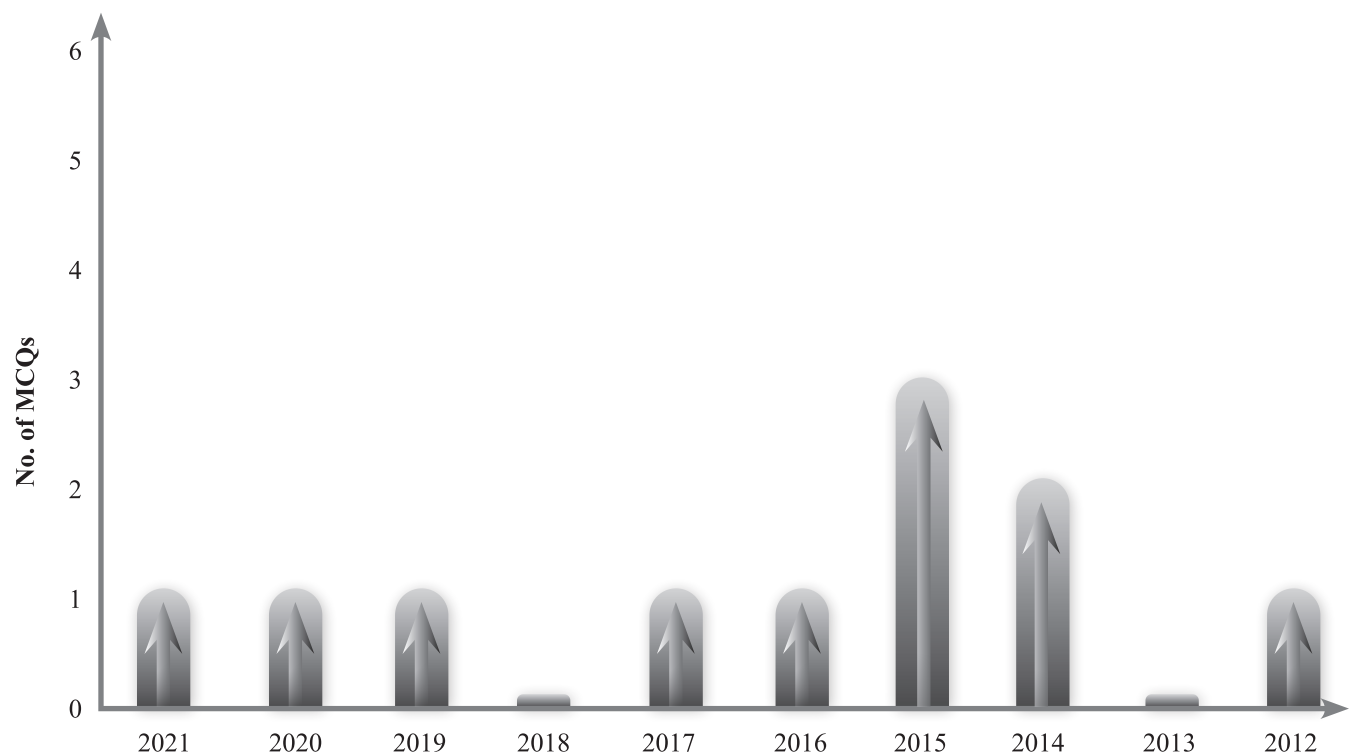
3 Classification of Elements and Periodicity in Properties

Past Years NEET Trend



Investigation Report

TARGET EXAM	PREDICTED NO. OF MCQs	CRITICAL CONCEPTS
NEET	0-1	• Acid base behaviour of oxides and hydroxides, Periodic trends in physical properties of element

Perfect Practice Plan

Topicwise Questions	Learning Plus	Multiconcept MCQs	NEET Past 10 Years Questions	Total MCQs
51	26	19	11	107

GENESIS OF PERIODIC CLASSIFICATION

(a) Dobereiner's Triads

He arranged similar elements in the groups of three elements known called as **triads**, in which the atomic mass of the central element was merely the arithmetic mean of atomic weight of other two elements or all the three elements possessed nearly the same atomic weight.

Dobereiner's Triads

Li	Na	K	
7	23	39	(7 + 39)/2 = 23
Fe	Co	Ni	
55.85	58.93	58.71	Nearly same atomic masses

It was restricted to few elements, hence discarded.

(b) Newland's Law of Octave

He was the first to correlate the chemical properties of the elements with their atomic masses.

According to octave, if the elements are arranged in the order of their increasing atomic weight the eighth element starting from given one is similar in properties to the first element.

This arrangement of elements is called as **Newland's law of Octave**.

Newlands's Octaves

Element	Li	Be	B	C	N	O	F
At. wt.	7	9	11	12	14	16	19
Element	Na	Mg	Al	Si	P	S	Cl
At. wt.	23	24	27	29	31	32	35.5
Element	K	Ca					
At. wt.	39	40					

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

(c) Lothar Meyer's Classification

He determined the atomic volumes by dividing atomic masses with their densities in solid states.

He plotted a graph between atomic masses against their respective atomic volumes for a number of elements. He found the following observations.

Elements with similar properties occupied similar positions on the curve.

On the basis of his observations he concluded that **the atomic volumes (a physical property) of the elements are the periodic functions of their atomic weights.**
It was discarded as it lacks practical utility .

(d) Mendeleev's Periodic Table

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

He arranged the known elements in order of their increasing atomic masses considering the facts that elements with similar properties should occupied the same vertical columns and leaving out blank spaces where necessary.

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

Periods	Number of Elements	Called as
$(1)^{st}$ n = 1	2	Very short period
$(2)^{nd}$ n = 2	8	Short period
$(3)^{rd}$ n = 3	8	Short period
$(4)^{th}$ n = 4	18	Long period
$(5)^{th}$ n = 5	18	Long period
$(6)^{th}$ n = 6	32	Very long period
$(7)^{th}$ n = 7	19	Incomplete period

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

- Mendeleev's predicted the properties of those missing elements from the known properties of the other elements in the same group.
- Eka - aluminium and Eka-silicon names were given for gallium and germanium (not discovered at the time of Mendeleev's). Later on it was found that properties predicted by Mendeleev's for these elements and those found experimentally were almost similar.

Property	Eka-aluminium (predicted)	Gallium (found)	Eka-silicon (predicted)	Germanium (found)
Atomic Mass	68	70	72	72.6
Density/(g/cm^3)	5.9	5.94	5.5	5.36
Melting point (K)	Low	30.2	High	1231
Formula of oxide	E_2O_3	Ga_2O_3	EO_2	GeO_2
Formula of chloride	ECl_3	$GaCl_3$	ECl_4	$GeCl_4$

Table: Mendeleev's Predictions for the Elements Eka-aluminium (Gallium) and Eka-silicon (Germanium)

Merits of Mendeleev's Periodic table

He has simplified and systematised the study of elements and their compounds.

He has helped in predicting the discovery of new elements on the basis of the blank spaces given in its periodic table.

Demerits in Mendeleev's Periodic Table

- Position of hydrogen is uncertain. It has been placed in IA and VIIA groups because of its resemblance with both the groups.
- No separate positions were given to isotopes.
- Anomalous positions of lanthanides and actinides in periodic table.

MODERN PERIODIC LAW AND PRESENT FORM OF PERIODIC TABLE

Moseley studied (1909) the frequency of the X-ray produced by the bombardment of a strong beam of electrons on metal target. He found that the square root of the frequency of X-rays $\left(\sqrt{v}\right)$ is directly proportional to number of effective nuclear charge (Z) of metal i.e. to atomic number and not to atomic mass of the atom of that metal (as nuclear charge of metal atom is equal to atomic number), i.e. $\left(\sqrt{v}\right)$ = a (Z - b).

Where 'a' and 'b' are constant. Thus, he, concluded that atomic number was a better fundamental property of an element than its atomic weight. **Then he suggested that the atomic number (Z) instead of atomic weight should be basis of the classification of the elements.**

Modern Periodic Law (Moseley's Periodic Law)

Physical and chemical properties of the elements are the periodic functions of their atomic number.

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

Periodicity

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

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

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

Periods

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

- Each period consists of a series of elements having same valence shell.
- Each period corresponds to a particular principal quantum number of the valence shell present in it.
- Each period starts with an alkali metal having outermost electronic configuration as ns^1.

Groups

- There are eighteen groups numbered as 1, 2, 3, 4, 5, ... 13, 14, 15, 16, 17, 18.
- Group consists of a series of elements having similar valence shell electronic configuration.

IUPAC NOMENCLATURE FOR ELEMENTS WITH Z $>$100

- Nomenclature of elements CNIC (commission on nomenclature of inorganic chemistry) appointed by IUPAC in 1994, approved a nomenclature scheme as well as also gave official names for elements after Z > 100 (upto atomic number 104 to 109 discovered by that time). This nomenclature is to be followed for naming the elements until their names are officially recognised.
- The names are derived by using roots for the three digits in the atomic number of the element and adding "ium" at the end. The roots for the numbers are.

Table: Notation for IUPAC Nomenclature of Elements

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

Table: Name and Symbols in current Use (or proposed) for Trans-fermium Elements (Z=101-118)

Atomic number	Systematic 1977	IUPAC 1997
101	Unnilunium (Unu)	Mendelevium (Md)
102	Unnilbium (Unb)	Nobelium(No)
103	Unniltrium(Unt)	Lawrencium(Lr)
104	Unnilquadium(Unq)	Rutherfordium(Rf)
105	Unnipentium(Unp)	Dubnium(Db)
106	Unnilhexium (Unh)	Seaborgium(Sg)
107	Unnilseptium(Uns)	Bohrium(Bh)
108	Unniloctium(Uno)	Hassium (Hs)
109	Unnilennium(Une)	Meitnerium(Mt)
110	Ununnillium(Uun)	Darmstadtium(Ds)
111	Unununium(Uuu)	Rontgenium(Rt)
112	Ununbium(Uub)	Copernicium (Cn)
113	Ununtrium(Uub)	Nihonium(Nn)
114	Ununquadium(Uuq)	Flerovium (Fl)
115	Ununpentium(Uup)	Moscovium(Mc)
116	Ununhexium(Uuh)	Livermorium (Lv)
117	Ununseptium(Uus)	Tennessine(Ts)
118	Ununoctium(Uuo)	Oganesson(Og)

CLASSIFICATION OF THE ELEMENTS

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

(a) s-block elements: When shells upto (n – 1) are completely filled and the last electron enters the *s*-orbital of the outermost (n^{th}) shell, the elements of this class are called **s-block elements**.

