

EXERCISE - 1 [A]

- (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.
- (c)
Transition elements are those in which electron are filled in d-orbitals.
- (d)
P-block elements are also known as representative element.
- (a)
Covalent radii < crystal radii < Vander Waal radii.
- (b)
Noble gases are placed in zero group but has 8 electrons in its valence shell.
- (b)
 $N_2 = 14e^-$ & $CO = 14e^-$
- (c)
18 electrons of s, p & d-block.
- (c)
Fourth period of periodic table has 3d electron filling, not 4d.
- (d)
 Fe^{2+} has 4 unpaired electron.
- (b)
 Fe^{2+} has 4 unpaired electrons, rest has zero.
- (d)
C & N^+ has 6 electrons in it.
- (b)
All the species has $14e^-$ in it.
- (b)
 ${}_{86}[Rn]7s^2 5f^{14} 6d^4$
- (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)
 ${}_{20}\text{Ca}^{2+} : 2, 8, 8$
17. (a)
All species are isoelectronic with argon i.e have 18 electrons in them.
18. (c)
 $[\text{Xe}] 4f^7 5d^1 6s^2$
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)

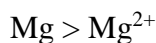
$$r_{\text{Na}^+} < r_{\text{F}^-}$$

$$r_{\text{F}^-} < r_{\text{Cl}^-}$$

$$r_{\text{Cl}^-} < r_{\text{S}^{2-}}$$

$$\therefore r_{\text{Na}^+} < r_{\text{F}^-} < r_{\text{Cl}^-} < r_{\text{S}^{2-}}$$
26. (b)
Across a period, as Z_{eff} increases. Atomic radii decreases.
27. (a)
For isoelectronic species
Radius of cation < radius of neutral
Atom < radius of anion
28. (c)
For same atom

$$r_{\text{cation}} < r_{\text{atom}} < r_{\text{anion}}$$
29. (a)
 $\text{Na} > \text{Mg}$ (Across a period size decreases)
 $\text{Na} > \text{Na}^+$



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)
 $\text{K}^+ > \text{Li}^+ > \text{Be}^{2+}$
 $\text{Mg}^{2+} > \text{Be}^{2+}$
 $\text{Mg}^{2+} > \text{Al}^{3+}$
32. (a)
 $\text{Mg}^{2+} < \text{Na}^+ < \text{F}^-$
 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)
 $\text{Na}^+ > \text{Li}^+ \text{ \& \; } \text{Mg}^{2+} > \text{Be}^{2+}$
 $\text{Na}^+ > \text{Mg}^{2+}$
 $\text{Li}^+ > \text{Be}^{2+}$
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)

$$\begin{aligned} \text{Mg} &\longrightarrow \text{Mg}^+ + \text{e}^- & \Delta H_1 = 1P_1 = 178 \text{ kcal mol}^{-1} \\ \text{Mg}^+ &\longrightarrow \text{Mg}^{2+} + \text{e}^- & \Delta H_2 = 1P_2 = 348 \text{ kcal mol}^{-1} \\ \text{Mg} &\longrightarrow \text{Mg}^{2+} + 2\text{e}^- & \Delta H = \Delta H_1 + \Delta H_2 \\ & & = 178 + 348 \\ & & = 526 \text{ kcal mol}^{-1} \end{aligned}$$
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)
 $\text{IE}_n > \text{IE}_{n-1} > \text{IE}_{n-2} \dots > \text{IE}_2 > \text{IE}_1$
39. (b)
 I.E for electron in 3p > 4p
 3p³ is half filled so has higher IE than p¹ & p²
40. (b)

$$\begin{aligned} \text{Mg}^+ &\longrightarrow \text{Mg}^{2+} + \text{e}^- \text{ (I.E}_2\text{)Mg} \\ \text{(2,8,1)} & \quad \quad \quad \text{(2,8)} \\ \text{Na}^+ &\longrightarrow \text{Na}^{2+} + \text{e}^- \text{ (I.E}_2\text{)Na} \\ \text{(2,8)} & \quad \quad \quad \text{(2,7)} \end{aligned}$$

