UNIT-3 PERIODIC CLASSIFICATION OF ELEMENTS

MY REVISION TIMELINE:-

SUMMARY:-

- Currently 118 elements are known of which elements with atomic no 1 to 92 are found in nature.
- The modern periodic law – the physical and chemical properties of the elements are periodic functions of their atomic numbers.
- The repetition of physical and chemical properties at regular intervals is called periodicity.
- There are 7 periods and 18 groups.
- The groups can be combined as s, p, d & f block elements on the basis of the orbital in which the last valence electron enters.
- Group 1 & 2 – s-block elements.
- Group 13 to 18 – p-block elements.
- Group 3 to 12 – d-block elements.
- Lanthanides and Actinides are f-block elements.
- Elements of the same group have similar electronic configuration in the outer shell.
- Atomic radius is the distance between the centre of its nucleus and outermost shell containing valence electron.
- Covalent radius is one-half of internuclear distance between two identical atoms linked together by a single covalent bond.
- Covalent radius < Actual atomic radius
- Covalent radius of individual atom in case of hetero nuclear diatomic molecules can be calculated using $d_{A-B} = r_A + r_B - 0.09(X_A - X_B)$
- Metallic radius is one-half of the distance between two adjacent metal atoms in the closely packed metallic crystal lattice.
- Atomic radius decreases from left to right along a period and increases down the group.
- Shielding effect – The inner shell electrons act as a shield between the nucleus and the valence electrons.
- Effective nuclear charge – The net nuclear charge experienced by valence electrons in the outermost shell. $Z_{eff} = Z - S$
- Ionic radius: Distance from the centre of the nucleus of the ion up to which it exerts its influence on the electron cloud of the ion.
- Ionisation energy: The minimum amount of energy required to remove the most loosely bound electron from the valence shell of the isolated neutral gaseous atom in its ground state.

  Unit: $kJmol^{-1}$ or eV

- Second ionisation energy: Minimum amount of energy required to remove an electron from a unipositive cation.
- Successive ionisation energy order: $IE_1 < IE_2 < IE_3 < \ldots\ldots$
- Ionisation energy increases along a period (Exceptions: Beryllium > Boron and Nitrogen > Oxygen) and decrease down a group.
- Increased shielding effect decreases ionisation energy.
- Electron affinity – Amount of energy released (required in the case of noble gases) when an electron is added to the valence shell of an isolated neutral gaseous atom in its ground state to form its anion.

  Unit: $kJmol^{-1}$
- From alkali metals to halogens in a period electron affinity increases. (Exception: Beryllium and nitrogen have zero electron affinity)
- Halogens have highest electron affinity. (high negative value)
- In noble gases, addition of electron is unfavourable and will require energy.
- Electron affinity decreases down the group. (Exception: Oxygen < Sulphur and Fluorine < Chlorine)
- Electronegativity: Relative tendency of an element present in a covalently bonded molecule to attract the shared pair of electrons towards itself.
- Electronegativity increases from left to right across a period and decreases down a group.
- Noble gases have zero electronegativity.
- Valence/ Oxidation states is the combining capacity of an atom relative to hydrogen atom.
- In a period valence/ oxidation state increases.
- Anomalous properties of second period elements:
  - Lithium and Beryllium form more covalent compounds but other alkali and alkaline earth metals form ionic compounds.
  - Second period elements maximum covalency is 4 but other members of subsequent periods show higher valences.
- Diagonal relationship: Similarity in properties existing between the diagonally placed elements.
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Elements on left side of periodic table	Elements on right side of periodic table
➢ Less ionisation energy	➢ Have high electron affinity
➢ Readily lose their valence electrons	➢ Readily accept electrons
➢ Lower left elements show metallic character	➢ Top right elements are non-metallic in nature

- On moving down the group electropositive character increases, hydroxides become more basic.

FORMULAS:-

- Covalent radius:
  Homo nuclear diatomic molecules
  $$r_A = \frac{d_{A-A}}{Z}$$
  Hetero nuclear diatomic molecules
  $$d_{A-B} = r_A + r_B - 0.09(X_A - X_B)$$
- Metallic radius = $\frac{distance\ between\ adjacent\ atoms}{2}$
- Z_{eff} = Z - S
- Ionic radius:
  $$d = r_{C^+} + r_{A^-}$$
  where $r_{C^+} \propto \frac{1}{Z_{eff_{C^+}}}$ $\quad r_{A^-} \propto \frac{1}{Z_{eff_{A^-}}}$
  $$\frac{r_{C^+}}{r_{A^-}} = \frac{Z_{eff_{A^-}}}{Z_{eff_{C^+}}}$$
- To calculate electronegativity values
  $$X_A - X_B = 0.182\sqrt{E_{AB} - (E_{AA} \times E_{BB})^{\frac{1}{2}}}$$