- Group 1 & 2 elements constitute the *s*-block.
- General electronic configuration is ns^{1-2}
- *s*-block elements lie on the extreme left of the periodic table.
- Includes metals.

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

- Group 13 to 18 elements constitute the *p*-block.
- General electronic configuration is $ns^2\ np^{1-6}$
- *p*-block elements lie on the extreme right of the periodic table.

- This block includes some metals, all non metals and metalloids.
- *s*-block and *p*-block elements are collectively called **main group or representative elements**.

(c) d-Block elements

When outermost (n^{th}) and penultimate shells ($n–1$)th shells are incompletely filled and last electron enters the ($n–1$) d orbitals (i.e., *d*-orbital of penultimate shell) then elements of this class are called *d*-block elements.

- Group 3 to 12 elements constitute the *d-block*.
- General electronic configuration is ($n–1$) d^{1-10} ns^{0-2} (except, palladium which has valence shell electronic configuration $4d^{10}$ $5s^0$).
- *d*-block elements are classified into four series.
- Those elements which have partially filled d-orbitals in neutral state or in any stable oxidation state are called transition elements.
- All the transition elements are metals and most of them form coloured complexes or ions.

(d) *f*-Block elements

When n, ($n–1$) and ($n–2$) shells are incompletely filled and last electron enters into *f*-orbital of antipenultimate i.e., ($n–2$)th shell, elements of this class are called ***f*-block elements**.

- General electronic configuration is $(n-2)\,f^{1-14}\,(n-1)\,d^{0-1}\,ns^2$
- The elements coming after uranium are called transuranium elements.
- They are also called as inner-transition elements as they contain three outer most shell in complete and were also referred to as **rare earth elements** since their oxides were rare in earlier days.

The elements of *f*-blocks have been classified into two series.

1. Ist inner transition or *4 f*-series, contains 14 elements $_{58}Ce$ to $_{71}Lu$. Filling of electrons takes place in 4f subshell.
2. IInd inner transition or *5 f-series*, contains 14 elements $_{90}$Th to $_{103}$Lr. Filling of electrons takes place in *5f* subshell.

- The actinides and lanthanides have been placed at the bottom of the periodic table to avoid the undue expansion of the periodic table.

> ✍ **KEY NOTE**
>
> **Prediction of period, group and block**
>
> - Period of an element corresponds to the principal quantum number of the valence shell.
> - The block of an element corresponds to the type of subshell which receives the last electron.
> - The group is predicted from the number of electrons in the valence shell or/and penultimate shell as follows.
>   (*a*) For s-block elements, Group number = the number of valence electrons
>   (*b*) For p-block elements, Group number = 10 + number of valence electrons
>   (*c*) For d-block elements, Group number = number of electrons in (n – 1) d sub shell + number of electrons in valence shell.

TRAIN YOUR BRAIN

Q. Elements A, B, C, D and E have the following electronic configurations :

A : $1s^2\,2s^2\,2p^1$ B : $1s^2\,2s^2\,2p^6\,3s^2\,3p^1$

C : $1s^2\,2s^2\,2p^6\,3s^2\,3p^3$ D : $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$

E : $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2$

Which among these will belong to the same group in the periodic table ?

Ans. Out of these, elements A and B will belong to the same group of the periodic table because they have same outer electronic configuration, $ns^2\,np^1$.

Q. 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 ?

Ans. (*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., $7s^2$).

PERIODIC TRENDS IN PROPERTIES OF ELEMENT

Trends in Physical Properties

(i) Effective Nuclear Charge: In an polyelectronic atom between the outer most valence electrons and the nucleus of an atom, there exists number of shells containing electrons. Due to the presence of these inner electrons, the valence electrons are unable to experience the attractive pull of the actual number of protons in the nucleus. These inner electrons act as shield between the valence electrons and protons in the nucleus.

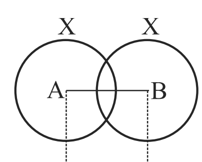
Thus, the presence of intervening (shielding) electrons reduces the electrostatic attraction between the protons in the nucleus and the valence electrons because intervening electrons repel the valence electrons.

The effective nuclear charge (Z_{eff}) is the charge actually felt by the valence electron. Z_{eff} is given by $Z_{eff} = Z - \sigma$, (where Z is the actual nuclear charge (atomic number of the element) and σ is the shielding (screening) constant).

(ii) Atomic radius: Atomic radius is taken as the effective size that is the distance of the closest approach of one atom to another atom in a given bonding state.

Types of Atomic Radius

(a) Covalent radius : It is one-half of the distance between the centres of two nuclei (of like atoms) bonded by a single covalent bond as shown in figure.



$$r_{covalent} = \frac{1}{2} AB\,(of\ \ element\ \ x)$$

Covalent radius is generally used for non metals.

Single Bond Covalent Radius, SBCR (bond length)

(*a*) For **homodiatomic molecules** $d_{A-A} = r_A + r_A$ or $2r_A$

so, $r_A = \frac{a_{A-A}}{2}$

(*b*) For **heterodiatomic molecules** in which electronegativity remains approximately same.

$d_{A-B} = r_A + r_B$

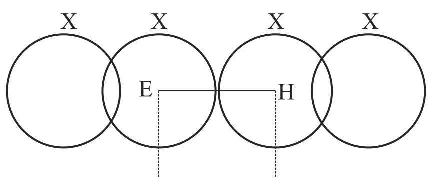
For heteronuclear diatomic molecule, A–B, where difference between the electronegativity values of atom A and atom B is relatively larger,

$$d_{A-B} = r_A + r_B - 0.09\,\Delta\chi$$

Electronegativity values are given in Pauling units and radius in picometers.

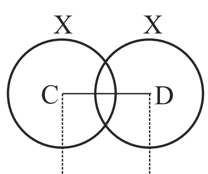
$\Delta\chi = X_A - X_B$ where X_A and X_B are electronegativity values of high electronegative element A and less electronegative element B.

Van der Waal's radius (Collision radius) : It is one-half of the internuclear distance between two adjacent atoms in two nearest neighbouring molecules of the substance in solid state as shown in figure.



$$r_{vander\ Waal's} = \frac{1}{2}EH$$

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



$$r_{crystal} = \frac{1}{2}CD(of\ \ element\ \ x)$$

Thus, the covalent, Vander Waal's and metallic radius magnitude wise follows the order,

$$\mathbf{r_{covalent} < r_{crystal} < r_{vander\ waal's}}$$

Table: Atom Radius Variation

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

The atomic radius of inert gas (zero group) is given largest in a period because it is represented by vander Waals's radius is generally larger than the covalent radius.

The Van der Waal's radius of inert gases also increases from top to bottom in a group.

In the transition series (e.g. in first transition series), the covalent radii of the elements decrease from left to right across a row until near the end when the size increases slightly.

The radii of the elements from Cr to Cu, are very close to one another as the successive addition of *d*-electrons screen the outer electrons (4s) from the inward pull of the nucleus. As a result of this, the size of the atom does not change much in moving from Cr to Cu.

There are 14 lanthanide elements between lanthanum and hafnium, in which the antipenultimate 4f shell of electrons (exert very poor shielding effect) is filled. There is a gradual decrease in size of the 14 lanthanide elements from cerium to lutetium. This is called **lanthanide contraction**. This lanthanide contraction cancels out the normal size increase on descending a group in case of transition elements.

(iii) Ionic radius: It is the effective distance from the centre of nucleus of the ion that is cation /anion up to which it has an influence in the ionic bond is know ionic radius.

	Cl^- (ionic radius 1.84 Å)	Cl (covalent radius 0.99 Å)
Number of electrons :	17	18
Number of protons :	17	17

So, there is reduction in effective nuclear charge and hence the electron cloud expands in case of Cl^-.

The sizes of ions increases as we go down a group (considering the ions of same charge). For example :

Li^+ (0.76) < Na^+ (1.02) < K^+ (1.38) < Rb^+ → (in Å)

$Be^{2+} < Mg^{2+} < Ca^{2+} < Sr^{2+}$

$F^- < Cl^- < Br^- < I^-$

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

Within a series of isoelectronic species as the nuclear charge increases, the force of attraction by the nucleus on the electrons also increases. So, the ionic radii of isoelectronic species decrease with increases in the magnitude of nuclear charges. For example, as shown in figure

$$\xrightarrow{\quad Al^{3+} \quad Mg^{2+} \quad Na^{+} \quad F^{-} \quad O^{2-} \quad N^{3-} \quad}$$

Ionic radii increase

$$\xrightarrow{\qquad\qquad\qquad\qquad\qquad\qquad}$$

As effective nuclear charge decrease.

- Following are the examples of isoelectronic series
  (*i*) S^{2-}, Cl^-, K^+ , Ca^{+2} , Sc^{+3} (*ii*) SO_2, NO_3^-, CO_3^{2-},
  (*iii*) N_2 , CO, CN^- (*iv*) NH_3 , H_3O^+
- Pauling's empirical formula for ionic radius

  $\propto \dfrac{1}{nuclear\ charge}$ (only for isoelectronic species)

TRAIN YOUR BRAIN

Q. X – X bond length is 1.00 Å and C–C bond length is 1.54 Å. If electronegativities of X and C are 3.0 and 2.0 respectively, then C–X bond length is likely to be ? (using Stevenson & Schomaker formula).