 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 $(\text{I.E}_2 - \text{I.E}_1) > 16\text{ev} / \text{atom}$, 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)

$$\text{I.E} \propto \frac{1}{\text{size}}$$
 \therefore 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.
 $\text{I.E}_4 \gg \gg \text{I.E}_3$, means valence electron is 3
47. (d)
For alkali earth metal $(\text{I.E})_3 \gg (\text{I.E})_2$.
48. (d)
IE of S < 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)
 $\text{Na} \longrightarrow \text{Na}^+ + \text{e}^- \quad \Delta\text{H}_1 = (\text{I.E}_1)_{\text{Na}}$
 $\text{Na}^+ + \text{e}^- \longrightarrow \text{Na} \quad \Delta\text{H}_2 = \text{EG}_{\text{Na}^+}$
 $|\Delta\text{H}_1| = |\Delta\text{H}_2|$

52. (c)
 $O \xrightarrow{+e^-} O^- \quad \Delta H_{eg_1} = (-ve)$
 $O^- \xrightarrow{+e^-} O^{2-} \quad \Delta H_{eg_2} = +ve$
 $O \xrightarrow{+2e^-} O^{2-} \quad (\Delta H_{eg})_{total} = +ve$
 $\therefore \Delta H_{eg_2}$ 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

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

64. (d)

$$X_p = \frac{X_m}{2.8} \quad X_p: \text{electronegativity on Pauling scale}$$

$X_m: \text{electronegativity on Mulliken's scale.}$

65. (c)

Electronegativity increases along a period and decreases down the group.

66. (a)

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

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 right, 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)

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

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

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

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

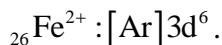
$\therefore (I.E)_2$
 $O > F > N > C.$

75. (b)
Hydration energy $\propto \frac{\text{charge}}{\text{size}}$
76. (a)
 $(\text{I.E.}_1)_N > (\text{I.E.}_1)_O$
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)
 $\text{I. E} \propto \frac{1}{\text{size}}$
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)
 $(\Delta H_{\text{eg}})_2$ is always positive.
12. (d)
 $\text{Fe} = \text{Co} = \text{Ni}$; Increase in atomic number is balanced by increase in repulsion between electrons.
 $\text{Ni} < \text{Cu} < \text{Zn}$ 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 $(\text{I.E})_2 \gg (\text{I.E})_1$
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)
 ${}_{24}\text{Cr}^{2+} : [\text{Ar}]3d^4$
 ${}_{23}\text{V}^{2+} : [\text{Ar}]3d^3$
 ${}_{25}\text{Mn}^{2+} : [\text{Ar}]3d^5 \leftarrow$ Half-filled 'd' orbitals are more stable than other configuration.



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

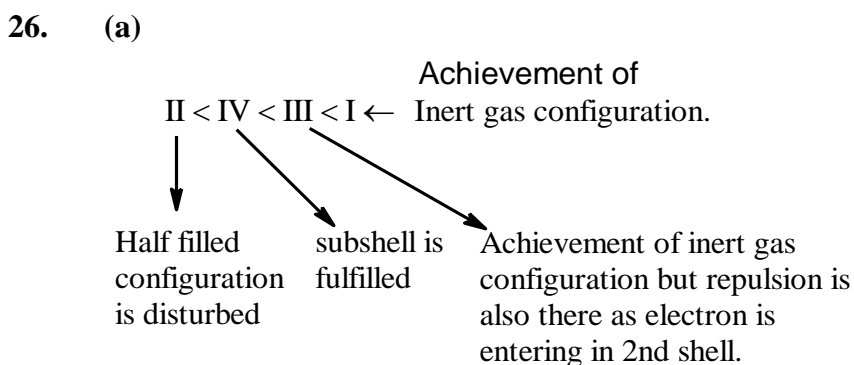
21. (d)
Lawrencium ${}_{103}\text{Lr} : [\text{Rn}]7s^2 6d^1 5f^{14}$.

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)
 ${}_{3}\text{Li}^+ : 1s^2$
 ${}_{4}\text{Be}^+ : 1s^2 2s^1$
 ${}_{5}\text{B}^+ : 1s^2 2s^2$
 $\therefore (\text{I.E})_2$ 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.



