Advanced Booklet Solution

Periodic Properties

EXERCISE - 1 [A]

1. **(b)**
Law of triads states that in the set of three elements arranged in increasing order of atomic weight, having similar properties, the atomic weight of middle element is the arithmetic mean of other two.

2. **(c)**
Transition elements are those in which electron are filled in d–orbitals.

3. **(d)**
P–block elements are also known as representative element.

4. **(a)**
Covalent radii < crystal radii < Vander Waal radii.

5. **(b)**
Noble gases are placed in zero group but has 8 electrons in its valence shell.

6. **(b)**
$N_2 = 14e^- \& CO = 14e^-$

7. **(c)**
18 electrons of s, p & d-block.

8. **(c)**
Fourth period of periodic table has 3d electron filling, not 4d.

9. **(d)**
Fe^{2+} has 4 unpaired electron.

10. **(b)**
Fe^{2+} has 4 unpaired electrons, rest has zero.

11. **(d)**
C & N^+ has 6 electrons in it.

12. **(b)**
All the species has $14e^-$ in it.

13. **(b)**
$_{86}[Rn]7s^2 5f^{14} 6d^4$

14. **(c)**
$_{29}Cu : 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

15. **(a)**
Phosphorous exist as solid at room temperature.

16. **(d)**
$_{20}Ca^{2+} : 2, 8, 8$

17. **(a)**
All species are isoelectronic with argon i.e have 18 electrons in them.

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

19. **(b)**
For iso–electronic species, more negatively charged ion is more bigger in size.

20. **(c)**
For iso–electronic species, more negatively charged ion is more bigger in size.

21. **(b)**
For iso–electronic species, more negatively charged ion is more bigger in size.

22. **(d)**
For same atom more positive charged species is smaller one.

23. **(b)**
For an isoelectronic species, negatively charged species is more bigger than positively charged species.

24. **(a)**
More is the P/e ratio smaller is the radius.

25. **(b)**
$r_{Na^+} < r_{F^-}$
$r_{F^-} < r_{Cl^-}$
$r_{Cl^-} < r_{S^{2-}}$
$\therefore r_{Na^+} < r_{F^-} < r_{Cl^-} < r_{S^{2-}}$

26. **(b)**
Across a period, as Zeff increases. Atomic radii decreases.

27. **(a)**
For isoelectronic species
Radius of cation < radius of neutral
Atom < radius of anion

28. **(c)**
For same atom
$r_{cation} < r_{atom} < r_{anion}$

29. **(a)**
Na > Mg (Across a period size decreases)
Na > Na^+

$Mg > Mg^{2+}$

30. **(d)**

For isoelectronic species, more positive charged ion, smaller is the size. More is the negative charge on anion, more is the size.

31. **(d)**

$K^+ > Li^+ > Be^{2+}$

$Mg^{2+} > Be^{2+}$

$Mg^{2+} > Al^{3+}$

32. **(a)**

$Mg^{2+} < Na^+ < F^-$

Isoelectronic with Ne. For isoelectronic species, more positive charged ion, smaller is the size. More is the negative charge on anion, more is the size.

33. **(b)**

$Na^+ > Li^+$ & $Mg^{2+} > Be^{2+}$

$Na^+ > Mg^{2+}$

$Li^+ > Be^{2+}$

34. **(c)**

For isoelectronic species, more positive charged ion, smaller is the size. More is the negative charge on anion, more is the size.

35. **(c)**

It is difficult to remove s-orbital electron compared to p-orbital electron.

36. **(b)**

$Mg \longrightarrow Mg^+ + e^- \qquad \Delta H_1 = 1P_1 = 178\,\text{kcal mol}^{-1}$

$Mg^+ \longrightarrow Mg^{2+} + e^- \qquad \Delta H_2 = 1P_2 = 348\,\text{kcal mol}^{-1}$

$Mg \longrightarrow Mg^{2+} + 2e^- \qquad \Delta H = \Delta H_1 + \Delta H_2$

$= 178 + 348$

$= 526\ \text{kcal mol}^{-1}$

37. **(b)**

$\because 1P_5$ is very high i.e removal of 5th electron is very difficult. It means after removal of $4e^-$ species has acquired very stable configuration.

38. **(b)**

$I.E_n > I.E_{n-1} > I.E_{n-2} ... > I.E_2 > I.E_1$

39. **(b)**

I.E for electron in 3p > 4p

$3p^3$ is half filled so has higher IE than p^1 & p^2

40. **(b)**

$\underset{(2,8,1)}{Mg^+} \longrightarrow \underset{(2,8)}{Mg^{2+}} + e^- \left(I.E_2\right) Mg$

$\underset{(2,8)}{Na^+} \longrightarrow \underset{(2,7)}{Na^{2+}} + e^- \left(I.E_2\right) Na$

Inert gas configuration.

$\therefore (\text{I.E})_2\ \text{Na} > (\text{I.E}_2)\,\text{Mg}$

41. **(a)**
I. E is less for left side metal as their size is more, so they are more reactive.
down the group. I. E decreases & hence reducing power increases.

42. **(c)**
${}_5\text{B}: 1s^2\ 2s^2\ 2p^1$

${}_4\text{Be}: 1s^2\ 2s^2$

In B electron is removed from 2p which has lesser penetrating power than 2s.

43. **(b)**
If $(\text{I.E}_2 - \text{I.E}_1) >$ 16ev / atom, then lower oxidation state is more stable.

44. **(d)**
For isoelectronic species,

$$\text{I.E} \propto \frac{1}{\text{size}}$$

S^{2-} has maximum ionic radii & hence lowest I.E.

45. **(c)**

$$\text{I.E} \propto \frac{1}{\text{size}}$$

$\because$ size of Na is between Li & K.
Its I.E will also be in between the I.E. value of Li & K.

46. **(c)**
I.E. increases tremendously when electron is removed from noble gas configuration.
$\text{I.E}_4 >>> \text{I.E}_3$, means valence electron is 3

47. **(d)**
For alkali earth metal $(\text{I.E})_3 >> (\text{I.E})_2$.

48. **(d)**
$\text{IE of S} < \text{IE of P}$ as P has half filled 3p subshell.

49. **(b)**
I.E. : Ca > Ba, S > Se
I.E. of noble gas is highest.

50. **(c)**
Down the group I.E. decreases.