Ans. $r_{C-X} = r_C + r_X - 0.09\,\Delta\chi$

$$= \frac{1.00}{2} + \frac{1.54}{2} - 0.09\left[\Delta\chi = 1\right] = 1.27 - 0.09$$

C–X bond length = 1. 18 Å.

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

Ans. Mg^{2+} and O^{2-} both are isoelectronic i.e., have same number of electrons. But Mg^{2+} having 12 protons in its nucleus apply a 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-}.

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

$$M(g) \xrightarrow{E_1} M^+(g) + e^- : M^+(g) + IE_2 \rightarrow M^{2+}(g) + e^-$$

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

IE_1, IE_2 & IE_3 are the Ist, IInd & IIIrd ionization energies to remove electron from a neutral atom, monovalent and divalent cations respectively.

In general, $(IE)_1 < (IE)_2 < (IE)_3 <$ ...., as the number of electrons decreases, the attraction between the nucleus and the remaining electrons increases considerably and thus subsequent ionization energies increase.

- It is measured in kJ mol^{-1}, k Cal mol^{-1}, eV (electron volt).
- **Factors Influencing Ionisation energy**

  It is influenced by the following factors.

(a) Size of the Atom: Ionisation energy decreases with increase in atomic size.

As the distance between the outermost electrons and the nucleus increases, the force of attraction between the valence shell electrons and the nucleus decreases. As a result, outer most electrons are held loosely and lesser amount of energy is required to break them.

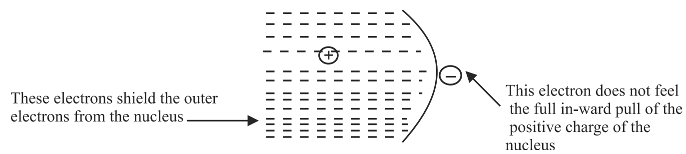
(b) Nuclear Charge: The ionisation energy increases with increase in the nuclear charge.

It is due to the fact that with increase in the nuclear charge, the electrons of the outermost shell are tightly held by the nucleus and hence large amount of energy is required to pull out an electron from the atom.

Ionisation energy increases as we move from left to right along a period due to increase in nuclear charge.

(c) Shielding or screening effect: The electrons in the innermost shells will act as a screen or shield between the nucleus and the electrons in the outermost shell. This is called **shielding effect**. The larger the number of electrons in the inner shells, greater is the screening effect and smaller the force of attraction and thus ionization energy (IE) decreases.

Screening Effect



(d) Penetration effect of the electron: The ionization energy also depends on the type of electron that is removed. *s, p, d and f* electrons have orbitals with different shapes. An *s* electron penetrates closer to the nucleus, and is therefore more tightly held than a *p* electron. Similarly *p*-orbital electron is more tightly held than a d-orbital electron and a d-orbital electron is more tightly held than an *f*-orbital electron.
If other factors being equal, penetration order are in the order $s > p > d > f$.

(e) Electronic Configuration: If an atom has exactly half-filled or completely filled orbitals, then such an arrangement has extra stability.
The removal of an electron from such an atom requires more energy then expected.

(v) Electron Affinity: The electron gain enthalpy ($\Delta_{eg}H$) is the change in standard molar enthalpy when a neutral gaseous atom gains an electron to form an anion.

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

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

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

$$E_a = E(X, g) - E(X^-, g),$$

An element has a high electron affinity if the additional electron can enter a shell where it experiences a strong effective nuclear charge.

Across a period, with increase in atomic number, electron gain enthalpy becomes more negative because left to right across a period effective nuclear charge increases and as a result it will be easier to add an electron to a small atom.

As we move in a group from top to bottom, electron gain enthalpy becomes less negative because the size of the atom increases and the added electron would be at larger distance from the nucleus.

> ✍ **KEY NOTE**
> - Group 17 elements (halogens) have very high negative electron gain enthalpies (i.e. high electron affinity) because they can attain stable noble gas electronic configuration by picking up an electron.
> - Noble gases have large positive electron gain enthalpies because the electron has to enter the next higher energy level leading to a very unstable electronic configuration.

Negative electron gain enthalpy of O or F is less than S or Cl. It is due to the fact that when an electron is added to O or F, the added electron goes to the smaller n = 2 energy state and experiences significant repulsion from the other electrons present in this level. In S or Cl, the electron goes to the larger n = 3 energy state and consequently occupies a larger region of space leading to much less electron-electron repulsion.

Nitrogen has very low electron affinity because there is high electron repulsion when the incoming electron enters an orbital that is already half filled.

Electron affinity $\propto \dfrac{1}{\textit{Atomic size}}$

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

Electron affinity $\propto \dfrac{1}{\textit{Screening effect}}$.

Stability of half filled and completely filled orbitals of a subshell is comparatively more and the addition of an extra electron to such an system is difficult and thus, the electron affinity value decreases.

TRAIN YOUR BRAIN

Q. Consider the elements N, P, O and S and arrange them in order of increasing negative electron gain enthalpy.

Ans. Order of increasing negative electron gain enthalpy is N < P < O < S.

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

The magnitude of electronegativity of an element depends upon its ionisation potential & electron affinity. Higher ionisation potential & electron affinity values indicate higher electronegativity value. When atomic size increases, the distance between nucleus and valence shell electrons increases, thus the force of attraction between the nucleus and the valence shell electrons decreases and so, the electronegativity values also decrease.

With increase in nuclear charge force of attraction between nucleus and the valence shell electrons increases and, hence, electronegativity value increases

In higher oxidation state, the element has higher magnitude of positive charge.

So, due to more positive charge on element, it has higher polarising power.

Thus, with increase in the oxidation state of element, its electronegativity also increases.

The electronegativity also increases as the *s*-character in the hybrid orbitals increases.

Hybrid orbital	sp^3	sp^2	sp
s−character	25%	33%	50%

$\xrightarrow[\textit{Electronegativity increases}]{}$

Variation of Electronegativity in a Group

On moving down the groups, Z increases but Z_{eff} almost remains constant, number of shells (n) increases, r_n (atomic radius) increases. Thus, electronegativity decreases moving down the groups.

Variation of Electronegativity in a Period

While moving across a period left to right, Z, Z_{eff} increases & r_n decreases. Thus, electronegativity increases along a period.

Measurement of Electro negativity

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

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

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

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

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

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

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

> ✍ **KEY NOTE**
>
> **Application of Electronegativity**
>
> **(a) Nomenclature**
>
> Name of more electronegative element is written at the end and 'ide' is suffixed to it. The name of less electronegative element is written before the name of more electronegative element of the formula. For example-
>
> | Correct formula | Name |
> |---|---|
> | (a) $I^+ Cl^-$ | Iodine chloride |
>
> **(b) Partial Ionic Character in Covalent bonds**
>
> Partial ionic characters are induced in covalent compounds by the difference of electronegativities. **Hanney and Smith** calculated percentage of ionic character from the difference of electronegativity.
>
> Percentage of ionic character
>
> $= 16(X_A - X_B) + 3.5(X_A - X_B)^2 = 16\Delta + 3.5\Delta^2$
>
> $= (0.16\Delta + 0.035\Delta^2) \times 100$
>
> X_A is electronegativity of element A (Higher)
>
> X_B is electronegativity of element B (Lower)
>
> $\Delta = X_A - X_B$

TRAIN YOUR BRAIN

Q. If electronegativity of x be 3.2 and that of y be 2.2, the percentage ionic character of xy is –

(*a*) 19.5 (*b*) 18.5

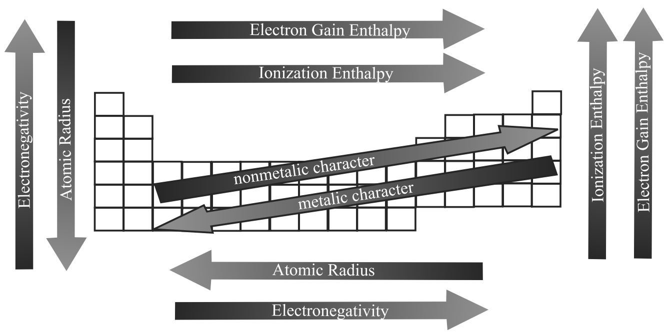
(*c*) 9.5 (*d*) 29.5.

Ans. (a) $EN_x - EN_y = 3.2 - 2.2 = 1$.

$\Delta = 1$

[Δ = difference of electronegativity values between x and y]

% ionic character = $16\Delta + 3.5\Delta^2 = 19.5$.



The periodic trends of elements in the periodic table

Periodic Trends in Chemical Properties

1. Periodicity of Valence or Oxidation States: The valence of representative elements is usually (though not necessarily) equal to the number of electrons in the outermost orbitals and / or equal to eight minus the number of outermost electrons. Frequently used for valence.

Being highest electronegative element, fluorine is given oxidation state –1.

Since there are two fluorine atoms in this molecule, oxygen with outer electronic configuration $2s^2 2p^4$ shares two electrons with fluorine atoms and thereby exhibits oxidation state +2.

On the other hand sodium with electronic configuration $3s^1$ loses one electron to oxygen and is given oxidation state +1.

There are many elements which exhibit variable valence. This is particularly characteristic of transition elements and actinoids.