27. (a)
Hydrated radius $\propto \frac{\text{Charge}}{\text{size}}$
 $\therefore \text{K}^+(\text{aq}) < \text{Na}^+(\text{aq}) < \text{Li}^+(\text{aq})$

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 = $\frac{110 \times 10^{-3}}{7}$
Energy required = $\frac{110 \times 10^{-3}}{7} \times 720 \text{ kJ}$
= 11.3 kJ.
32. (a)

$$\text{Cl} \xrightleftharpoons[(\text{I.E})_{\text{Cl}^-}]{(\text{E.A})_{\text{Cl}}} \text{Cl}^-$$

$$\therefore |(\text{E.A})_{\text{Cl}}| = |(\text{I.E})_{\text{Cl}^-}|$$
33. (b)
Cl⁻ is more stable than F⁻
∴ (I.E)_{Cl⁻} > (I.E)_{F⁻}
(I.E)_{Cl} > (I.E)_{Cl⁻} - (I.E)_{neutral atom} > (I.E)_{anion}.
(I.E)_F > (I.E)_{Cl}. I.E ∝ $\frac{1}{\text{size}}$
34. (b)
Na⁺ has inert gas configuration.
35. (c)
O⁻ $\xrightarrow{+e^-}$ O²⁻
S⁻ $\xrightarrow{+e^-}$ S²⁻
Incoming electron is repelled by anionic electrons.
36. (a)
Electronegativity ∝ $\frac{1}{\text{size}}$
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

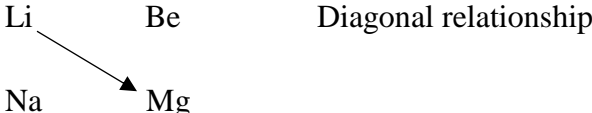
$$X = \frac{\text{I.E} + \text{E.A}}{2}$$

$$2X - Y = \text{I.E}$$

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: $\text{Na} < \text{Al} < \text{Mg} < \text{Si}$
 $\therefore 496 < \text{IE}(\text{Al}) < 737 < 786$
 Only option (c), matches the condition.
8. (c)
Rb and Cs have nearly same electron enthalpy $\Delta H_{\text{eg}} = -46 \text{ kJ mol}^{-1}$.
 Ar and Kr have same eq. $\Delta H_{\text{eg}} = +96 \text{ kJ mol}^{-1}$.
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 $\text{Na} > \text{Mg} > \text{Be} > \text{Si} > \text{P}$.
10. (d)
 ${}_{31}\text{Ga}$ is in group 13 with e^- configuration $[\text{Ar}] 3d^{10} 4s^2 4p^1$ whereas ${}_{30}\text{Zn}$ had e^- configuration as $[\text{Ar}] 3d^{10} 4s^2$. Removal of e^- from 4p orbital is easier than that from 4s 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, $\text{P} > \text{S}$.
12. (a)

 Li Be Diagonal relationship
 ↘
 Na Mg
 $\text{Li}^+ \rightarrow$ 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.
SO₃, SiO₂ → Acidic oxide
CaO → Basic oxide
Al₂O₃ → Amphoteric oxide
16. (b)
Halogens generally formed acidic oxides.
Cl₂O₇ – Acidic oxide
Na₂O – Basic oxide
Al₂O₃ – Amphoteric oxide
N₂O – 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.
N³⁻ > O²⁻ > F⁻ > Na⁺ > Mg²⁺ (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, O²⁻, 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 H₃PO₃ are acidic.
Al(OH)₃ and Be(OH)₂ are amphoteric.
NaOH and Ca(OH)₂ are basic.
22. (b)
-
- 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 $\text{Na} < \text{Mg} < \dots < \text{Cl} < \text{Ar}$.

24. (d)
Correct order of electron gain enthalpy is
 $\text{S} > \text{Se} > \text{Te} > \text{O}$

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 : $\text{Na} < \text{Al} < \text{Mg} < \text{Si}$.
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 ionisation 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 \Rightarrow O^{2-} F^- Na^+ Mg^{2+}
 $8 + 2$ $9 + 1$ $11 - 1$ $12 - 2$
No. of $e^- \Rightarrow$ 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)

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 ium

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IUPAC symbol = Unu

Atomic no. (Z) = 101