Unstable configuration

So, the correct order of ionisation potential of given elements is Li < B < Be < C

32 Electron gain enthalpy becomes less negative as the size of an atom increases down the group. This is because within a group, screening effect increases on going downward and the added electron would be farther away from the nucleus.

33 Noble gases have large positive electron gain enthalpy because the electron has to enter the next high principal quantum level leading to a very unstable electronic configuration.

34 Reason is the correct explanation of Assertion.

SESSION 2

1 First Lonisation enthalpy increases from left to right in a period but due to half-filled electronic configuration of nitrogen atom, N has higher first ionisation enthalpy than of oxygen. Hence (b) is not in agreement with the given order.

2 Ionisation energy increases as we move from left to right in any period but the value of ionisation energy increases, if the given species have half to full-filled electronic configuration w.r.t the next member is the period. Thus, ionisation energy of nitrogen is more than of oxygen.

3 In alkali metals, their chemical reactivity increases due to presence of only one valence electron, which it can lose easily, whereas halogen have seven valence electrons. So it can complete its octate by easily accepting one electron and becomes stable. Thus, it is also a reactive species but its reactivity decreases as we move down the group due to increase in atomic number and atomic size.

4 The element hydrogen, chlorine and bromine can form (+) ve as well as (–) ve ions, thus can show (+) ve as well as (–) ve oxidation states.

Hydrogen $\rightarrow \text{H}^{+}$ and H^{-}

Chlorine $\rightarrow \text{Cl}^{-}$ and (+) 3, (+) 5, and (+7)

Bromine → and (+ 1), (+ 3), (+ 5) and (+ 7).

5 Electronic configuration for the elements (*A*), (*B*) and (*C*) are

(*A*), ($Z = 20$) = 2, 8, 8, 2 (4th period)

(*B*), ($Z = 36$) = 2, 8, 18, 8 (4th period)

(*C*), ($Z = 35$) = 2, 8, 18, 7 (4th period)

$\because$ Element (*B*) has completely full-filled outermost orbitals and ionisation energy increases from left to right in any period. Thus, (*B*) has highest ionisation potential.

6 (a) $^{35}Cl^-$ and $^{37}Cl^-$ is wrong,

∵ Both have same number of protons and electrons.

(b) K^+ and Cl^- are isoelectronic,

∵ for isoelectronic species, the size of anion is more than of cation.

(c) Ti^{4+} has $4d^\circ$ configuration, while Mn^{7+} has $4d^\circ\, 4s^\circ$ configuration but

$$\text{Ionic radii} \propto \frac{1}{(+)\text{ ve charge}}$$

Thus, size of $Ti^{4+} > Mn^{7+}$

(d) For the same element, the ion having more (+) ve charge is comparatively smaller in size. i.e. $P^{5+} > P^{3+}$ incorrect.

∴ (c) is the correct option.

7 ∵ Number of electronic in

$C^{4-} = 6 + 4 = 10$

$O^{2-} = 8 + 2 = 10$

$N^{3-} = 7 + 3 = 10$

Thus, all are isoelectronic,

For isoelectronic species, more be the (–) ve charge over the ion, more is the size of that species.

Hence, the size of N^{3-} should be greater than O^{2-} but less than C^{4-}.

∴ (b) is the correct answer.

8 ∵ (i) Ionisation energy increases with the decreases in size.

(ii) Completely filled or half-filled valency orbital has higher ionisation energy. As, electronic configuration of option (d) has only 3 main shells with two electrons in $3s^2$. Therefore, after the removal of 2nd electron i.e. IE_2 the configuration of element becomes full-filled. Hence, IE_3 for this species is very high, therefore the value of ioniration enthalpy from (2) to (3) jumps with a sudden rise.

9 Increasing order of first ionization enthalpy is, Ba < Ca < Se < S < Ar.

$$r_n \propto \frac{1}{Z} \propto \frac{1}{\text{I.E}}$$

10 Electronic configuration (E.C) for Na ($Z = 11$) is $1s^2, 2s^2, 2p^6, 3s^1$ (show atom Na) and for Na^+ is $1s^2 2s^2 2p^6$ (show Na^+).

Similarly,

EC for Mg = $1s^2, 2s^2, 2p^6, 3s^2$ and for $Mg^+ = 1s^2, 2s^2\ sp^6 3s^1$ show Mg^+

∵ (i) IE is more for Half of full-filled electronic configuration.

(ii) I.E. increases in the period, as we move from left to right (for atoms).

Hence,

$$II_{Na} > II_{Mg}$$

∵ for II_{Na}, E.C. is full-filled while for II_{Mg} it is not.

11 In modern periodic table. The groups are made on the basis of number of valence electrons.

∵ Member of group-17 (fluorine, chlorine, bromine and iodine) have seven valence electrons thus are placed in group (17)

12 Electron affinity for oxygen (O) is less than of sulphur (S), because of presence of six valence electrons is a very small volume, which offers large electron-electron repulsion on the incoming electron.

On the other hand due to comparatively larger size of sulphur, it offers less repulsion. So, has more electron affinity than of oxygen.

13 Given,

(i) $A(g) + e^- \longrightarrow A^-(g); \Delta H = x$

(ii) $A^-(g) \longrightarrow A(g) + e^-; \Delta H = y$

∵ In relation (i), an atom is changing to anion by accepting an electron, thus it represent electron affinity and its value is x. On the other hand relation (ii) represent the ionisation energy (I.E) and represent the value equal to y. Also, in general

∵ Value of electron affinity = (–) ve of I.E.

∴ (d) is the correct option.

14 (i) The acidic nature of oxides, depends on the value of oxidation number. (O.N) of central atom. More (+) ve be the value, more acidic nature of the given oxide.

(ii) For same value of oxidation number, more electronegative central atom represents more acidic oxide due to its smaller size.

∵ O.N. of P in $P_4O_{10} = 5$

O.N. of As is $As_4O_6 = 3$

O.N. of As in $As_4O_{10} = 5$

O.N. of P in $P_4O_6 = 3$

Also,

∵ between P_4O_6 and As_4O_6, the As is less electronegative, thus As_4O_6 is least acidic.

15 Among the given electronic configuration is.

(1) is $2s^2\, 2p^5$ i.e. has
total no. of electrons = 2 + 7 = 9
Hence, it is fluorine (F)

(2) is $3s^2 3p^4$ i.e. has
total no. of electrons
= 2 + 8 + 6 = 16
Hence, it is sulphur (5)

(3) is $3s^2 3p^5$ i.e. has
total no. of electrons
= 2 + 8 + 7 = 17
Hence, it is chlorine (Cl)

(4) is $2s^2 2p^4$ i.e. has
total no. of electrons = 2 + 6 = 8
Hence, it is oxygen.

∵ Election affinity (E.A) is more for Cl than of F due to electron repulsion in fluorine and electron affinity for sulpher is more than of oxygen for the same reason.

Also, electron affinity increases from left to right in any period. Hence, the correct order is (3) > (1) > (2) > (4).

i.e. $4 < 2 < 1 < 3$

16 More number of unpaired electrons in any species, show more magnetic moment. Thus, option (d) is not in correct sequence. The correct sequence is d^1, d^3, d^4, d^5.