51. **(a)**

$$\text{Na} \longrightarrow \text{Na}^+ + e^- \qquad \Delta H_1 = (\text{I.E}_1)_{\text{Na}}$$

$$\text{Na}^+ + e^- \longrightarrow \text{Na} \qquad \Delta H_2 = \text{EG}_{\text{Na}^+}$$

$$|\Delta H_1| = |\Delta H_2|$$

52. **(c)**

$O \xrightarrow{+e^-} O^- \quad \Delta H_{eg_1} = (-ve)$

$O^- \xrightarrow{+e^-} O^{2-} \; \Delta H_{eg_2} = +ve$

$O \xrightarrow{+2e^-} O^{2-} \left(\Delta H_{eg}\right)_{total} = +ve$

$\therefore \Delta H_{eg_2}$ for all element is positive.

53. **(a)**

$(E.A)_O > (E.A)_N$ because N has half filled P configuration.

$(E.A)_S > (E.A)_O$ Repulsion offered to incoming

e^- in 3P orbital will be very less as compared to 2P orbital.

54. **(d)**

Negative anion species will repel incoming electron due to same charge repulsion.

55. **(b)**

Be is more stable due to fulfilled subshell. Be^- is formed over coming this stability, so more energy is taken by Be. Hence it is least stable.

56. **(c)**

EA decreases down the group

But $(E.A)_{Cl} > (E.A)_F$ Repulsion offered to incoming

e^- in 3P orbital will be very less as compared to 2P orbital.

57. **(b)**

ΔH_{eg_2} for all element is positive.

58. **(b)**

In O, electronic repulsion is very high. So EA is very less. For S & Se, repulsion difference is not significant.

$S > Se$

$\therefore$ S > Se > O

59. **(c)**

Electron affinity increases on moving right in period.

60. **(d)**

Cation will have higher EA than neutral atoms or anions.

61. **(c)**

Due to high electronegativity

$(E.N)_F > (E.N)_{Br}$

$\therefore$ Fluorine undergo reduction easily than bromine, so it is better oxidizing agent.

62. **(c)**

Electronegativity decreases down the group, but decreases is less in case of alkali metal as compared to halogen groups.

63. **(b)**

According to Pauling scale

$\chi_P = \frac{I.E + E.A}{5.6}$ where I.E & E.A is expressed in ev/atom.

64. **(d)**

$X_P = \frac{X_m}{2.8}$ X_P: electronegativity on Pauling scale

X_m: electronegativity on Mulliken's scale.

65. **(c)**

Electronegativity increases along a period and decreases down the group.

66. **(a)**

$X_M = \frac{I.P + E.A}{2}$ I.P & E.A in eV/atom.

67. **(a)**

More is the electronegativity difference, more is the ionic character. For same atom, electronegativity ≈ 0.

68. **(d)**

$$\text{Hydration energy} \propto \frac{\text{charge}}{\text{size}}$$

69. **(a)**

$$\text{Hydration energy} \propto \frac{\text{charge}}{\text{size}}$$

70. **(d)**

Electronegativity difference increases, it means acidic strength of different element decreases.

71. **(c)**

B_2O_3 is acidic & Al_2O_3 is amphoteric.

72. **(a)**

In a period, from left to hight, atomic radius decreases while E. A increases.

73. **(d)**

In alkali metal, I. E. decreases down the group.
So, it oxidises easily. Hence reactivity increases down the group.
In halogens, EA$\downarrow$ down the group, so reactivity decreases.

74. **(c)**

${}_6C^+ : 1s^2 2s^2 2p^1$

${}_7N^+ : 1s^2 2s^2 2p^2$

${}_8O^+ : 1s^2 2s^2 2p^3 \leftarrow$half filled subshell is more stable.

${}_9F^+ : 1s^2 2s^2 2p^4$

$\therefore (I.E)_2$

$O > F > N > C$.

75. **(b)**

Hydration energy $\propto \frac{\text{charge}}{\text{size}}$

76. **(a)**

$\left(\text{I.E}_1\right)_N > \left(\text{I.E}_1\right)_O$

77. **(a)**

For F radius is covalent radius & for Ne radius vander waal radius.

78. **(c)**

Electronegativity $\propto \frac{1}{\text{size}}$

79. **(d)**

There are exception to Aufbau principle. Ex: Cr & Cu

80. **(b)**

Due to inert pair effect.

81. **(b)**

I. E $\propto \frac{1}{\text{size}}$

82. **(b)**

Size of Ga is less than Al due to 3d-orbitals in Ga.

EXERCISE - 1 [B]

1. **(d)**

Exception to aufbau's principle.

2. **(d)**

Valency of X is 1
Valency of Y is 3

3. **(d)**

Oxygen group is known as chalcogens.

4. **(b)**

$M^{3+} : [\text{Ar}]3d^{10}\,4s^2$

$M : [\text{Ar}]3d^{10}\,4s^2\,4p^3$.

5. **(b)**

${}_{26}\text{Fe} : [\text{Ar}]4s^2\;3d^6$

${}_{26}\text{Fe}^{3+} : [\text{Ar}]3d^5$

6. **(b)**

Same group elements have same outer electronic configuration.

7. **(c)**
Due to inert pair effect Bi^{3+} is more stable as compared to +3 state for others.

8. **(a)**
w.r.t oxygen
Halogen family : +7
Oxygen family : +6
Nitrogen family : +5
Boron family : +3

9. **(d)**
Valency of A is 3
Valency of B is 2
$\therefore$ compound formed is A_2B_3.

10. **(c)**
I.E increases from left to right.

11. **(b)**
$(\Delta Heg)_2$ is always positive.

12. **(d)**
Fe = Co = N; Increase in atomic number is balanced by increase in repulsion between electrons.
Ni < Cu < Zn Increase in Z_{eff} is overcome by electronic repulsion and size increases.

13. **(b)**
As P/e ratio increase, z_{eff} increases so 2nd ionization energy is always higher than 1st I. E.

14. **(c)**
For alkali metal $(I.E)_2 >> (I.E)_1$

15. **(a)**
Atomic radius is determining factor for I. E down the group z_{eff} and stable configuration important factor for I. E across period.

16. **(b)**
Smaller size of atoms, more is the I. E

17. **(b)**
Hydration energy depends on $\frac{\text{charge}}{\text{size}}$ ratio.

18. **(a)**
Acidic strength depends on $\frac{\text{charge}}{\text{size}}$ ratio.

19. **(c)**
$_{24}Cr^{2+} : [Ar]3d^4$
$_{23}V^{2+} : [Ar]3d^3$
$_{25}Mn^{2+} : [Ar]3d^5 \leftarrow$ Half-filled 'd' orbitals are more stable than other configuration.