HINTS TO SOLVE PROBLEMS:-

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

- General outer electronic configuration of elements in a group:
  s-block elements
  - Group 1 → ns^1
  - Group 2 → ns^2

  d-block elements
  - Group 3 → $ns^2(n-1)d^1$
  - Group 4 → $ns^2(n-1)d^2$
  - Group 5 → $ns^2(n-1)d^3$
  - Group 6 → $ns^2(n-1)d^4$
  - Group 7 → $ns^2(n-1)d^5$
  - Group 8 → $ns^2(n-1)d^6$
  - Group 9 → $ns^2(n-1)d^7$
  - Group 10 → $ns^2(n-1)d^8$

- Group 11 → $ns^2(n-1)d^9$
- Group 12 → $ns^2(n-1)d^{10}$

p-block elements

- Group 13 → ns^2np^1
- Group 14 → ns^2np^2
- Group 15 → ns^2np^3
- Group 16 → ns^2np^4
- Group 17 → ns^2np^5
- Group 18 → ns^2np^6

f-block elements

- Lanthanides → $4f^{1-14}5d^{0-1}6s^2$
- Actinides → $5f^{0-14}6d^{0-2}7s^2$

TEXTBOOK EVALUATION

Multiple choice questions:-

1. What would be the IUPAC name for an element with atomic number 222?
(a) bibibiium (b) bididium
(c) didibium **(d) bibibium**

2. The electronic configuration of the elements A and B are $1s^2, 2s^2, 2p^6, 3s^2$ and $1s^2, 2s^2, 2p^5$, respectively. The formula of the ionic compound that can be formed between these elements is
(a) AB (b) AB_2
(c) A_2B (d) none of the above.

3. The group of elements in which the differentiating electron enters the anti-penultimate shell of atoms are called
(a) p-block elements (b) d-block elements
(c) s-block elements **(d) f-block elements**

4. In which of the following options the order of arrangement does not agree with the variation of property indicated against it? (NEET 2016 Phase 1)
(a) I < Br < Cl < F (increasing electron gain enthalpy)
(b) Li < Na < K < Rb (increasing metallic radius)
(c) $Al^{3+} < Mg^{2+} < Na^+ < F^-$ (increasing ionic size)
(d) B < C < O < N (increasing first ionization enthalpy)

5. Which of the following elements will have the highest electro negativity?
(a) Chlorine (b) Nitrogen
(c) Cesium **(d) Fluorine**

6. Various successive ionization enthalpies (in kJ mol^{-1}) of an element are given below. The element is
(a) phosphorus (b) sodium
(c) aluminium (d) silicon table

7. In the third period, the first ionization potential is of the order
(a) Na > Al > Mg > Si > P **(b) Na < Al < Mg < Si < P**
(c) Mg > Na > Si > P > Al (d) Na< Al < Mg < Si < P

8. Identify the wrong statement
(a) Among-st the iso electronic species, smaller the positive charge on cation, smaller

is the ionic radius

(b) Among-st iso electric species greater the negative charge on the anion, larger is the ionic radius

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

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

9. Which one of the following arrangements represent the correct order of least negative to most negative electron gain enthalpy?

(a) $Al < O < C < Ca < F$ (b) $Al < Ca < O < C < F$

(c) $C < F < O < Al < Ca$ **(d) $Ca < Al < C < O < F$**

10. The correct order of electron gain enthalpy with negative sign of F, Cl, Br and I having atomic number 9, 17, 35 and 53, respectively is

(a) J > Br > Cl >F (b) F > Cl > Br >I

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

11. Which one of the following is the least electro negative element?

(a) Bromine (b) Chlorine

(c) Iodine **(d) Hydrogen**

12. The element with positive electron gain enthalpy is

(a) hydrogen (b) sodium

(c) argon (d) fluorine

13. The correct order of decreasing electro negativity values among the elements X, Y, Z and A with atomic numbers 4, 8, 7 and 12 respectively

(a) Y > Z > X > A (b) Z > A > Y > X

(c) X > Y > Z > A (d) X > Y >A >Z

14. Assertion : Helium has the highest value of ionization energy among all the elements known Reason: Helium has the highest value of electron affinity among all the elements known –

