UNIT – III – PERIODIC CLASSIFICATION OF ELEMENTS

| Mr. S.JOHNSON., M.Sc., M.Sc., B.Ed.,

UNIT – 3 – PERIODIC CLASSIFICATION OF ELEMENTS

II. WRITE BRIEF ANSWER TO THE FOLLOWING QUESTIONS.

24. Define modern periodic law. [QY-18, QY,HY-19, HY-22, MAR-23]

The modern periodic law states that, "the physical and chemical properties of the elements are periodic functions of their atomic numbers."

25. What are isoelectronic ions? Give examples. [FMT,HY-18, QY-19]

Ions of different elements having the same number of electrons are called isoelectronic ions.

Ions of different elements	Na^+	Mg^{2+}	Al^{3+}	F^-	O^{2-}	N^{3-}
No. of electrons	10	10	10	10	10	10

26. What is effective nuclear charge? [CRT-22, QY-23]

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

$$Z_{eff} = Z - S$$

Where **Z** is the atomic number and 'S' is the screening contant.

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" [GMQ-18]

No the above definition is incorrect.

The correct definition is Ionisation energy 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 ionisation energy and why?

Neutral atom $Mg \rightarrow Mg^+ + e^- \ (I.E_1 = X_1)$

Unipositive atom $Mg^+ \rightarrow Mg^{2+} + e^- \ (I.E_2 = X_2)$

Dipositive atom $Mg^{2+} \rightarrow Mg^{3+} + e^- \ (I.E_3 = X_3)$

The third step will have the highest ionization energy $I.E_3 > I.E_2 > I.E_1$

Because from a neutral gaseous atom, the electron removal is easy and less amount of energy is required. But from a dipositive cation, there will be more number of protons than the electrons and more is more forces of attraction between the nucleus and electron. So, the removal of electron in a dipositive cation, becomes highly different and more energy is required.

29. Define electronegativity. [FMT-18, SEP-21, MAR-24]

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

30. How would you explain the fact that the second ionisation potential is always higher than first ionisation potential? [FMT-18, QY-22] (or) Successive ionization energy values increase. Why? [QY-18]

The minimum amount of energy required to remove a unipositive cation is called second ionization energy. It is represented by the following equation, $M^+_{(g)} + IE_2 \rightarrow M^{2+}_{(g)} + 1e^-$

The total number of electrons is less in the cation than the neutral atom while the nuclear charge remains the same. Therefore, the effective nuclear charge of the cation is higher than the corresponding neutral atom. Thus, the successive ionization energies, always increase in the following order $IE_1 < IE_2$. Hence, the second ionization potential is always higher than the first ionization potential.

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

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

$$H \rightarrow H^+ + e^-$$

Energy required to ionize 1 mole of hydrogen atoms, we multiply by the Avogadro constant.

E = 2.18 x 10^{-18} x 6.023 x 10^{23} = 13.130 x 10^5 J mol^{-1} = 1.31 x 10^6 J mol^{-1}

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

- Electronic configuration of an atom affects the value of ionization potential and electron given enthalpy.
- Half-filled valence shell electronic configuration and completely filled valence shell electronic configuration are more stable than partially filled electronic configuration.
- For, eg. Beryllium (Z=4) $1s^2$, $2s^2$ (completely filled electronic configuration)
  Nitrogen (Z=7) $1s^2$, $2s^2$, $2p_x^1$, $2p_y^1$, $2p_z^1$ (Half-filled electronic configuration)
  Both Beryllium and Nitrogen have high ionization energy due to more stable nature.
- In the case of Beryllium ($1s^2$, $2s^2$), Nitrogen ($1s^2$, $2s^2$, $2p^3$) the addition of extra electron will disturb their almost zero electron affinity.
- Noble gases have stable ns^2, np^6 configuration and addition of further electron is unfavourable and they have zero electron affinity.

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

The element with atomic number Z = 118, is present in 7^{th} period and 18^{th} group.

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

- According to Aufbau's principle 5^{th} period has nine orbital (one 5s, five 4d and three 5p) to be filled.
- Nine orbitals can accommodate a maximum of 18 electrons. Hence fifth period of the periodic table should have 18 elements from rubidium (Z=37) to Xenon (Z=54).

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?

(a): $1s^2$, $2s^2$, $2p^6$ (Z=10) - Neon **(Ne)**
(b): $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^1$ (Z=13) - Aluminium **(Al)**
(c): $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$ (Z=18) - Argon **(Ar)**
(d): $1s^2$, $2s^2$, $2p^1$ (Z=5) - Boron **(B)**

- In the above elements, **Ne** and **Ar** belong to same group (Noble gases – 18^{th} group)
- **Al** and **B** belong to the same group (13^{th} group)

36. Give the general electronic configuration of lanthanides and actinides? [JUN19, MAR24]

- The electronic configuration of lanthanides is **$4f^{1-14}$, $5d^{0-1}$, $6s^2$**
- The electronic configuration of actinides is **$5f^{0-14}$, $6d^{0-2}$, $7s^2$**

37. Why halogens act as oxidising agents? [FMT-18, HY-22]

Halogens act as oxidising agents because their electronic configuration is **ns^2 np^5**. So all halogens are ready to gain one electron to attain the nearest inert gas configuration. An oxidising agent is the halogens act as electronegative with low disorientation energy and high negative electron gain enthalpies. Therefore, halogens have a high tendency to gain an electrons.