$_{26}Fe^{2+} : [Ar]3d^6$.

20. **(b)**
More is the no. of unpaired electrons, more is the paramagnetism.

21. **(d)**
Lawrencium $_{103}Lr : [Rn]7s^2 6d^1 5f^{14}$.

22. **(c)**
In a group, atomic number differ by 8, 8, 18, 18, 32 ………. Etc.

23. **(b)**
For isoelectronic species
Anion > atom > cation

24. **(b)**

$_3Li^+ : 1s^2$

$_4Be^+ : 1s^2 2s^1$

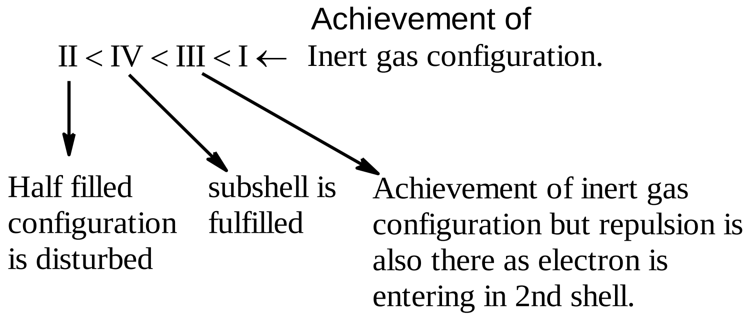
$_5B^+ : 1s^2 2s^2$.

$\therefore (I.E)_2$ order

$Be < B < Li$ or $Li > B > Be$

25. **(c)**
Difference between (I. E)$_6$ and (I. E)$_7$ is very large. So, there is $6e^-$ in valence shell of atom of this element.
$\therefore$ It belongs to 16th group.

26. **(a)**



Achievement of
$II < IV < III < I \leftarrow$ Inert gas configuration.

Half filled configuration is disturbed

subshell is fulfilled

Achievement of inert gas configuration but repulsion is also there as electron is entering in 2nd shell.

27. **(a)**

Hydrated radius $\propto \dfrac{\text{Charge}}{\text{size}}$

$\therefore K^+(aq) < Na^+(aq) < Li^+(aq)$

28. **(c)**
E.N. : P < S < Br < Cl

29. **(b)**
For alkaline, earth metal
4, 12, 20, 38, 56, 88, 120.

30. **(c)**

Less is electronegativity difference between atoms forming a bond, more is the covalent character. So

If $|X_M - X_C| > |X_O - X_H|$, the M – O bond break in polar solvent.

31. **(c)**

Moles of $X = \dfrac{110\times10^{-3}}{7}$

Energy required $= \dfrac{110\times10^{-3}}{7}\times720\,\text{kJ}$

$= 11.3\,\text{kJ}$.

32. **(a)**

$$Cl \underset{(I.E)_{Cl^-}}{\overset{(E.A)_{Cl}}{\rightleftharpoons}} Cl^-$$

$$\therefore |(E.A)_{Cl}| = |(I.E)_{Cl^-}|$$

33. **(b)**

Cl^- is more stable than F^-

$\therefore (I.E)_{Cl^-} > (I.E)_{F^-}$

$(I.E)_{Cl} > (I.E)_{Cl^-}$ - $(I.E)_{\text{neutral atom}} > (I.E)_{\text{anion}}$.

$(I.E)_F > (I.E)_{Cl}$. $I.E \ \alpha \dfrac{1}{\text{size}}$

34. **(b)**

Na^+ has inert gas configuration.

35. **(c)**

$O^- \xrightarrow{+e^-} O^{2-}$

$S^- \xrightarrow{+e^-} S^{2-}$

Incoming electron is repelled by anionic electrons.

36. **(a)**

Electronegativity $\alpha \dfrac{1}{\text{size}}$

37. **(b)**

$|X_A - X_B| \geq 2.1$ A – B is ionic bond

< 2.1 A – B is covalent bond theoretically.

Practically

$|X_A - X_B| < 1.7$ covalent bond will be formed

> 1.7 Ionic bond is formed.

38. **(b)**

According to Mulliken's scale

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

$$2X - Y = I.E$$

39. **(b)**

Hydration energy $\times \frac{\text{charge}}{\text{size}}$

40. **(d)**

Fe	Co	Ni	Cu
1.17	1.16	1.15	1.17

41. **(d)**

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

$O > F$ along period size decreases.

42. **(d)**

Be^{2+} has maximum charge / size ratio

43. **(b)**

N : $1s^2\ 2s^2\ 2p^3$ (stable configuration)

44. **(c)**

Nitrogen has half – filled 2p – orbitals.

45. **(a)**

In a period size decreases, So I. E increases. It is difficult to remove s-orbital electron compared to p-orbital electron. Therefore, I.E. of Mg > Al.

JEE Main : PYQ

1. **(d)**

Atomic no. of the element [Rn] $5f^{14}\ 6d^1\ 7s^2$

= 86 + 14 + 1 + 2 = 103

IUPAC name = Unniltrium

2. **(a)**

Period of the element just above E is 3 i.e., n = 3

Group number = 16, so orbital is p and it contains $4e^-$ in last subshell.

$E \Rightarrow [Ar]3d^{10}\ 4s^2\ 4p^4$

Element above $E \Rightarrow [Ne]3s^2\ 3p^4$

3. **(b)**

At the time of D.I. Mendeleev, structure of atom was not known.

4. **(b)**

Phosphorus has atomic number 15. Its group number is 15, number of valence electrons is 5 and valency is 3.

5. **(b)**

Elements with Z = 120 will belong to alkaline earth metals.

Its electronic configuration may be represented as $[Og]8s^2$.

6. **(a)**

If elements are arranged in order of their increasing atomic numbers, element coming at intervals of 2, 8, 8, 18, 18, 32 and 32 will have similar physical and chemical properties and thus grouped in one particular group.

7. **(c)**

I.E. order: Na < Al < Mg < Si

$\therefore$ 496 < IE (Al) < 737 < 786

Only option (c), matches the condition.

8. **(c)**

Rb and Cs have nearly same electron enthalpy $\Delta H_{eg} = -46\,\text{kJ mol}^{-1}$.

Ar and Kr have same eq. $\Delta H_{eg} = +96\,\text{kJ mol}^{-1}$.

9. **(a)**

In the periodic table; on moving down the group, the size of atom increase and tendency to remove an electrons increase. Hence, metallic character increases and on moving left to right in a period, the metallic character decreases.