2. Anomalous properties of Second period Elements: Some elements of certain groups of 2nd period resemble much in properties with the elements of third period of next group i.e. elements of second and third period are diagonally related in properties. This phenomenon is known as **diagonal relationship**. For example, the similarity between lithium (the first member of group 1) and magnesium (the second element in group 2) is called a diagonal relationship. Diagonal relationship also exist between other pairs of elements Be and Al, B and Si as shown in figure ;

Diagonal Relationship

2nd period	Li	Be	B	C
3rd period	Na	Mg	Al	Si

PERIODIC TRENDS AND CHEMICAL REACTIVITY

The ionization enthalpy of the extreme left element in a period is the least and the electron gain enthalpy of the element on the extreme right is the highest negative (except noble gases which having completely filled shells have rather positive electron gain enthalpy values). This results in high chemical reactivity at the two extremes and the lowest in the centre.

So, the maximum chemical reactivity at the extreme left (among alkali metals) is represented by the loss of an electron leading to the formation of cation and at the extreme right (among halogens) represented by the gain of an electron forming an anion.

The loss and gain of electron can be related with the reducing and oxidizing behaviour of the elements respectively. Therefore, it can also be directly related to the metallic and nonmetallic character of elements.

Thus, the metallic character of an element, that is highest at the extremely left decreases and the nonmetallic character increases while moving from left to right across the period.

The chemical reactivity of an element will be best understood by its reactions with oxygen and halogens.

Nature of oxide

Elements on two extremes of a period easily combine with oxygen to form oxides.

The normal oxide formed by the element on extreme left is the most basic (e.g. Na_2O), whereas that formed by the element on extreme right is the acidic (e.g. Cl_2O_7).

Oxides of elements in the centre are amphoteric (e.g. Al_2O_3, As_2O_3) or neutral (e.g., CO, NO, N_2O).

Amphoteric oxides behave as acidic with bases and as basic with acids, while neutral oxides have no acidic or basic properties.

In general, metallic oxides (O^{2-}), peroxides (O_2^{2-}) and super oxides (O_2^{1-}) are ionic solids.

Oxides of IA and IIA dissolve in water forming basic solutions where as other oxides do not dissolve in water.

$$Na_2O + H_2O \rightarrow 2NaOH$$

In a group, basic nature of oxides increases or acidic nature decreases.

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

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

Nature of Hydroxide

If electronegativity of E and O is larger than H and O in H_2O then EOH is basic.

$$E—O—H + H_2O \rightarrow [EOH_2]^+ + OH^-$$

If electronegativities of E and O is less than that of H and O in H_2O then EOH is acidic due to the formation of H_3O^+.

$$E—O—H + H_2O \rightarrow H_3O^+ + EO^-$$

TRAIN YOUR BRAIN

Q. Arrange the following in decreasing basic nature LiOH, NaOH, KOH, CsOH.

Ans. The basic nature of hydroxides of elements of group 1st increases on descending the group with increase in size of cation as CsOH > RbOH > KOH > NaOH > LiOH.

Topicwise Questions

GENESIS OF PERIODIC TABLE, MODERN PERIODIC LAW & NOMENCLATURE OF ELEMENTS

1. The third period of periodic table contains
(*a*) 8 elements (*b*) 32 elements
(*c*) 3 elements (*d*) 18 elements

2. The element Californium belongs to the family of
(*a*) Actinoids series (*b*) Alkali metals
(*b*) Lanthanoid series (*d*) Alkali Earth metals

3. The tenth element in the periodic table resembles the element with atomic number
(*a*) 2 as well as 30 (*b*) 2 as well as 54
(*c*) 8 as well as 18 (*d*) 8 only

4. Without looking at the periodic table select from each of the following list, the elements belonging to the same group
(*a*) Z = 12, 38, 4, 88 (*b*) Z = 9, 16, 3, 35
(*c*) Z = 5, 11, 27, 19 (*d*) Z = 24, 47, 42, 55

5. The plot of square root of frequency of X-rays emitted against atomic number led to the suggestion of which law (rule)?
(*a*) Mendeleev's periodic law
(*b*) Modern periodic law
(*c*) Hund's rule
(*d*) Newland's law

6. The discovery of which of the following group of element gave death to the Newland law of octave?
(*a*) Inert gas (*b*) Alkaline earth metal
(*c*) Rare earth (*d*) Actinoid series

7. The long form of periodic table was based on.
(*a*) Atomic number (*b*) Atomic mass
(*c*) Atomic volume (*d*) Effective nuclear charge

8. What is the name of element with atomic number 105?
(*a*) Kurchatovium (*b*) Dubnium
(*c*) Nobelium (*d*) Holmium

ELECTRONIC CONFIGURATION

9. The electronic configuration, element of group 18 can be represented by.
(*a*) ns^2np^5 (*b*) ns^1
(*c*) ns^2np^6 (*d*) $(n-1)d^8\ ns^2$

10. In transition element , the incoming electron occupies (n-1)d subshell, in preference to.
(*a*) *np*-level (*b*) *ns*-level
(*c*) (*n-1*) p-level (*d*) (n+1) s-level

11. Electronic configuration of the 4th transitional element is:
(*a*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$ (*b*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^4\ 4s^2$
(*c*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^4$ (*d*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$

12. Which of the following ions does not have the configuration of Argon?
(*a*) Cl^- (*b*) K^+
(*c*) Ca^{2+} (*d*) I^-

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

14. Name an element of p-block of the periodic table in which last electron goes to the s-orbital of valence shell instead of p-orbital:
(*a*) As
(*b*) Ga
(*c*) No such element is there
(*d*) He

15. Find the total no. of d-block elements in the following Atomic numbers?
48, 28, 70, 100, 55, 45, 34, 36
(*a*) 5 (*b*) 3
(*c*) 6 (*d*) 2

16. Which of the following electronic configuration 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*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$

17. Which block of the periodic table contains the element with configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$?
(*a*) s-block (*b*) p-block
(*c*) d-block (*d*) f-block

18. The electronic configuration of gadolinium (atomic number 64) is:
(*a*) $[Xe]\ 4f^8\ 5d^9\ 6s^2$ (*b*) $[Xe]\ 4f^7\ 5d^1\ 6s^2$
(*c*) $[Xe]\ 4f^3\ 5d^5\ 6s^2$ (*d*) $[Xe]\ 4f^6\ 5d^2\ 6s^2$

19. Which of the following will have total number of d-electrons equal to the difference in the number of total p and s electrons?
(*a*) He (*b*) Ne
(*c*) Ar (*d*) Kr

20. The radii of the F, F^-, O and O^{2-} are in the order:
(*a*) $O^{2-} > O > F^- > F$ (*b*) $F^- > O^{2-} > F > O$
(*c*) $O^{2-} > F^- > F > O$ (*d*) $O^{2-} > F^- > O > F$

21. Which of the following configuration represent atoms of element having the highest second ionisation energy?
(*a*) $1s^2\ 2s^2\ 2p^4$ (*b*) $1s^2\ 2s^2\ 2p^6$
(*c*) $1s^2\ 2s^2\ 2p^6\ 3s^1$ (*d*) $1s^2\ 2s^2\ 2p^6\ 3s^2$

22. Which of the following ions are paramagnetic in character?
(*a*) Zn^{2+} (*b*) Cu^+
(*c*) Ni^{2+} (*d*) Ag^+

23. Which of the following is not correct for isoelectronic ions?
(*a*) They have the same number of electrons around their nuclei
(*b*) Higher the atomic number, higher will be positive charge in a series of isoelectronic ions of same period
(*c*) Isoelectronic ions have same electric charge
(*d*) An isoelectronic series may have both positively and negatively charged ions

24. Which of the following is correct?
(*a*) $r_{ionic} \propto Z$ (*b*) $r_{ionic} \propto Z_{eff}$
(*c*) $r_{ionic} \propto \dfrac{1}{Z_{eff}}$ (*d*) $r_{ionic} \propto Z_{eff}^2$

25. Pd has exceptional electronic configuration of $4d^{10}\ 5s^0$. It belong to
(*a*) 4^{th} period, group 11 (*b*) 5^{th} period, group 10
(*c*) 6^{th} period, group 9 (*d*) 3^{rd} period, group 16

26. Atomic number of Ag is 47. In the same group, the atomic numbers of elements placed above and below Ag in long form of periodic table will be
(*a*) 29, 65 (*b*) 39, 79
(*c*) 29, 79 (*d*) 39, 65

PERIODIC TRENDS AND PROPERTIES OF THE ELEMENTS

27. Which of the following isoelectronic ions has lowest ionisation energy?
(*a*) K^+ (*b*) Ca^{2+}
(*c*) Cl^- (*d*) S^{2-}

28. The elements which occupy the peaks of ionisation energy curve are
(*a*) Na, K, Rb, Cs (*b*) Na, Mg, Cl, I
(*c*) Cl, Br, I, F (*d*) He, Ne, Ar, Kr

29. Which out of the following has the largest ionisation energy?
(*a*) $_{11}Na$ (*b*) $_{19}K$
(*c*) $_{12}Mg$ (*d*) $_{37}Rb$

30. Which of the process requires largest energy among them?
(*a*) $Al(g) \rightarrow Al^+(g) + e^-$
(*b*) $Al^{2+}(g) \rightarrow Al^{3+}(g) + e^-$
(*c*) $Al^+(g) \rightarrow Al^{2+}(g) + e^-$
(*d*) All the processes require same amount of energy

31. Which of the following process refers to the ionisation potential?
(*a*) $X(s) \rightarrow X^+(g) + e^-$ (*b*) $X(g) + 2e^- \rightarrow X^{2-}(g)$
(*c*) $X(g) \rightarrow X^+(g) + e^-$ (*d*) $X(g) + e^- \rightarrow X^-(g)$

32. Which out of the following has the largest size?
(*a*) Rb^+ (*b*) Mg^{2+}
(*c*) Li^+ (*d*) Na^+

33. Which of the following units are used frequently for atomic radii?
(*a*) Meter (*b*) Picometers
(*c*) Kilometers (*d*) Centimeters