(a) Both assertion and reason are true and reason is correct explanation for the assertion

(b) Both assertion and reason are true but the reason is not the correct explanation for the assertion

(c) Assertion is true and the reason is false

(d) Both assertion and the reason are false

15. The electronic configuration of the atom having maximum difference in first and second ionization energies is

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

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

16. Which of the following is second most electro negative element?

(a) Chlorine (b) Fluorine

(c) Oxygen (d) Sulphur

17. IE_1 and IE_2 of Mg are 179 and 348 k cal mol^{-1} respectively. The energy required for the reaction

$Mg \rightarrow Mg^{2+} + 2e^-$ is

(a) +169 kcal mol^{-1} (b) -169 kcal mol^{-1}

(c) +527 kcal mol^{-1} (d) -527 kcal mol^{-1}

18. In a given shell the order of screening effect is

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

(c) f > d > p > s (d) f > p > s > d

19. Which of the following orders of ionic radii is correct?

(a) $H^- > H^+ > H$ (b) $Na^+ > F^{"} > O^-$
(c) $F > O^{2-} > Na^+$ **(d) None of these**

20. The first ionization potential of Na, Mg and Si are 496, 737 and 786 $kJmol^{-1}$ respectively. The ionization potential of Al will be closer to
(a) 760 $kJ\ mol^{-1}$ **(b) 575 $kJ\ mol^{-1}$**
(c) 801 $kJ\ mol^{-1}$ (d) 419 $kJ\ mol^{-1}$

21. Which one of the following is true about metallic character when we move from left to right in a period and top to bottom in a group?
(a) Decreases in a period and increases along the group
(b) Increases in a period and decreases in a group
(c) Increases both in the period and the group
(d) Decreases both in the period and in the group

22. How does electron affinity change when we move from left to right in a period in the periodic table?
(a) Generally increases (b) Generally decreases
(c) Remains unchanged (d) First increases and then decreases

23. Which of the following pairs of elements exhibit diagonal relationship?
(a) Be and Mg (b) Li and Mg
(c) Be and B **(d) Be and Al**

Write brief answers to the following questions:-

24. Define modern periodic law.
The **physical and chemical properties** of the elements are periodic functions of their atomic numbers.

25. What are isoelectronic ions? Give examples.
Ions having **same number of electrons** are called isoelectronic ions.
Example: Na^+ and F^- ($10e^-$ each)
K^+ and Cl^- ($18e^-$ each)

26. What is effective nuclear charge?
The net nuclear charge experienced by **valence electrons** in the **outermost shell** is called effective nuclear charge.

27. Is the definition given below for ionisation enthalpy is correct?
"Ionisation enthalpy is defined as the energy required to remove the most loosely bound electron from the valence shell of an atom"
No. It is defined as the **minimum amount of energy** required to remove the most loosely bound electron from the valence shell of the **isolated neutral gaseous atom** in its ground state.

28. Magnesium loses electrons successively to form Mg^+, Mg^{2+} and Mg^{3+} ions. Which step will have the highest ionization energy and why?
- $IE_3 > IE_2 > IE_1$. The last step will have highest IE.
- The total number of electron are **less** in the **cation** than the neutral atom while the nuclear charge remain the same.
- The effective nuclear charge on the **cation** (Mg^{2+}) is much **higher** than the other Mg^+ or neutral Mg atom. Since the electron has to be removed against the **net positive charge** on the ions.

29. Define electronegativity.

The relative tendency of an element present in a **covalently bonded molecule** to attract the **shared pair of electrons** towards itself is called electronegativity.

30. How would you explain the fact that the second ionisation potential is always higher than first ionisation potential?

- The total number of electrons are **less** in the **cation** (M^+) than the neutral atom (M), while the nuclear charge remains the same.
- In M^+ the electron has to be removed against the **net positive charge** on the ions.

31. Energy of an electron in the ground state of the hydrogen atom is -2.18×10^{-18} J. Calculate the ionisation enthalpy of atomic hydrogen in terms of kJ mol-1.

Given:

Energy of an electron in the ground state of the hydrogen atom = -2.18×10^{-18} J

Ionisation enthalpy of atomic hydrogen = ?

Solution:

Ionisation enthalpy of atomic hydrogen = -2.18×10^{-18} J/atom

Ionisation enthalpy of atomic hydrogen = $-2.18 \times 10^{-18} \times 6.023 \times 10^{23} \times 10^{-3}$ kJ/mol

$\boxed{\textbf{Ionisation enthalpy of atomic hydrogen = 1312 kJ/mol}}$

32. The electronic configuration of atom is one of the important factor which affects the value of ionisation potential and electron gain enthalpy. Explain.