Therefore; the decreasing order for metallic character is Na > Mg > Be > Si > P.

10. **(d)**

$_{31}$Ga is in group 13 with e$^-$ configuration [Ar] $3d^{10}4s^2 4p^1$ whereas $_{30}$Zn had e$^-$ configuration as [Ar] $3d^{10}4s^2$. Removal of e$^-$ from 4*p* orbital is easier than that from 4*s* orbital. Hence, I.E. of Ga is lower than that of Zn.

11. **(b)**

(A) $3s^2 \rightarrow \text{Mg}$

(B) $3s^2\,3p^1 \rightarrow \text{Al}$

(C) $3s^2\,3p^3 \rightarrow \text{P}$

(D) $3s^2\,3p^4 \rightarrow \text{S}$

All elements have $n = 3$. Along the period I.E. increase but half-filled orbital stability of P results increase in I.E.

Thus, P > S.

12. **(a)**

$$\begin{tikzcd}
\text{Li} \arrow[dr] & \text{Be} \\
\text{Na} & \text{Mg}
\end{tikzcd}$$

Diagonal relationship

Li$^+$ → Maximum hydration enthalpy in group 1 due to small size.

So, 'B' is Mg.

13. **(d)**

On moving from left to right period size of atom decrease, hence ionization energy increases.

Nitrogen have half-filled *p*-orbitals and Beryllium have fully filled *s*-orbital so their ionization energy are greater than oxygen and Boron respectively. So the order is :

$$\underset{(2p^3)}{\text{N}} > \underset{(2p^4)}{\text{O}} > \underset{(2s^2)}{\text{Be}} > \underset{(2p^1)}{\text{B}}$$

14. **(c)**

Down the group, the size of atom increases, so electron gain enthalpy decreases.

In case of oxygen, due to its small size, which leads to electron-electron repulsion, results in less electron gain enthalpy than Sulphur.

15. **(c)**

Non-metals, metals generally form acidic oxide, basic oxides respectively.

SO_3, SiO_2 → Acidic oxide

CaO → Basic oxide

Al_2O_3 → Amphoteric oxide

16. **(b)**

Halogens generally formed acidic oxides.

Cl_2O_7 – Acidic oxide

Na_2O – Basic oxide

Al_2O_3 – Amphoteric oxide

N_2O – Neutral oxide

17. **(a)**

For the isoelectronic species the ionic radius increase with increase of negative charged and ionic radius decreases with increase of positive charged.

$N^{3-} > O^{2-} > F^{-} > Na^{+} > Mg^{2+}$ (Radii)

18. **(d)**

Metalloids have intermediate properties between metals and non-metals

Sc, Pb, Bi are metals Te is a metalloid

19. **(d)**

Isoelectronic species have same no. of electrons, e.g., Al^{+3}, O^{2-}, Mg^{+2} all have 10 electrons.

20. **(b)**

The magnitude of e^- gain enthalpy decrease down the group but its value for Cl is more than that of F.

Element	F	Cl	Te	Po
$\Delta_{eg}H$ (kJ/mol)	–328	–349	–190	–174

21. **(b)**

$B(OH)_3$ and H_3PO_3 are acidic.

$Al(OH)_3$ and $Be(OH)_2$ are amphoteric.

NaOH and $Ca(OH)_2$ are basic.

22. **(b)**

$$\begin{tikzcd}
\text{Li} \arrow[dr] & \text{Be} \arrow[dr] & \text{B} \arrow[dr] & \\
\text{Na} & \text{Mg} & \text{Al} & \text{Si}
\end{tikzcd}$$

B–Si, Be–Al and Li–Mg exhibit diagonal relationship, whereas Li–Na belongs to same group.

23. **(c)**

Ionization energy generally increases moving from left to right across the period.

So, the order is Na < Mg < …… Cl < Ar.

24. **(d)**

Correct order of electron gain enthalpy is

S > Se > Te > O

Electron gain enthalpy of oxygen is even less than Te because of inter-electronic repulsion and effective nuclear charge in the oxygen atom.

25. **(b)**

Group Number

3	4	5	6	7	…	18
$_{21}\text{Sc}$				$_{25}\text{Mn}$		$_{36}\text{Kr}$
$_{39}\text{Y}$			$_{42}\text{Mo}$			$_{54}\text{Xe}$
$_{57}\text{La}$	$_{72}\text{Hf}$					

26. **(c)**

$_{90}\text{Th} \rightarrow {}_{103}\text{Lr}$

Belongs to actinoids series and they all belongs to 3rd group. So atomic no. 101 element is actinoids and atomic number 104 element belongs to 4th group.

27. **(c)**

As difference in 3rd and 4th ionisation energies is high, so atom contains 3 valence electrons.

28. **(d)**

On moving left to right along a period in the periodic table, atomic radius decreases while electronegativity, electron gain enthalpy and ionisation enthalpy increases.

29. **(c)**

Generally, non-metal oxides are acidic in nature and metal oxides are basic in nature, Al_2O_3 is amphoteric.

30. **(a)**

All given elements belongs to period III and generally their ionisation energy will increase along the period but Mg will show higher ionisation potential compared to Al due to its stable configuration. Thus, correct order of ionisation energy will be : Na < Al < Mg < Si.

Ionisation energy (kJ/mol) of the given metals are

Na : 496; Al : 577; Mg : 737; Si : 786

31. **(b)**

Generally, electron affinity decreases on moving down a group. Chlorine has more electron affinity than F because of very small size of fluorine. Therefore chlorine, Sulphur and Li has higher electron affinity among given groups.

32. **(d)**

Alkali metal have high difference in first and second ionsation energy as they achieve stable noble gas configuration after first ionisation.

33. **(a)**

On going down the group, size increases while going from left to right in a period, size decreases.

So order is : C < S < Al < Cs.

34. **(b)**

Correct order of electronegativity values of the elements is

Si > Al; S > P; Se > Te; Ge > Ga.

35. **(c)**

$_{71}X = [Xe]\ 6s^2\ 4f^{14}\ 5d^1$

$\therefore$ Orbital occupied by last e^- is 5d.

36. **(d)**

On moving from left to right across a period, the electron affinity becomes more negative. On moving from top to bottom in a group, the electron affinity becomes less negative.

Fluorine has exceptionally less negative electron affinity than chlorine, because adding an electron to fluorine (2p orbital) causes greater repulsion than adding an electron to chlorine (3p orbital) which is larger in size.

37. **(a)**

Isoelectronic species have same no. of electrons.