34. Which of the following does not affect the ionisation potential of the atom?
(*a*) Nuclear charge (*b*) Stable configuration
(*c*) Penetration effect (*d*) None of these

35. The correct order of ionic radii for the ions S^{2-}, Cl^-, P^{3-}, Ca^{2+} is
(*a*) $Ca^{2+} > Cl^- > S^{2-} > P^{3-}$ (*b*) $S^{2-} > P^{3-} > Cl^- > Ca^{2+}$
(*c*) $P^{3-} > S^{2-} < Cl^- > Ca^{2+}$ (*d*) $Ca^{2+} < Cl^- < S^{2-} < P^{3-}$

36. The correct order of electron affinity of halogens is:
(*a*) F > Cl > Br > I (*b*) I > Br > Cl > F
(*c*) Cl > F > Br > I (*d*) Cl > F > I > Br

37. Which of the following pairs show diagonal relationship?
(*a*) Li, Mg (*b*) Be, Al
(*c*) B, Si (*d*) All of these

38. Which of the following element is expected to have highest electron affinity?
(*a*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ (*b*) $1s^2\ 2s^2\ 2p^3$
(*c*) $1s^2\ 2s^2\ 2p^4$ (*d*) $1s^2\ 2s^2\ 2p^5$

39. Which of the following oxide is amphoteric?
(*a*) Na_2O (*b*) Al_2O_3
(*c*) SO_3 (*d*) P_2O_5

40. The correct order of electronegativity of N, O, F and P is.
(*a*) F > N > P > O (*b*) F > O > P > N
(*c*) F > O > N > P (*d*) N > O > F > P

41. Which set of elements have strongest tendency to form anions?
(*a*) Na, Cl, Al (*b*) Cu, Ag, Au
(*c*) Be, F, N (*d*) F, Cl, Br

42. Outer most configuration of most electronegative element of the periodic table is.
(*a*) $3s^2\ 3p^6$ (*b*) $2s^2\ 2p^5$ (*c*) $4s^2\ 4p^5$ (*d*) $2s^2\ 2p^4$

43. In which of the following pairs, the ionisation energy of the first species is less than that of the second:
(*a*) N, P (*b*) Be^{2+}, Be
(*c*) N, N^- (*d*) S, P

44. Isoelectronic ions are those which have:
(*a*) Same size
(*b*) Same ionisation energy
(*c*) Same electronic configuration
(*d*) Same nuclear charge

45. The correct order of the size of C, N, P, S following the order:

(*a*) N < C < P < S (*b*) C < N < S < P

(*c*) C < N < P < S (*d*) N < C < S < P

46. Which of the following has smallest size?

(*a*) Al^{3+} (*b*) Al^{2+}

(*c*) Al^{+} (*d*) Al

47. Which one of the following is an incorrect statement?

(*a*) The ionisation potential of nitrogen is greater than that of oxygen

(*b*) The electron affinity of fluorine is greater than that of chlorine

(*c*) The ionization potential of beryllium is greater than that of boron

(*d*) The electronegativity of fluorine is greater than that of chlorine

48. Identify the correct order of the size of the following

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

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

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

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

49. Which is mismatched regarding the position of the element as given below?

(*a*) X(Z = 89) - f block, 6th period

(*b*) Y(Z = 100) - f block, 7th period

(*c*) Z(Z = 115) - d block, 7th period

(*d*) Both (a) & (c)

50. Correct order of Ist ionization potential among elements Be, B, C, N, O is

(*a*) B < Be < C < O < N

(*b*) B < Be < C < N < O

(*c*) Be < B < C < N < O

(*d*) Be < B < C < O < N

51. An atom of an element has electronic configuration 2, 8, 1. Which of the following statement is correct?

(*a*) The valency of element is 7

(*b*) The element exists as a triatomic molecule

(*c*) The element is metalloid

(*d*) The element forms basic oxide

Learning Plus

1. The size of ionic species is correctly given in the order

(*a*) $Na^+ > Mg^{+2} > Cl^{+7} > Si^{4+}$

(*b*) $Na^+ > Mg^{+2} > Si^{4+} > Cl^{+7}$

(*c*) $Cl^{+7} > Si^{+4} > Mg^{+2} > Na^+$

(*d*) $Cl^{+7} > Na^+ > Mg^{+2} > Si^{+4}$

2. A neutral atom (A) is converted to (A^{3+}) by the following process : $A \xrightarrow[-e^-]{E_1} A^+ \xrightarrow[-e^-]{E_2} A^{2+} \xrightarrow[-e^-]{E_3} A^{3+}$

The correct order of E_1, E_2 and E_3 energies is

(*a*) $E_1 < E_2 < E_3$ (*b*) $E_1 > E_2 > E_3$

(*c*) $E_1 = E_2 = E_3$ (*d*) $E_1 > E_2 < E_3$

3. Which of the following is a favourable factor for cation formation?

(*a*) High electronegativity (*b*) High electron affinity

(*c*) Low ionisation potential (*d*) Smaller atomic size

4. Generally, the first ionisation 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.

5. Consider the isoelectronic species, Na^+, Mg^{2+}, F^- and O^{2-}. The correct order of increasing length of their radii is:

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

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

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

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

6. Which of the following is not an actinoid?

(*a*) Curium (Z = 96) (*b*) Californium (Z = 98)

(*c*) Uranium (Z= 92) (*d*) Terbium (Z= 65)

7. 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

8. The first ionisation enthalpies of Na, Mg, Al and Si are in the order:

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

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

9. The electronic configuration of gadolinium (Atomic number 64) is:

(*a*) [Xe] $4f^3\ 5d^5\ 6s^2$ (*b*) [Xe] $4f^7\ 5d^2\ 6s^1$

(*c*) [Xe] $4f^7\ 5d^1\ 6s^2$ (*d*) [Xe] $4f^8\ 5d^6\ 6s^2$

10. The statement that is not correct for periodic classification of elements is:

(*a*) The properties of elements are periodic function of their atomic numbers

(*b*) Non-metallic elements are less in number than metallic elements

(*c*) For transition elements, the 3d-orbitals are filled with electrons after 3p-orbital and before 4s-orbitals

(*d*) The first ionisation enthalpies of elements generally increase with increase in atomic number as we go along a period

11. Among halogens, the correct order of amount of energy released in electron gain (electron gain enthalpy) is:

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

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

12. 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

13. The elements in which electrons are progressively filled in 4f-orbital are called:

(*a*) Actinoids (*b*) Transition elements

(*c*) Lanthanoids (*d*) Halogens

14. Which of the following is the correct order of size of the given species:

(*a*) $I > I^- > I^+$ (*b*) $I^+ > I^- > I$

(*c*) $I > I^+ > I$ (*d*) $I^- > I > I^+$

15. The element with atomic number 57 belongs to:

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

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

16. The last element of the p-block in 6th period is represented by the outermost electronic configuration:

(*a*) $7s^2\ 7p^6$ (*b*) $5f^{14}\ 6d^{10}\ 7s^2\ 7p^0$

(*c*) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$ (*d*) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^4$

17. Which of the elements whose atomic numbers are given below, cannot be accommodated in the present set up of the long form of the periodic table?

(*a*) 107 (*b*) 118 (*c*) 126 (*d*) 102

18. The electronic configuration of the element which is just above the element with atomic number 43 in the same group is:

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

(*b*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^3\ 4p^6$

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

(*d*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$

19. The elements with atomic numbers 35, 53 and 85 are all:

(*a*) Noble gases (*b*) Halogens

(*c*) Heavy metal (*d*) Light metals

20. Electronic configuration of four elements A, B, C, and D are given below:

(*a*) $1s^2\ 2s^2\ 2p^6$ (*b*) $1s^2\ 2s^2\ 2p^4$

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

Arrange them in increasing order of electron gain enthalpies.

(*a*) A < C < B < D (*b*) A < B < C < D

(*c*) D > B < C < A (*d*) A > B < C > D

21. Which one of the following represents a d-block element?

(*a*) [Rn] $6d^{10}\ 7s^2\ 7p^2$

(*b*) [Xe] $4f^1\ 5d^1\ 6s^2$

(*c*) [Xe] $4f^{14}\ 5d^1\ 6s^2$

(*d*) [Xe] $5d^1\ 6s^2$

22. The electronic configuration having maximum difference in first and second ionization energies is

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

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

23. The successive ionization energies for element X is given below

IE_1: 250 kJ mol^{-1}

IE_2: 820 kJ mol^{-1}

IE_3: 1100 kJ mol^{-1}

IE_4: 1400 kJ mol^{-1}

Find out the number of valence electrons for the element X.

(*a*) 3 (*b*) 4 (*c*) 2 (*d*) 1

24. What is the value of electron gain enthalpy of Na^+ if IE_1 of Na = 5.1 eV?

(*a*) +2.55 eV (*b*) +10.2 eV

(*c*) –5.1 eV (*d*) –10.2 eV

25. An element X with Z = 112 has been recently discovered. To which group and period will it belong, respectively-

(*a*) 12, 7 (*b*) 10, 6

(*c*) 8, 4 (*d*) 6, 6

26. Which of the following has lowest Ist ionisation energy?

(*a*) Li (*b*) C

(*c*) O (*d*) F

Multiconcept MCQs

1. Ionisation energies of element A are given below in kJ/mol

$IE_1 = 120$, $IE_2 = 430$, $IE_3 = 540$

If A reacts with different elements, which compound of A is not possible?

(*a*) AF (*b*) A_2O

(*c*) A_3N (*d*) A_3N_2

2. In 4th period of periodic table, how many elements have one or more 4d electrons?

(*a*) 2 (*b*) 0 (*c*) 13 (*d*) 6

3. Few reactions are given below:

1. $O_{(g)} + e^- \longrightarrow O^-_{(g)}, \Delta_1H$ 2. $F_{(g)} + e^- \longrightarrow F^-_{(g)}, \Delta_2H$

3. $Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)}, \Delta_3H$ 4. $O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)}, \Delta_4H$

Which of the following statement(s) is/are correct with respect to the reactions mentioned above?