- Be has more IE than Boron as Be has **completely filled 2s configuration**. The IE decreases in the order: **s>p>d>f**. The first IE of Al is lower than Mg. It is easier to remove p electron in Al than s electron in Mg. The IE is **greater** for **half-filled** and **fully filled electronic configuration**. Half-filled configuration of N-atom has more IE than oxygen, also **noble gases have most stable configuration** and hence maximum IE.
- Be ($1s^2 2s^2$) and N ($1s^2 2s^2 2p^3$) have almost **zero EA** because of **full filled and half-filled orbitals** respectively. EA of **noble gases** are **zero** as they have **np^6** configuration. In each period halogens have **high negative EA** due to **ns^2np^5** configuration. (Ready to accept e^-)

33. In what period and group will an element with Z = 118 will be present?

- The electronic configuration of Z = 188 is
  $1s^2\ 2s^2\ 2p^6\ 3s^2 3p^6\ 4s^2\ 3d^{10}\ 4p^6 5s^2\ 4d^{10} 5p^6\ 6s^2 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2 5f^{14}\ 6d^{10}\ 7p^6$
- So it is a **noble gas**. It belongs to **7th period** and **18th group**.

34. Justify that the fifth period of the periodic table should have 18 elements on the basis of quantum numbers.

On the basis of quantum numbers and electronic configuration, the fifth period has the orbitals in the order **5s 4d** and **5p** to be filled up. (i.e.) **$5s^{1\text{-}2} 4d^{1\text{-}10} 5p^{1\text{-}6}$** (or) **18 electrons**. So 18 elements will occupy this period.

35. Elements a, b, c and d have the following electronic configurations:

a: $1s^2, 2s^2, 2p^6$

b: $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$

c: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

d: $1s^2, 2s^2, 2p^1$

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

- Element a and c belong to same group (Noble gases – 18th group) as they have **np^6** configuration.
- Element b and d belong to the same group (13th group) as they have **p^1** configuration.

36. Give the general electronic configuration of lanthanides and actinides?

- Lanthanides → $4f^{1-14}5d^{0-1}6s^2$
- Actinides → $5f^{0-14}6d^{0-2}7s^2$

37. Why halogens act as oxidising agents?

- Halogens are **highly electronegative** with low dissociation energies and high negative electron gain enthalpies.
- They have a high tendency to **gain an electron**.
- By gaining one electron they complete their **octet**.
- Hence they act as **strong oxidising agents**.

38. Mention any two anomalous properties of second period elements.

- Lithium and Beryllium form more **covalent compound** but other alkali and alkaline earth metals form **ionic compounds**.
- Second period elements show **maximum covalency** of **4** but other members of subsequent periods show **higher valences**.

39. Explain the Pauling method for the determination of ionic radius.

- Ionic radius is defined as the distance from the **center of the nucleus of the ion** up-to which it exerts its influence on the **electron cloud of the ion**.
- Ionic radius of **uni-univalent crystal** can be calculated from the inter-ionic distance between the nuclei of the cation and anion.
- Pauling assumed that ions present in a crystal lattice are **perfect spheres** and they are in contact with each other, therefore

  $d = r_{C^+} + r_{A^-} \rightarrow 1$

  where, d = distance between the center of the nucleus of cation C+ and the anion A-

  r_{C^+} = radius of cation

  r_{A^-} = radius of anion

- Pauling assumed that the radius of the ion having **noble gas configuration** (Na^+ and F^- having $1s^2$, 25^2, $2p^6$ configuration) is **inversely proportional** to the **effective nuclear charge felt at the periphery of the ion.**

  i.e. $r_{C^+} \propto \frac{1}{Z_{eff_{C^+}}} \rightarrow 2$

  $r_{A^-} \propto \frac{1}{Z_{eff_{A^-}}} \rightarrow 3$

  where, **Z_{eff}** is the effective nuclear charge

  $Z_{eff} = Z - S$

- Dividing equation 2 by 3

$$\frac{r_{C^+}}{r_{A^-}} = \frac{Z_{eff_{A^-}}}{Z_{eff_{C^+}}} \rightarrow 4$$

On solving equation 1 and 4, the values of r_{C^+} & r_{A^-} can be obtained.