Ions ⇒	O^{2-}	F^-	Na^+	Mg^{2+}
	8 + 2	9 + 1	11 – 1	12 – 2
No. of e^- ⇒	10	10	10	10
Therefore	O^{2-},	F^-,	Na^+,	Mg^{2+}

are isoelectronic.

38. **(b)**

Alkali metals have the lowest ionization energy in each period, on the other hand, Sc is a d-block element.

Transition metals have smaller atomic radii and higher nuclear charge, leading to high ionisation energy.

39. **(c)**

Tellurium (Te) has $5s^2 5p^4$ valence shell configuration.

It belongs to group 16 and present in period 5 of the periodic table.

40. **(b)**

In O_2^{2-}, there are two atoms, therefore, it will be largest in size.

41. **(d)**

For chlorine, the atomic radius and covalent radius are taken as same. Hence, statement I is incorrect.

In case of anionic species, extra electrons are added in the parent atom, so effective nuclear charge decrease. Hence, anionic radius always greater than their parent atomic radius.

- **Numeric Based Questions**

42. **(101)**

$\boxed{\text{Un}}\ \boxed{\text{nil}}\ \boxed{\text{un}}$ ium

$\downarrow \quad \downarrow \quad \downarrow$

1 0 1

IUPAC symbol = Unu
Atomic no. (Z) = 101

EXERCISE - 2 [A]

1. **(b)**

Sum of oxidation numbers of all atoms in a neutral molecules is zero.

$\rightarrow A_3(B_4C)_2 : +2\times3+2\times[4\times5-2]=42\neq0$

$\rightarrow A_3(BC_4)_2 : +2\times3+2\times[+5-8]=0$

$\rightarrow A_2(BC_3)_2 : +2\times2+2\times[+5-6]=+2\neq0$

$\rightarrow ABC_2 : +2+5+(-2)\times2=+3\neq0$

2. **(d)**

W : Phosphorus	*Y*: Oxygen	*X* : Sulphur	*Z* : Chlorine
Electronegativity :	O > Cl > S > P	Catenation :	S > P > O > Cl
Electron Affinity :	Cl > O > S > P	Oxygen exhibits covalency of two only	

3. **(b)**

(a) Correct order → $Ca^{2+} > K^{+} > Cl^{-} > S^{2-}$ (Ionisation energy)
For isoelectronic species (I.E. $\propto Z_{\text{eff}}$)

(b) Correct order → C < N < F < O (2nd I.E.)
Second electron removal form oxygen requires more energy as it acquired stable $2s^2 2p^3$ configuration after removal of one electron.

(c) Correct order → B > Al > In > Ga > Al (Electronegativity)
In general EN increases in boron family from top to bottom due to increase in Z_{eff} on valence shell while boron has highest E.N. due to its vary small size.

(d) Correct order → $Na^{+} > Li^{+} > Mg^{2+} > Al^{3+} > Be^{2+}$ (Ionic radius)
Ionic radius depends on Z_{eff} and number of shells.

4. **(d)**

(i) $[Kr]5s^1$, shows only single oxidation state +1

(ii) $[Rn]5f^{14}6d^{1}7s^{2}$, it is *f*-block element (*Z* = 103)

(iii) The compound formed between I and III is ionic.

(iv) $[Ar]3d^{6}4s^{2}$, (*Z* = 26) Fe shows variable oxidation state.

5. **(c)**

	Possible elements	Period number	Group number	Electronic configuration
(i)	H	1	1	$1s^1$
(ii)	Be	2	2	$1s^2 2s^2$

6. **(b)**

As value of n increases, energy gap decreases due to increasing $Z_{\text{eff.}}$ on valence shell.

7. **(c)**

Atomic radius increases from top to bottom because number of shell increases down the group.

8. **(b)**

(a) $[\text{Ar}]3d^5 4s^1 \rightarrow \text{Cr}(24) \rightarrow$ 4th period, 6th group

(b) $[\text{Kr}]4d^{10} \rightarrow \text{Pd}(46) \rightarrow$ 5th period, 10th group

(c) $[\text{Rn}]6d^2 7s^2 \rightarrow \text{Th}(90) \rightarrow$ 7th period, 3rd group

(d) $[\text{Xe}]4f^{14} 5d^2 6s^2 \rightarrow \text{Hf}(72) \rightarrow$ 6th period, 4th group

9. **(a)**

Correct order : $\text{La}^{3+} > \text{Gd}^{3+} > \text{Eu}^{3+} > \text{Lu}^{3+}$

10. **(d)**

Second ionization energy is amount of energy required to take out an electron from the monopositive cation.

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

11. **(d)**

$A \Rightarrow \text{H}(1s^1)$

$B \Rightarrow \text{He}(1s^2)$

$C \Rightarrow \text{Li}(1s^2 2s^1)$

$A_1 = \text{IE}_1(A)$ $\quad B_2 = \text{IE}_2(B)$

$B_1 = \text{IE}_1(B)$ $\quad C_2 = \text{IE}_2(C)$

$C_1 = \text{IE}_1(C)$ $\quad C_3 = \text{IE}_3(C)$

$B_1 > A_1 > C_1$ $\quad C_3 > B_2 > A_1$ $\quad C_3 > C_2 > C_1$

$\text{He} > \text{H} > \text{Li}$ $\quad \text{Li}^{2+}\ \text{He}^+\ \text{H}$ $\quad \text{Li}^{2+}\ \text{Li}^+\ \text{Li}^+$

$1s^2\ 1s^1\ 2s^1$ $\quad 1s^1\ 1s^1\ 1s^1$ $\quad 1s^2\ 1s^2\ 2s^1$

12. **(d)**

(a) $\underset{4p^4}{\text{Se}} \xrightarrow{\text{I.E.}_1} \underset{4p^3}{\text{Se}^+} \xrightarrow{\text{I.E.}_2} \underset{4p^2}{\text{Se}^{2+}}$ $\qquad \underset{4p^3}{\text{As}} \xrightarrow{\text{I.E.}_1} \underset{4p^2}{\text{As}^+} \xrightarrow{\text{I.E.}_2} \underset{4p^1}{\text{As}^{2+}}$

(b) $\underset{2p^2}{\text{C}} \longrightarrow \underset{2p^1}{\text{C}^+} \xrightarrow{\text{I.E.}_2} \underset{2s^2}{\text{C}^{2+}}$ $\qquad \underset{2p^3}{\text{N}} \longrightarrow \underset{2p^2}{\text{N}^+} \longrightarrow \underset{2p^1}{\text{N}^{2+}}$

(c) $\underset{2p^5}{\text{F}} \xrightarrow{\text{I.E.}_1} \underset{2p^4}{\text{F}^+} \xrightarrow{\text{I.E.}_2} \underset{2p^3}{\text{F}^{2+}} \xrightarrow{\text{I.E.}_3} \text{F}^{3+}$ $\qquad \underset{2p^4}{\text{O}} \xrightarrow{\text{I.E.}_1} \underset{2p^3}{\text{O}^+} \xrightarrow{\text{I.E.}_2} \underset{2p^2}{\text{O}^{2+}} \xrightarrow{\text{I.E.}_3} \text{O}^{3+}$

(d) In respective period, noble gases have highest I.E.