(*a*) Δ_3H is more –ve than Δ_1H and Δ_2H

(*b*) Δ_1H is less negative than Δ_2H

(*c*) Δ_1H , Δ_2H and Δ_3H are –ve whereas Δ_4H is +ve

(*d*) All are correct

4. Which of the following statement is not correct?

(*a*) Ionisation enthalpy for a neutral atom is always positive

(*b*) 2^{nd} I.E. of monovalent cation is equal to the first I.E. of bivalent cation.

(*c*) Ionisation enthalpy of an ion is always positive

(*d*) Both (a) and (b)

5. Which of the following is/are incorrect?

(*a*) I.P. of $O_{(g)}$ is less than I.P of $O^-_{(g)}$

(*b*) I.P. of $Ne_{(g)}$ is greater than I.P. of $Ne^+_{(g)}$

(*c*) I.P. of $N_{(g)}$ is greater than I.P. of $N^+_{(g)}$

(*d*) All are incorrect

6. Which of the following reactions will involve release of energy?

(*a*) $S^-_{(g)} \longrightarrow S^{2-}_{(g)}$

(*b*) $N_{(g)} \longrightarrow N^-_{(g)}$

(*c*) $\underset{(g)}{O} + \underset{(g)}{e^-} \longrightarrow \underset{(s)}{O^-}$

(*d*) $Al^{2+}_{(g)} \longrightarrow Al^{3+}_{(g)}$

7. Out of N, O, Ne, Na and Na^+, which of these species will have the maximum and minimum ionization energy respectively:

(*a*) Na^+ , O (*b*) Na^+ , Na

(*c*) N , Ne (*d*) Ne , N

8. Which of the following is/are correct with respect to reactions given below?

$$M_{(g)} \longrightarrow M^+_{(g)} + e^- ; \ \Delta H = 100 \text{ eV}$$

$$M^+_{(g)} \longrightarrow M^{2+}_{(g)} + e^- ; \ \Delta H = 250 \text{ eV}$$

(*a*) ΔH_1 of $M_{(g)}$ is 100 eV (*b*) ΔH_1 of $M^+_{(g)}$ is 150 eV

(*c*) ΔH_2 of $M_{(g)}$ is 250 eV (*d*) All are correct

9. Which of the following statement is incorrect?

(*a*) The first ionization potential of nitrogen is greater than that of oxygen

(*b*) The electron affinity of fluorine is greater than that of chlorine

(*c*) The first ionization potential of Mg is greater than aluminium

(*d*) The electronegativity of fluorine is greater than that of chlorine

10. Identify the wrong statement in the following.

(*a*) Atomic radius of the elements increases as one moves down the first group of the periodic table

(*b*) Atomic radius of the elements decreases as one moves across from left to right in the 2nd period of the periodic table

(*c*) Amongst isoelectronic species, smaller the positive charge on the cation, smaller is the ionic radius

(*d*) Amongst isoelectronic species, greater the negative charge on the anion, larger is the ionic radius

11. The diagram below is a part of the skeleton of the periodic tables in which elements are indicated by letter which are not their usual symbols.

	D	
A	C	F
52	B	E

The correct option is :

(*a*) 'D' has maximum electron affinity

(*b*) 'B' exist in nature in a liquid state

(*c*) 'A' is a 5^{th} period element in a periodic table

(*d*) (b) & (c)

12. From the following given electronic configuration, identify the correct order of electron affinity:

(I) [He] $2s^2\ 2p^5$ (II) [He] $2s^2\ 2p^3$

(III) [Ne] $3s^2\ 3p^5$ (IV) [Ne] $3s^2\ 3p^3$

(*a*) I > II > III > IV (*b*) III > I > IV > II

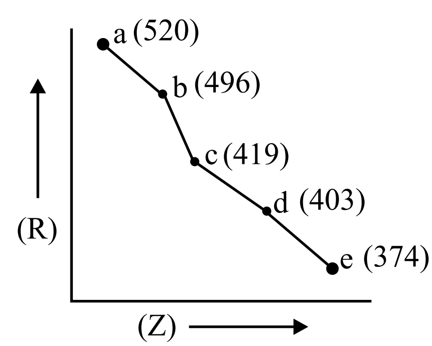
(*c*) I < II < III < IV (*d*) II > III > IV > I

13. Few elements are matched with their successive ionisation energies. Identify the elements.

Element	IE_1 (kJ/mol)	IE_2 (kJ/mol)
X	2372	5251
Y	520	7297
Z	900	1758

	X	Y	Z
(*a*)	A noble gas	Alkali metal	Alkaline earth metal
(*b*)	Alkali metal	A noble gas	Alkaline earth metal
(*c*)	Alkaline earth metal	Alkali metal	A noble gas
(*d*)	Alkali metal	Alkaline earth metal	A noble gas

14. In the given graph, a periodic property (R) is plotted against atomic numbers (Z) of the elements. Which property is shown in the graph and how is it correlated with reactivity of the elements?



(*a*) Ionisation enthalpy in a group decrease reactivity decreases from a → e.
(*b*) Ionisation enthalpy in a group decreases reactivity increases from a → e.
(*c*) Atomic radius in a group decrease reactivity decreases from a→ e.
(*d*) Metallic character in a group decrease reactivity increases from a → e.

15. Match the entries of List I with appropriate entries of List II and select the correct answer using the codes given below the lists:

	List I		**List II**
P.	Rutherfordium (At. No. = 104)	1.	Period number = 7
Q.	Roentgenium (At. No. = 111)	2.	Group number = 4
R.	Thorium (At. No. = 90)	3.	d-block elements
S.	Neptunium (At. No. = 93)	4.	f-block elements

	P	Q	R	S
(*a*)	3	2,4	2,3	1,2,3
(*b*)	1,2,3	1,3	1,4	1,4
(*c*)	1,2	2,1	4,3	3,1
(*d*)	3,1	1,3	2,4	4,1,2

16. Five ionization energy values in kJ/mol are listed below: $E_1 = 870$, $E_2 = 830$, $E_3 = 1010$, $E_4 = 1290$, $E_5 = 376$. These are
(*a*) Successive ionization energies for the element with atomic number 5
(*b*) The first I.E. of successive elements in group 15, 16, 17, 18 and 1 respectively
(*c*) The first I.E. of elements with atomic number 1 to 5
(*d*) Successive I.E. for transition elements with four electrons in d-subshell.

17. In a periodic table, the basic character of oxides
(*a*) Increases from left to right and decreases from top to bottom
(*b*) Decreases from right to left and increases from top to bottom
(*c*) Decreases from left to right and increases from top to bottom
(*d*) Decreases from left to right and increases from bottom to top

18. Identify the incorrect statement.
(*a*) The first ionisation potential of Al is less than the first ionisation potential of Mg.
(*b*) The second ionisation potential of Mg is greater than the second ionisation potential of Na.
(*c*) The first ionisation potential of Na is less than the first ionisation potential of Mg.
(*d*) The third ionisation potential of Mg is greater than that of Al.

19. Elements X, Y and Z have atomic numbers 19, 37 and 55 respectively. Which of the following statements is true about them?
(*a*) Their ionization potential would increase with increasing atomic number.
(*b*) Y would have an ionization potential between those of X and Z.
(*c*) Z would have the highest ionization potential.
(*d*) Y would have the highest ionization potential.

NEET Past 10 Years Questions

1. From the following pairs of ions which one is **not** an iso-electronic pair? ***(2021)***
(*a*) Na^+, Mg^{2+}
(*b*) Mn^{2+}, Fe^{3+}
(*c*) Fe^{2+}, Mn^{2+}
(*d*) O^{2-}, F^-

2. Identify the incorrect match ***(2020)***

Name	**IUPAC Official Name**
(A) Unnilunium	(*i*) Mendelevium
(B) Unniltrium	(*ii*) Lawrencium
(C) Unnilhexium	(*iii*) Seaborgium
(D) Unununnium	(*iv*) Darmstadtium

(*a*) (B), (*ii*) (*b*) (C), (*iii*)
(*c*) (D), (*iv*) (*d*) (A), (*i*)

3. For the second period elements the correct increasing order of first ionisation enthalpy is: ***(2019)***
(*a*) Li < Be < B < C < N < C < F < Ne
(*b*) Li < B < Be < C < O < N < F < Ne
(*c*) Li < B < Be < C < N < O < F < Ne
(*d*) Li < Be < B < C < O < N < F < Ne

4. The element Z = 114 has been discovered recently. It will belong to which of the following family group and electronic configuration? ***(2017-Delhi)***
(*a*) Nitrogen family, [Rn] $5f^{14}6d^{10}7s^27p^6$
(*b*) Halogen family, [Rn] $5f^{14}6d^{10}7s^27p^5$
(*c*) Carbon family, [Rn] $5f^{14}6d^{10}7s^27p^2$
(*d*) Oxygen family, [Rn] $5f^{14}6d^{10}7s^27p^4$

5. In which of the following options the order of arrangement does not agree with the variation of property indicated against it? ***(2016 - I)***

(*a*) $Li < Na < K < Rb$ (increasing metallic radius)

(*b*) $Al^{3+} < Mg^{2+} < Na^{+} < F^{-}$ (increasing ionic size)

(*c*) $B < C < N < O$ (increasing first ionization enthalpy)

(*d*) $I < Br < Cl < F$ (increasing electron gain enthalpy)

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

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

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

Thus, process of formation of O^{2-} in gas phase is unfavourable even though O^{2-} is isoelectronic with neon. It is due to the fact that, ***(2015 Re)***