40. Explain the periodic trend of ionization potential.

Ionization potential is defined as the **minimum amount of energy** required to remove the most loosely bound electron from the **valence shell** of the **isolated neutral gaseous atom in its ground state.**

Its unit is **kJ/mol** or **eV**.

- **Variation in a period:**

  Ionization potential **increases** across a **period from left to right**. This is due to the following reasons
  - Increase of **nuclear charge** in a period.
  - Decrease of **atomic size** in a period.

  Because of these reasons, the valence electrons are tightly held by the nucleus and more amount of energy is require to remove the electron.

- **Variation in a group:**

  Ionization potential **decreases down a group**. This is due to the following reasons
  - Increase of **screening effect** on the outermost electrons due to the increase in the number of inner electrons, decreases the attractive force acting on the valence electron by the nucleus.
  - Gradual increase of **atomic size** in a group.

  Because of these reasons, the valence electrons are loosely held by the nucleus and less amount of energy is require to remove the electron.

41. Explain the diagonal relationship.

- On moving diagonally across the periodic table, the **second and third period elements show certain similarities**. Even though the similarity is not same as we see in a group, it is quite pronounced in the following pair of elements.

  Li Be B C

  Na Mg Al Si

- The similarity in properties existing between the diagonally placed elements is called '**diagonal relationship**'.

42. Why the first ionization enthalpy of sodium is lower than that of magnesium while its second ionization enthalpy is higher than that of magnesium?

- The first IE of sodium is less than that of magnesium because Mg has **more proton** in its nuclear to hold as to the electrons in the **3s** orbital and further Mg has **2e$^-$** with a **completely filled 3s - orbital** which is **stable**. Na has only one valence electron.

  $_{11}Na \rightarrow {}_{10}Na^+ + e^-$

  (2,8,1) (2,8)

- Sodium attains **neon** (octet) configuration easily by losing an e$^-$.

  $_{12}Mg \rightarrow {}_{11}Mg^1 + e^-$

  (2,8,2) (2,8,1)

- But, on losing one as using are electron Mg does not get stable configuration. But **Mg$^+$ attains stable Neon** (octet) **configuration** easily by losing the next electron.

$_{10}Mg^{+} \rightarrow {}_{10}Mg^{2+} + e^{-}$

- But Na cannot lose the second electron easily as the mono positive Na^{+} ion is having neon (octet) configuration. The electrons are **more strongly** attached to **Na^{+} ion** and it requires more energy.

43. By using Pauling's method calculate the ionic radii of K^{+} and Cl^{-} ions in the potassium chloride crystal. Given that $d_{K^+-Cl^-}$ = 3.14 Å.

Given:

$d_{K^+-Cl^-} = 3.14$ Å.

i.e. $r_{K^+} + r_{Cl^-} = 3.14$ Å $\rightarrow$ 1

Formula used:

$$\frac{r_{K^+}}{r_{Cl^-}} = \frac{Z_{eff_{Cl^-}}}{Z_{eff_{K^+}}}$$

Solution:

$Z_{eff_{Cl^-}} = Z - S$

$Z_{eff_{Cl^-}} = 17 – [(0.35×7)+(0.85×8)+(1×2)]$

$Z_{eff_{Cl^-}} = 17 – 11.25$

$Z_{eff_{Cl^-}} = 5.75$

$Z_{eff_{K^+}} = Z - S$

$Z_{eff_{K^+}} = 19 – [(0.35×7)+(0.85×8)+(1×2)]$

$Z_{eff_{K^+}} = 19 – 11.25$

$Z_{eff_{K^+}} = 7.75$

$$\frac{r_{K^+}}{r_{Cl^-}} = \frac{Z_{eff_{Cl^-}}}{Z_{eff_{K^+}}} = \frac{5.75}{7.75} = 0.74$$

$r_{K^+} = 0.74 r_{Cl^-}$

Substitute the value of r_{K^+} in equation 1

$0.74 r_{Cl^-} + r_{Cl^-} = 3.14$ Å

$r_{Cl^-} = \frac{3.14}{1.74}$

$r_{Cl^-} = 1.80$ Å

$r_{K^+} + 1.80 = 3.14$

$r_{K^+} = 3.14 – 1.80$

$r_{K^+} = 1.34$ Å

$\boxed{\therefore\ r_{Cl^-} = 1.80 \text{ Å and } r_{K^+} = 1.34 \text{ Å}}$

44. Explain the following, give appropriate reasons.

i) Ionisation potential of N is greater than that of O

ii) First ionisation potential of C-atom is greater than that of B atom, where as the reverse is true is for second ionisation potential.

iii) The electron affinity values of Be and Mg are almost zero and those of N (0.02 eV) and P (0.80 eV) are very low

iv) The formation of F^{-}(g) from F(g) is exothermic while that of $O2^{-}$(g) from O (g) is endothermic.