13. **(c)**

R is p-block element, because difference between IE_2 and IE_3 is not very high as compared to between IE_1 and IE_2; hence stable oxidation state of R will be higher than +2.

14. **(b)**

In the isoelectronic species, all isoelectronic anions belong to the same period and cations to the next period.

15. **(d)**

Order of electron gain enthalpy : Cl > F > O

Second electron gain enthalpy for an element is always positive.

16. **(a)**

The electronegativity difference between M_1 and O is 0.1, which indicates M_1 – O bond will be covalent, since O–H bond having more ionic character thus bond will break and H^+ ions will release and acidic solution is formed. Whereas difference between electronegativity of M_2 – O bond is 2.3, thus, M_2 – OH bond will break. Hence, solution will be basic in nature.

17. **(a)**

$$\text{I.E. + E.A.} = 275 + 86 = 361 \text{ kcal mol}^{-1}$$

$$= 361 \times 4.184 = 1510.42 \text{ kJ mol}^{-1}$$

$$\therefore \quad \text{Electronegativity} = \frac{1510.42}{540} = 2.797 = 2.8$$

EXERCISE - 2 [B]

- **One or More than One Option Correct**

1. **(a, b, c, d)**

(A) S^- is more stable than O^-

$\therefore (\text{E.A})_S > (\text{E.A})_O$

(B) Z_{eff} increases.

(C) $(\text{E.A})_{Cl} > (\text{E.A})_F$

(D) 'As' belongs to nitrogen family, so it has half filled subshell.

2. **(b, d)**

Argon (Ar) has stable configuration.

3. **(a, b, c, d)**

Isotopes have same atomic number.

4. **(a, c, d)**

(A) $Z_{eff} \propto \frac{1}{\text{size}}$ E.A $\propto \frac{1}{\text{size}}$

(CD) I.E $\propto \frac{1}{\text{size}}$.

For isoelectronic species

$r_{cation} < r_{atom} < r_{anion}$.

5. **(c, d)**

$Z_{\text{eff}} \propto \frac{1}{\text{size}}$

6. **(b, d)**

As oxidation state increases acidic strength increases.

7. **(a, b, c)**

Greater energy is required to remove electrons having lower value of n.

8. **(b)**

$(\text{I.E})_3 >> (\text{I.E})_2$

+2 oxidation state of metal is more stable.

9. **(b)**

Cl has large size which results in less electronic repulsion and high electron affinity

10. **(a)**

$X \equiv Ne : 1s^2 2s^3 2p^6$.

11. **(a)**

For oxides, if electronegativity difference between elements increases, acid strength of different elements decreases. As positive change on any element increases, electronegativity increases.

12. **(a, b, c, d)**

Due to poor shielding effect of electron in f – orbital, is $r_{Nb} \approx r_{Ta}$

13. **(a, b, c)**

Halogens has most electron gain enthalpy. Alkaline earth metal has oxidation state $(+2)$

14. **(a, b, c, d)**

$\text{I.E} \propto \frac{1}{\text{size}}$

Along period half – filled shell stability is considered

15. **(b)**

Compounds of Mn	oxidation state of Mn.
MnO_2	+4
$KMnO_4$	+7
MnO	+2

As oxidation state increase size of atom decreases.

16. **(a, b, c, d)**

$\text{I.E} \propto \frac{1}{\text{size}}$

Metallic radii > covalent radii

17. **(c)**

Due to inert pair effect

18. **(c)**
Fe is transition element.

19. **(a, b, c)**
The difference between the I. P is very large when there is change in no. of shell

20. **(a, b, c)**
$\text{I.E} \propto \frac{1}{\text{size}}$ $\therefore$ along a period, I. E increases.
$\left(\text{I.E}_1\right)_{\text{Na}} < \left(\text{I.E}_1\right)_4$

21. **(d)**
Ionic mobility $\propto \frac{1}{\text{Hydrated radius}}$
Hydrated radius $\propto \frac{\text{charge}}{\text{size}}$

22. **(a, b, c, d)**
Down the 14th group, stability +4 oxidation state decreases as +2 increases.
For $\text{Pb} : \text{Pb}^{2+} > \text{Pb}^{4+}$

23. **(a, c, d)**
Exceptional configuration of Cr & Cu.

24. **(b)**
$\text{M}^{+}(\text{g}) + 280\,\text{Kcal} \longrightarrow \text{M}^{2+}(\text{g}) + \text{e}^{-}$

25. **(a, b, c)**
Fe = Co = N; Increase in atomic number is balanced by increase in repulsion between electrons.
Ni < Cu < Zn Increase in Z_{eff} is overcome by electronic repulsion and size increases.

26. **(b, c, d)**
d & f block element are between s & p block.

- **Comprehension Type**

1. **(d)**
Metals undergo oxidation easily
So I.E should be less.

2. **(d)**
Non – metals undergo reduction easily.
So | EA | should be high.

3. **(d)**
High ionisation energy.

4. **(d)**
For E (I. E_1) is less than others.

5. **(a)**
If difference in two successive Ionisation energy is less than higher oxidation state is more stable.

6. **(c)**

E.A. for noble gas is positive.

7. **(c)**

2×1^2

2×2^2

2×2^2

2×3^2

;

;

;

For 10th

Period $2\times6^2=72$

8. **(c)**

For $2e^-$ in each orbital maximum element in 4th shell is 18.

$\therefore$ For $3e^-$ in each orbital $=\frac{18}{2}\times3=27$

Elements in 4th shell.

9. **(d)**

General electronic configuration of outer most shell of f – block element is

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

10. **(c)**

For any shell / orbit no. ‘n’ maximum electron in it is $2n^2$.