(*a*) O^{-} ion has comparatively smaller size than oxygen atom

(*b*) Oxygen is more electronegative

(*c*) Addition of electron in oxygen results in larger size of the ion

(*d*) Electron repulsion outweighs the stability gained by achieving noble gas configuration

7. The number of d-electrons in Fe^{2+} (Z = 26) is not equal to the number of electrons in which one of the following? ***(2015)***

(*a*) p-electrons in Cl (Z = 17)

(*b*) d-electrons in Fe (Z = 26)

(*c*) p-electrons in Ne (Z = 10)

(*d*) s-electrons in Mg (Z = 12)

8. The species Ar, K^{+} and Ca^{2+} contain the same number of electrons. In which order do their radii increase? ***(2015)***

(*a*) $Ca^{2+} < Ar < K^{+}$

(*b*) $Ca^{2+} < K^{+} < Ar$

(*c*) $K^{+} < Ar < Ca^{2+}$

(*d*) $Ar < K^{+} < Ca^{2+}$

9. Be^{2+} is isoelectronic with which of the following ions? ***(2014)***

(*a*) Li^{+} (*b*) Na^{+}

(*c*) Mg^{2+} (*d*) H^{+}

10. Which of the following orders of ionic radii is correctly represented? ***(2014)***

(*a*) $Na^{+} > F^{-} > O^{2-}$

(*b*) $O^{2-} > F^{-} > Na^{+}$

(*c*) $Al^{3+} > Mg^{2+} > N^{3-}$

(*d*) $H^{-} > H^{+} > H$

11. Identify the wrong statement in the following: ***(2012 Pre)***

(*a*) Atomic radius of the elements decreases as one moves across from left to right in the 2^{nd} period of the periodic table

(*b*) Amongst isoelectronic species, smaller the positive charge on the carbon, smaller is the ionic radius

(*c*) Amongst isoelectronic species, greater the negative charge on the anion, larger is the ionic radius

(*d*) Atomic radius of the elements increases as one moves down the first group of the periodic table.

ANSWER KEY

Topicwise Questions

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

Learning Plus

1. (*b*)	**2.** (*a*)	**3.** (*c*)	**4.** (*a*)	**5.** (*b*)	**6.** (*d*)	**7.** (*a*)	**8.** (*a*)	**9.** (*c*)	**10.** (*c*)
11. (*c*)	**12.** (*c*)	**13.** (*c*)	**14.** (*d*)	**15.** (*c*)	**16.** (*c*)	**17.** (*c*)	**18.** (*a*)	**19.** (*b*)	**20.** (*a*)
21. (*d*)	**22.** (*d*)	**23.** (*d*)	**24.** (*c*)	**25.** (*a*)	**26.** (*a*)				

Multiconcept MCQs

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

NEET Past 10 Years Questions

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

Solution

Topicwise Questions

1. (*a*) According to periodic table.

2. (*a*) According to modern periodic table.

3. (*b*) Element with atomic number (2) and (54) (Xe) are also noble gas. The tenth atom is Neon (Ne). It is a noble gas element with atomic number 10.

4. (*a*) They have same valence shell configuration or their atomic no. should differ by magic number. 2, 8, 8, 18, 18, 32 .

5. (*b*) Moseley gave the modern periodic law. He showed that atomic number is more fundamental property of an element than its atomic mass. He found that the square root of the frequency of a line (of a X-Ray spectrum) is related to the atomic number (Z) of target material; as

$\sqrt{v} = a(z-b)$

6. (*a*) With the discovery of inert gases (group zero in Mendeleev's periodic table), the law of octaves lost its original significance since it was now the 9th element which had properties similar to the 1st one, which is not according to the law.

7. (*a*) The long form of periodic table is based on atomic number of element.

8. (*b*) In IUPAC nomenclature, it is known as un-nil-pentium.

Db (105) = $[Rn]^{86}\ 5f^{14}\ 6d^{3}\ 7s^{2}$

9. (*c*) Periodic table ($ns^2\ np^6$)

10. (*a*) $(n-1)d < np$

11. (*d*) 4th transition element has atomic no. 24

12. (*d*) I^- has electronic configuration of Xe.

13. (*d*) Ca^{2+} has electronic arrangement (2, 8, 8)

14. (*d*) Helium is a p-block element with configuration of s block.

15. (*b*) 28 → d-block

48 → d-block

70 → f-block

100 → f-block

55 → s-block

45 → d-block

34 → p-block

36 → p-block

16. (*d*) Electronic filling configurations is of transition elements.

17. (*c*) It is copper, a d-block transition element.

18. (*b*) Electronic filling configuration.

19. (*d*) The first inert gas which contains d electrons is in 4th period i.e., Kr with atomic no. = 36

E.C. of Kr = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

Total number of d-electrons = 10

Total number of p-electrons = 6 + 6 + 6 = 18

Total number of s-electrons = 2 + 2 + 2 + 2 = 8

Difference in p and s electron is = 18 – 8 = 10

20. (*d*) Radii are in the following order

$O^{2-} > F^- > O > F$

O^{2-} and F^- are isoelectronic and O,F belong to the same period.

21. (*c*) Electronic configuration.

22. (*c*) $Zn^{2+} : [Ar]^{18}\ 3d^{10}$; $Ag^+ : [Kr]^{36}\ 4d^{10}$

$Cu^+ : [Ar]^{18}\ 3d^{10}$; $Ni^{2+} : [Ar]^{18}\ 3d^{8}$

As such Ni^{2+} has two unpaired electrons and is paramagnetic.

23. (*c*) Isoelectronic species have same number of electrons not same isoelectric charge. For example N^{3-}, O^{2-}, F^- are isoelectronic but have different charge.

24. (*c*) On increasing effective nuclear charge ionic size decreases.

$$r_{ionic} \propto \frac{1}{Z_{eff}}$$

25. (*b*) Pd = $4d^{10}, 5s^0$ member of 4d series i.e., 5th period and 10th group.

26. (*c*) Silver belongs to 5th period. So, the atomic number of elements placed above and below will be 47 – 18 = 29 and 47 + 32 = 79 respectively.

27. (*d*) In isoelectronic species, greater the negative charge, lower the ionisation energy.

28. (*d*) Peaks in the IE curve are occupied by the noble gases because IE of noble gases is highest in the respective periods.

29. (*c*) Among Rb, K and Na, the IE_1 decrease as Rb < K < Na but IE_1 of Na < Mg.

30. (*b*) It refers to IE_3 which has a higher value.

31. (*c*) Based on definition of ionisation energy.

32. (*a*) Down the group; size increase with decrease in nuclear charge.

33. (*b*) Unit of atomic radii is picometer.

34. (*d*) None of these.

35. (*d*) $(15)P^{3-}$, $(16)S^{2-}$, $(17)Cl^{-}$, belong to 3rd period whereas 20 Ca^{2+} belong to 4th period.

These are isoelectronic ions and therefore size depends on nuclear charge. More is the nuclear charge, smaller is the size.

36. (*c*) Since the atomic size increases down the group, electron affinity generally decreases (At < I < Br < F < Cl)

37. (*d*)

	Diagonal Relationship				
	1	2	13	14	15
Period 2	Li	Be	B	C	N
Period 3	Na	Mg	Al	Si	P

38. (*a*) It represents halogen Cl. The E.A. of Cl is highest among halogens.

39. (*b*) Na_2O is basic, SO_3 and P_2O_5 are acidic but Al_2O_3 can react both with acids and bases.

40. (*c*) The correct order of electronegativities of N, O, F and P is F > O > N > P.

In a period, electronegativity increases from left to right. Hence, N < O < F.

In a group, electronegativity decreases down the group. Hence, N > P.

41. (*d*) Halogens have a high value of electron affinity also by gaining one electron, it acquire stable noble gas configuration.

42. (*b*) The most electronegative element is Fluorine.

43. (*d*) P has higher IE due to stable half filled configuration.

44. (*c*) Same number of electrons and hence same electronic configuration.

45. (*d*) Along a period size decrease while down the group size increases.

Grp 13	Grp 14	Grp 15	Grp 16
B	C	N	O
Al	Si	P	S

46. (*a*) Greater the positive charge on the ion, smaller is its size.

47. (*b*) The electron affinity of chlorine is greater than flourine.

48. (*a*) Among isoelectronic ions ionic radii of anions is more than that of cations. Further size of the anion increases with increase in negative charge and size of the cation decreases with increase in positive charge.

49. (*d*) $\begin{bmatrix} Z = 89 = Ac = d-\text{block } 7^{th} \text{ period} \\ Z = 115 = Uup = p-\text{block } 7^{th} \text{ period} \end{bmatrix}$

50. (*a*) Left to right in period I.E increases and half filled have more I.E.

1st I.E [B < Be (more I.E. due to half filled) < C < O < N (more I.E. than 'O' due to half filled)]

51. (*d*) Electronic configuration indicates that 1 e^- is present in outermost shell.

It will easily lose electrons

∴ It is metal and form basic oxide

$$2Na + \frac{1}{2}O_2 \rightarrow Na_2O$$

Learning Plus

1. (*b*) For isoelectronic more is the positive charge smaller will be the size

∴ $Na^+ > Mg^{2+} > Si^{4+} > Cl^{+7}$

2. (*a*) For a particular atom the successive ionisation potential always increases. Thus, $E_1 < E_2 < E_3$.

3. (*c*) The tendency to lose electron is higher with elements having lower ionisation potential.

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

5. (*b*) In case of isoelectronic species, more the number of protons smaller the size.

6. (*d*) Elements with atomic number, Z = 90 to 103 are called actinoids. Thus, terbium (Z = 65) is not an actinoid. Terbium belongs to lanthanoids.