Answer:

- N (Z = 7) $\mathbf{1s^2\ 2s^2\ 2p_x^1\ 12p_y^1\ 2p_z^1}$. It has **exactly half filled electronic configuration** and it is more **stable**. Due to stability, ionization energy of nitrogen is high.
  O (Z = 8) $\mathbf{1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1}$. It has **incomplete electronic configuration** and it requires less ionization energy.
  $I.E_1$ N > $I.E_1$ O
- $_6C$ – $\mathbf{1s^2 2s^2 2p^2}$ [Greater nuclear charge]
  $_5B$ - $\mathbf{1s^2 2s^2 2p^1}$
  The size of C-atom is smaller than that of B-atom. So it has large**r nuclear charge**. So it needs higher energy to remove the valence electron of C atom. So the **first IE of carbon is greater than B atom**. After the removal of one electron e^-
  $_6C^+ – 1s^2 2s^2 2p^2$ & $_5B^+$ - 1s22s22p1
  (partially filled) (fully filled)
  The fully filled orbital B^+ has extra stability. So it would require more energy to remove an e^- from B^+. So IE_2 of B atom is higher.
- Be (Z = 4) $\mathbf{1s^2\ 2s^2}$
  Mg (Z = 12) $\mathbf{1s^2\ 2s^2\ 2p^6\ 3s^2}$
  Noble gases $\mathbf{ns^2np^6}$
  These are **more stable atoms**. There is no chance for the addition of extra electrons. So they have **zero EA** values.
  Nitrogen (Z = 7) $\mathbf{1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1}$
  Phosphorus (Z = 15) $\mathbf{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^1}$
  N and P have **half-filled electronic configurations**. Inert gases have fully filled configurations. N and P are not as stable as fully filled or empty valence shell. So they have **0.02eV** (N) and **0.80eV** (P) electron affinity values.
- F (Z = 9) $\mathbf{1s^2\ 2s^2\ 2p^5}$
  F-atom is the most electronegative element. If it gains one electron it can form a stable octet configuration. The formation of F^- release energy because there is tremendous stability associated with the octet electronic configuration. So it is exothermic.
  O (Z = 8) $\mathbf{1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1}$
  To get octet configuration, it requires two electrons. When oxygen gains the first electron, some energy is released and it is **exothermic**. O^- is not having an octet configuration. The process of adding a negative electron to already negative O^- ion requires a lot of energy because of **repulsive forces**. So excess additional energy has to be supplied to force O^- ion to accept one more electron. So the formation of O^{2-} is **endothermic**.

45. What is screening effect?

- The repulsive force between the **inner shell electrons** and the **valence electrons** leads to a decrease in the electrostatic attractive forces acting on the valence electrons by the nucleus.

- Inner shell electrons acts as a **shield** between the **nucleus** and the **valence electrons**. This effect is called **shielding effect**.

46. Briefly give the basis for Pauling's scale of electronegativity.

- Pauling assigned the arbitrary value of electronegativities for **hydrogen** and **fluorine** as **2.1** and **4.0** respectively.
- Based on this the electronegativity value for other elements can be calculated using,

$$\mathbf{X_A - X_B = 0.182\sqrt{E_{AB} - (E_{AA} \times E_{BB})^{\frac{1}{2}}}}$$

47. State the trends in the variation of electronegativity in group and periods.

Electronegativity is defined as the relative tendency of an element present in a **covalently bonded molecule**, to **attract** the **shared pair of electrons** towards itself.

- **Variation of electronegativity in a period:**

  The electronegativity generally **increases across a period from left to right**. The **atomic radius decreases** in a period, as the attraction between the **valence electron** and the **nucleus increases**. Hence the tendency to attract shared pair of electrons increases. Therefore, electronegativity increases in a period.

- **Variation of electronegativity in a group:**

  The electronegativity generally **decreases down a group**. As we move down a group the **atomic radius increases** and the **nuclear attractive force** on the **valence electron decreases**. Hence, the electronegativity decreases.

- **Example:**

  Noble gases are assigned **zero** electronegativity. The electronegativity values of the elements of s-block show the expected decreasing order in a group. Except **13^{th}** and **14^{th} group** all other p-block elements follow the expected decreasing trend in electronegativity.