$\therefore$ Maximum no. of possible elements in nature $\sum_{n=1}^{n=1}2n^2$

$$=2\sum_{n=1}^{n}n^2$$

$$=2\times\frac{1}{6}n(n+1)(2n+1)$$

$$=\frac{1}{3}n(n+1)(2n+1)$$

11. **(b)**

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

$\nu=a(z-b)^2$

12. **(d)**

For Li : $\sqrt{\nu_4}=a(3-b)$

$3\times10^8=a(3-b)$……. (1)

For Na : $\sqrt{\nu_{Na}}=a(z-b)$

$4\times10^8=a(11-b)$…………… (2)

(2) ……….. (1)

$10^8=8a$

$a=\frac{1}{8}\times10^8$

$32 = 11 - b$
$b = -21$
For K
$\sqrt{\nu} = \frac{1}{8} \times 10^{8}(40)$
$\nu = 25 \times 10^{16}$ Hz

13. **(b)**
$E = h\nu$
$= 6.626 \times 10^{-34} \times 12 \times 10^{16}$
$= 7.92 \times 10^{-17}$ J.

14. **(d)**
As electronegativity difference between element increases, acidic strength of oxide is decreases & vice versa for basic strength.

15. **(d)**
As electronegativity difference between element increases, acidic strength of oxide is decreases & vice versa for basic strength.

16. **(d)**
All are amphoteric.

17. **(d)**
SO_2 is an acidic oxide.

18. **(a)**
Covalent radius (single bonded) is given.

19. **(b)**
$r_{x^-} > (r_x)_{metallic} \cong (r_x)_{covalent} > r_{x^+}$

20. **(c)**
$r_{x^-} > (r_x)_{metallic} \cong (r_x)_{covalent} > r_{x^+}$

- **Matrix Match:**

1. $A \rightarrow s$ ${}_{19}K : [Ar]4s^1$
$B \rightarrow r$ ${}_{22}Ti : [Ar]4s^2 3d^2$
$C \rightarrow p$ ${}_{32}Ge : [Ar]4s^2 3d^{10} 4p^2$
$D \rightarrow q$ ${}_{64}Gd : {}_{54}[Xe]6s^2 5d^1 4f^7$

2. $A \rightarrow p, q, r$ ${}_{3}Li : 1s^2 2s^1$
$B \rightarrow p, q, r$ ${}_{36}[Kr]5s^1$
$C \rightarrow p, q$ ${}_{56}Ba : [Xe]6s^2$
$D \rightarrow p, q, s$ ${}_{88}Ra : [Rn]7s^2$

3. $A \rightarrow p$ ${}_{15}P : [Ne]3s^2 3p^3$

$B \rightarrow p,q,s$ $\quad {}_{31}Ga:[Ar]4s^2 3d^{10} 3p^1$
$C \rightarrow p,q,r$ $\quad {}_{35}Br:[Ar]4s^2 3d^{10} 3p^5$
$D \rightarrow p,r$ $\quad {}_{85}At:[Xe]6s^2 4f^{14} 5d^{10} 6p^5$

4. $A \rightarrow p,s$ $\quad \mu = \sqrt{n(n+2)}$ B.M n: no. of unpaired electron
$B \rightarrow p,s$ $\quad \mu = 0 \Rightarrow$ diamagnetic
$C \rightarrow p,r$
$D \rightarrow q$ $\quad \mu \neq 0 \Rightarrow$ paramagnetic

5. $A \rightarrow p,q,r$ $\quad {}_{60}Nd:[Xe]6s^2 5d^1 4f^3$
$B \rightarrow p,s$ $\quad {}_{92}U:[Rn]7s^2 6d^1 5f^3$
$C \rightarrow p,s$ $\quad {}_{90}Th:[Rn]7s^2 6d^1 5f^1$
$D \rightarrow p,s$ $\quad {}_{100}Fm$ or ${}_{100}Uun:[Rn]7s^2 6d^1 5f^{11}$

6. A → 4; B → 3; C → 1; D → 2
20 : [Ar] $4s^2$
23 : [Ar] 4s2
^{2} $3d^3$
31 : [Ar] $4s^2$ $3d^{10}$ $4p^1$
64 : [Xe] $6s^2$ $5d^1$ $4f^7$

7. A → 2; B → 1; C → 4; D → 3
$ns^{1\text{-}2}$ or ns^2 $np^{0\text{-}6}$: Main group elements (s & p-block elements)
$1s^2$ or ns^2 np^6 : Inert gas element
$(n-2)f^{14}$ $(n-1)$ d^1, ns^2 : Inner transition elements (last electron enters into f-orbital)
$(n-1)d^{1-9}$ ns^2 : Transition elements (last electron enters into d-orbital)

8. A → 3; B → 4; C → 1; D → 2
Cs : Alkali metal
As : Metalloid
Ca : Alkaline earth metals
Ru : Radioactive element

- **Integer type:**

1. **(1)**
$\mu = \sqrt{n(n+2)} = 1.732 = \sqrt{3}$
$n = 1$

2. **(4)**
Maximum oxidation state shown by any element is equal to the number of valence electrons

3. **(6)**
Same as above.

4. **(4)**
Amphoteric oxide are Al_2O_3, PbO_2, SnO_2 & BeO

EXERCISE - 2 [C]

1. For isoelectronic species, atomic radii of more negatively charged species > neutral atom > cationic species

2. Anionic radii > atomic radii
For atomic radii ; vanderwall radii > covalent radii
$\therefore \text{II} < \text{I} < \text{III}$

3. Down the group shielding effect decreases.

4. (a) $Ni^{2+} < Cu^{2+} < Zn^{2+}$
(b) For isoelectronic species
Anion > atom > cation
$\therefore\ F^{-} > Ne > Na^{+}$

5. $Ca^{2+} < Ar < Cl^{-} < S^{2-}$ (Isoelectronic species)

6. $Al^{3+} < Mg^{2+} < Li^{+} < K^{+}$

$$\left(0.51\overset{0}{A}\right)\left(0.66\overset{0}{A}\right)\left(0.68\overset{0}{A}\right)\left(1.33\overset{0}{A}\right)$$

7. $Na \underset{(E.A)^{Na+}}{\overset{(I.P)_{Na}}{\rightleftarrows}} Na^{+} + e^{-}$

$\therefore$ | I.P of Na| = |E.A of Na^{+}|

8. After first ionization, alkali metals achieves a noble gas configurations so, $IP_2 > IP_1$.
Since it is true for alkaline earth metal after 2nd ionization.
$(I.P)_3 >> (I.P)_2$

9. It is difficult to remove s-orbital electron compared to p-orbital electron.

10. Poor shielding effect of 4f subshell.

11.