7. (*a*) For the same shell, screening effect decreases in the order s > p > d > f

8. (*a*) The E.C are as follows:

Na: [Ne] $3s^1$, Mg: [Ne]$3s^2$, Al: [Ne]$3s^23p^1$, Si: [Ne]$3s^23p^2$

The I.E of Mg will be larger than that of Na due to fully filled configuration. The I.E of Al will be smaller than that of Mg due to $1e^-$ extra than the stable configuration but smaller than Si due to increase in EAN of Si.

9. (*c*) The electronic configuration of Gd (Z = 64) is [Xe] $4f^7$ $5d^1$ $6s^2$.

10. (*c*) 3d orbital is filled when 4s orbital gets completely filled. But in case of transition elements (or any elements), the order of filling of electrons in various orbital is 3p < 4s < 3d.

11. (*c*) As we move from Cl to I, the electron gain enthalpy (i.e., energy released in electron gain) become less and less negative due to a corresponding increase in the atomic size. However, the electron gain enthalpy of F is less negative than that of Cl due to its small size.

12. (*c*) As each period starts with the filling of electrons in a new principal quantum number. Thus, the period number in the long form of the periodic table refers to the maximum principal quantum number of any element in the period.

Period number = maximum n of any element (where, n = principal quantum number).

13. (*c*) The elements in which electrons are progressively filled in 4f-orbital are called lanthanoids. Lanthanoids consist of elements from Z = 58 (cerium) to 71 (lutetium).

14. (*d*) Cation is formed after the loss of electron from outer shell and anion is formed after the gain of electron to the neutral atom. Hence, cation has smaller size but anion has bigger size than its neutral atom. Thus, $I^- > I > I^+$.

15. (*c*) The element with atomic number 57 belongs to d-block element as the last electron enters the 5d-orbital against the Aufbau principle. This anomalous behaviour can be explained on the basis of greater stability of the xenon (inert gas) core.

Thus, the outer electronic configuration of La(Z = 57) is $5d^1\ 6s^2$ rather than the expected $4f^1\ 6s^2$.

16. (*c*) 6^{th} period starts with the filling of 6s-orbital and ends when 6 p-orbitals are completely filled because each period starts with the filling of electrons in a new principal energy shell.

17. (*c*) The long form of the periodic table contain element with atomic number 1 to 118.

18. (*a*) The fifth period begins with Rb (Z = 37) and ends at Xe (Z = 54). Thus, the element with Z = 43 lies in the 5th period. Since, the 4th period has 18 elements, therefore, the atomic number of the element which lies immediately above the element with atomic number 43 is 43 – 18 = 25 .

Now, the electronic configuration of element with Z = 25 is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$

19. (*b*) Elements with atomic numbers 35 (36 – 1), 53 (54 – 1), and 85 (86 – 1), lie in a group before noble gases, i.e., halogens (group 17) elements because each period ends with a noble gas. The atomic number of noble gases (i.e., group 18 elements) are 2, 10, 18, 36, 54 and 86.

Thus, the elements with atomic number 35, 53 and 85 are all belongs to halogens.

20. (*a*) Electronic configuration of elements indicate that A is a noble gas (i.e., Ne), B is oxygen (group 16), C is sodium metal (group 1) and D is fluorine (group 17).

(*i*) Noble gases have no tendency to gain electrons since all their orbitals are completely filled. Thus, element A has the least electron gain enthalpy.

(*ii*) Since, element D has one electron less and element B has two electrons less than the corresponding noble gas configuration, hence, element D has the highest electron, gain enthalpy followed by element B

(*iii*) Since, element C has one electron in the s-orbital and hence needs one more electron to complete it, therefore, electron gain enthalpy of C is less than that of element B. Combining all the facts given above the electron gain enthalpies of the four elements increase in the order A < C < B < D.

21. (*d*) Last electrons enters in d-subshell

$\therefore$ It is d-block elements.

22. (*d*) After removing 1 e^- it will get stable noble gas configuration.

23. (*d*) Difference between IE_1 and IE_2 is high then the number of valence electron in the element is one.

24. (*c*) Electron gain enthalpy is negative of I.E. i.e., –5.1 eV

25. (*a*) At No. 112 will have configuration $(Rn)86\ 5f^{14}\ 6d^{10}\ 7s^2$. It's a d–block element, belongs to 7^{th} period & 12^{th} group.

26. (*a*) Order of I^{st} IP

F > O > C > Li

Multiconcept MCQs

1. (*a*) The maximum jump is between I^{st} and 2^{nd} I.E., therefore it will exhibit oxidation state of +1.

2. (*b*) 4^{th} period comprises of 4s, 3d and 4p.
4d elements will lie in 5^{th} period

3. (*d*) For option 1, Cl has most negative electron gain enthalpy.
For option 2, F has greater negative electron gain enthalpy than O.

For option 3, adding e^- in O^- will certainly require energy, so electron gain enthalpy is +ve.

4. (*d*) Do it yourself

5. (*d*) IE_3 is always greater than IE_2 which in turn is greater than IE_1.

6. (*c*) For option a, ΔH_{eg} is always +ve.
For option b, N has half filled configuration, so $\Delta_{eg}H$ is +ve
For option d, IE will be positive
In option c, lattice energy is released

7. (*b*) Na^+ and Ne are isoelectronic but Na^+ has more protons. So, its I.E. is highest.
Na is alkali metal with a very low I.E.

8. (*d*) $M_{(g)} \longrightarrow M^+_{(g)} + e^-$; $\Delta H_1 = 100\text{ev}$

$M^+_{(g)} \longrightarrow M^{2+}_{(g)} + e^-$; $\Delta H = 150\text{ev}$

$M_{(g)} \longrightarrow M^{2+}_{(g)} + e^-$; $\Delta H_2 = 250\text{ev}$

9. (*b*) Electron affinity of F is less than chlorine because of smaller size more will be the repulsion towards new electron.

10. (*c*) In isoelectronic species i.e., same number of electrons.

$$\left\{\begin{array}{l}\text{more the positive charge; smaller will be the size}\\ \text{more the negative charge; larger will be the size}\end{array}\right\}$$

11. (*a*)

	Cl	
Se	Br	Kr
Te	I	Xe

12. (*b*) I = F   III = Cl
II = N   IV = P
order   Cl > F > P > N

13. (*a*) X has highest IE_1 and IE_2 hence, it is a noble gas. Y has low IE_1, but very high IE_2 hence, it is an alkali metal.
Z has low IE_1 than IE_2 and IE_2 is even lower than IE_2 of alkali metal hence, it is an alkaline earth metal.

14. (*b*) I.E. in a group decreases and reactivity increases.

15. (*b*) (P $\rightarrow$ 1, 2, 3)

Rf (Z = 104) : [Rn] $5f^{14}6d^27s^2$

Period no. 7, d-block element, group no. 4.

(Q $\rightarrow$1, 3)

Rg (Z = 111) : [Rn] $5f^{14}\,6d^{10}7s^1$

Period no. 7, group no. 11, d-block element.

(R $\rightarrow$1, 4)

Th (Z = 90) : [Rn] $5f^0\,6d^27s^2$

Period no. 7, group no. 3 , f-block element.

(S $\rightarrow$1, 4)

Np (Z = 93) : [Rn] $5f^4\,6d^17s^2$

Period no. 7, group no 3, f-block element.

16. (*b*) Do it yourself.

17. (*c*) As the electronegativity of element increases, acidic character of oxides increases. So, in a group, basic nature increases on moving down and decreases along a period.

18. (*b*) IE_2 of Na is higher than that of Mg because in case of Na, the second electron has to be removed from the noble gas core while in case of Mg, removal of second electron gives a noble gas core.

19. (*b*) Elements X, Y and Z with atomic numbers 19, 37, 55 lie in group 1 (alkali metals). Within a group, IE decreases from top to bottom. Therefore, IE of Y could be between X and Z.

NEET Past 10 Years Questions

1. (*c*) Total no. of e^-

$_{26}Fe \rightarrow 3d^64s^2$, $Fe^{+2} \rightarrow 3d^6$ 24

$_{25}Mn \rightarrow 3d^54s^2$, $Mn^{+2} \rightarrow 3d^5$ 23

2. (*c*) Unununnium is the element that has Atomic number = 111
IUPAC official name of Unununium: Roentgenium
Thus option (*c*) is correct.

3. (*b*) 'Be' and 'N' have comparatively more stable valence sub-shell than 'B' and 'O'.
Generally Ionisation energies increases across a period.
Thus, correct increasing order of first ionisation enthalpy is:
Li < B < Be < C < O < N < F < Ne

4. (*c*) Carbon family: [Rn] $5f^{14}6d^{10}7s^27p^2$

5. (*c*, *d*) Increasing first ionization enthalpy will be B < C < O < N.
Electron gain enthalpy: I < Br < F < Cl

6. (*d*) There is a lot of repulsion when similar charges approach each other as O^-(g) and e^- are both negatively charged. To add an electron under such situation, the force of repulsion is to be overcome by applying external energy

7. (*a*) Number of *d* electrons in Fe^{2+} (26) = 6
Number of *p* electron in Cl (Z = 17) = 11
Number of *s* electron in Mg (Z = 12) = 6
Number of *p* electron in Ne = 6

8. (*b*) In case of isoelectronic species, radius decreases with increase in nuclear charge.

9. (*a*) $Be^{2+} = 2e^-$

$Li^+ = 2e^-$

Isoelectronic species means ions with same number of electron.

10. (*b*) Cations lose electrons and are smaller in size than the parent atom, whereas anions add electrons and are larger in size than the parent atom.
Hence, the order is $H^- > H > H^+$.
For isoelectronic species, the ionic radii decreases with increase in atomic number i.e., nuclear charge.
Hence, the correct orders are:
$O^{2-} > F^- > Na^+$ and $N^{3-} > Mg^{2+} > Al^{3+}$

11. (*b*) Among isoelectronic species the ion with the maximum positive charge will have the smallest radius.