$$\begin{array}{lcll} Mg & \rightarrow & Mg^{+} & + \ e^{-} \\ \frac{1}{24}\text{mole} & & \frac{1}{24}\text{mole} & \end{array}$$

$$\begin{array}{lcll} Mg^{+} & \rightarrow & Mg^{2+} & + \ e^{-} \\ \frac{1}{24} - x\text{ mole} & & x\text{ mole} & \end{array}$$

$$\frac{1}{24} \times 710 + x \times 1450 = 50$$

$$1450x = 50 - 30.83 = 19.17$$

$$x = \frac{19.17}{1450}$$

$$\% \text{ of } Mg^{2+} = \frac{x}{1/24} \times 100$$

$$= 2400x$$

$$= 240 \times \frac{19.17}{1450} = 31.73\%$$

12. $_6C:1s^2 2s^2 2p^2 \xrightarrow[1.P_1]{-e^-} {}_6C^+:1s^2 2s^2 2p^1$

More zeff $\qquad 1.P_2 \downarrow -e^-$

$_6C^{2+}:1s^2 2s^2$

$_5B:1s^2 2s^2 2p^1 \xrightarrow[(1.P)_1]{-e^-} {}_5B^+:1s^2 2s^2 \xrightarrow[(1.P)_2]{-e^-} {}_4B^{2+}:1s^2 2s^1$

Less Z_{eff} than $_6C$ $\qquad$ fully filled 2s orbital i.e stable Configuration.

13. $_4Be:1s^2 2s^2 \xrightarrow{(I.P)_1} {}_4Be^+:1s^2 2s^1 \xrightarrow{(I.P)_2} {}_4Be^{2+}:1s^2 \xrightarrow{(I.P)_3} {}_4Be^{3+}:1s^1$

Inert gas configuration.

$_5B:1s^2 2s^2 2p^1 \xrightarrow{(I.P)_1} {}_5B^+:1s^2 2s^2 \xrightarrow{(I.P)_2} {}_5B^{2+}:1s^2 2s^1 \xrightarrow{(I.P)_3} {}_5B^{3+}:1s^2$

Inert gas configuration.

14. $(I.P_1)_{Cu} < (I.P)_1\, Zn$

$(I.P_2)_{Cu} > (I.P_2)\, Zn$.

15. Be has fulfilled 2s subshell while N has half filled 2p subshell which are comparatively stable than Be^- & N^-

16. $Na_{(g)} + F(g) \longrightarrow Na^+(g) + F^-(g)$

$$\Delta H = +5.4eV - 3.4eV$$
$$= 2\,eV/atom$$

$\Delta H = 2\,eV/atom$

$= 2\times1.6\times10^{-19}\times6.023\times10^{23}\ J/mol$

$= \dfrac{2\times1.6\times10^{\cancel{19}}\times6.023\times10^{\cancel{23}}}{4.2\times\cancel{10^3}}\ Kcal/mol$

$= 45.88\,Kcal/mol$.

17. no. of moles of F ionized $= \dfrac{328}{1681}$

$\therefore$ no. of F – atoms ionized $= \dfrac{328}{1681} N_A$

18. In bonding electron of different subshell can participate.

19. $x \longrightarrow \frac{1}{2}x^+ + \frac{1}{2}x^- \qquad \Delta H_1 = -\frac{E.A}{2} + \frac{I.P}{2}$

$\frac{I.P}{2} - \frac{E.A}{2} = 409$

$\frac{1}{2}x^- \longrightarrow \frac{1}{2}x \longrightarrow \frac{1}{2}x^+$

$\frac{E.A}{2} + \frac{I.P}{2} = 733$

20. $Na < Al < Mg < Si$, the effect of half filled or fully filled subshell only extends upto neighbouring element

21. Greater the oxidation number, greater the acidic strength.

22. (I. E) K<< (I. E) Ag

23. $\underset{(2,7)}{F_{(g)}} \xrightarrow{+e^-} \underset{(2,8)}{F^-_{(g)}} \qquad \Delta H_{eg1} = -ve$

$O_{(g)} \xrightarrow{+e^-} O^-_{(g)} \qquad \Delta H_{eg1} = -ve$

$O^-_{(g)} \xrightarrow{+e^-} O^{2-} \qquad \Delta H_{eg2} = +ve$

$\because |\Delta H_{eg2}| > |\Delta H_{eg1}|$

$O \xrightarrow{+2e^-} O^{2-} \quad \Delta H_{eg} = +ve$

24. Diagonal relationship between Li and Mg.

25. Due to inert pair effect
$\therefore T\ell^{+1}$ is more stable than $T\ell^{+3}$.

26. Because it can loose $(n-1)$ d as well as ns electrons

JEE Advanced : PYQ

1. **(a, d)**
NO $\Rightarrow$ neutral
$B_2O_3 \Rightarrow$ Acidic
CrO $\Rightarrow$ Basic
All other oxides are amphoteric.

- **Assertion and reasoning**

2. **(c)**
Additional e^- are repelled more effectively by $2e^-$ in F than $3pe^-$ in Cl.

3. **(c)**
2p $(2+1=3)$ is more in energy than $2s(2+0=2)$ orbital.

4. **(c)**
The lower oxidation states are more stable due to inert pair effect.

- **Fill in the Blanks**

5. Electron gain enthalpy

6. Electronegativity $(\chi) = \frac{I.E + E.A}{2}$

7. Higher Z_{eff}
As p/e increases Z_{eff} increases.

8. Inert pair effect

- **T/F**

9. **(T)**
Density of alkali metal decreases down the group.

10. **(F)**
Less ionization potential means more reducing agent.

11. **(F)**
E. A. $Cl > F > Br$.

12. **(F)**
$|X_M - X_O|$ increases down the group, so basic nature increases.

- **Integer Value Answer**

13. **(9)**
By observing the values of ionization enthalpy for atomic number (*n* + 2), it is observed that $I_2 >> I_1$. Which shows that number of valence shell electrons is 1 for atomic number (*n* + 2). Therefore element with atomic number (*n* + 2) should be an alkali metal.

For atomic number (*n* + 3, $I_3 >> I_2$, which shows that it will be an alkaline earth metal. All the observations suggests that atomic number (*n* + 1) should be a noble gas and atomic number (*n*) should belong to halogen family. Since $n < 10$; hence $n = 9$.

14. **(2)**
Fluorine generally shows 0 and –1 oxidation states while sodium shows 0 and +1 oxidation state.