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38. Mention any two anomalous properties of second period elements. [CRT-22, QY-23]

- In the 1st group, lithium forms compounds with more covalent character while the other elements of this group form only ionic compounds.
- In the 2nd group, beryllium forms compounds with more covalent character while the other elements of this group form only ionic compounds.

39. Explain the pauling method for the determination of ionic radius. [SEP-20, MAR-22]

- Ionic radius is defined as the distance from the centre 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 using Pauling's method 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, $d = r_{C+} + r_{A-}$ ----------(1)
- Where d is the distance between the centre of the nucleus of cation C^+ and anion A^- and r_{C+}, r_{A-} are the radiusof the cation and anion respectively.
- Pauling also assumed that the radius of the ion having noble gas electronic configuration is inversely proportional to the effective nuclear charge.

$$r_{C^+} \propto \frac{1}{(Z_{eff})_{C+}} \quad \text{----------(2)}$$

$$r_{A^-} \propto \frac{1}{(Z_{eff})_{A^-}} \quad \text{----------(3)}$$

Where Z_{eff} is the effective nuclear charge and $Z_{eff} = Z - S$

Dividing the equation 2 by 3

$$\frac{r_{C^+}}{r_{A^-}} = \frac{(Z_{eff})_{A^-}}{(Z_{eff})_{C+}} \quad \text{----------(4)}$$

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

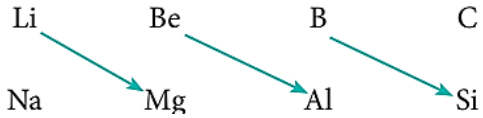
40. Explain the periodic trend of ionisation potential. [MAR-24]

Variation along a period: The ionisation energy usually increases along a period with few exceptions. As discussed earlier, when we move from left to right in a period.

Periodic variation in group: The ionisation energy decreases down a group. As we move down a group, the valence electron occupies new shells, the distance between the nucleus and the valence electron increases. So, the nuclear forces of attraction on valence electron decreases and hence ionisation energy also decreases down a group.

41. Explain the diagonal relationship. [FMT-18, QY,MAR-19, SEP-21, MAR-23]

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.



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

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

The first ionization enthalpy of sodium is lower than that of magnesium.

$$Na(1s^2, 2s^2, 2p^6, 3s^1) + IE_1 \rightarrow Na^+ (1s^2, 2s^2, 2p^6) + e^-$$

$$Mg(1s^2, 2s^2, 2p^6, 3s^2) + IE_1 \rightarrow Mg^+(1s^2, 2s^2, 2p^6, 3s^1) + e^-$$

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Magnesium has completely filled 3s orbital ($1s^2, 2s^2, 2p^6, 3s^2$), is more stable than partially filled valence shell electronic configuration of sodium ($1s^2, 2s^2, 2p^6, 3s^1$).

$$Na^+ (1s^2, 2s^2, 2p^6) + IE_2 \rightarrow Na^{2+} (1s^2, 2s^2, 2p^5) + e^-$$

$$Mg^+ (1s^2, 2s^2, 2p^6, 3s^1) + IE_2 \rightarrow Mg^{2+} (1s^2, 2s^2, 2p^6, 3s^2) + e^-$$

Na^+ has completely filled 2p orbital ($1s^2, 2s^2, 2p^6$), is more stable than partially filled valence shell electronic configuration of Mg^+ ($1s^2, 2s^2, 2p^6$). Hence the second ionization energy of sodium is higher than that of magnesium.

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

$d_{K^+-Cl^-} = 3.14$ Å; $r_{K^+} = ?$; $r_{K^+} + r_{Cl^-} = 3.14$ Å ----------(1)

$\mathbf{Z_{eff(Cl^-)}}$ = Z – S = 17 – [(0.35 x 7) + (0.85 x 8) + (1 x 2)] = 17 – 11.25 = **5.75**

$\mathbf{Z_{eff(K^+)}}$ = Z – S = 19 – [(0.35 x 7) + (0.85 x 8) + (1 x 2)] = 19 – 11.25 = **7.75**

We know that $\frac{r_{K^+}}{r_{Cl^-}} = \frac{(Z_{eff})_{Cl^-}}{(Z_{eff})_{K^+}} = \frac{5.75}{7.75} = 0.74$

$\therefore r_{K^+} = 0.74\, r_{Cl^-}$ Substitute the value of r_{K^+} in equation (1)

$0.74\, r_{Cl^-} + r_{Cl^-} = 3.14$ Å

$1.74\, r_{Cl^-} = 3.14$ Å

$\therefore r_{Cl^-} = \frac{3.14}{1.74}$ Å = **1.81 Å**

From Equation (1) $r_{K^+} = 3.14 - 1.81 =$ **1.33 Å** $\therefore r_{Cl^-} =$ **1.81Å** and $r_{K^+} =$ **1.33Å**

44. Explain the following, give appropriate reasons. (i) Ionisation potential of N is greater than that of O [FMT-18, JUN-19] (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 [SEP-20] (iv) The formation of F^- (g) from F(g) is exothermic while that of O^{2-}(g) from O (g) is endothermic.

(i) N (Z = 7) $1s^2, 2s^2, 2p_x^1, 12p_y^1, 2p_z^1$. It has exactly half-filled electronic configuration and it is more stable. Due to stability, ionization energy of nitrogen is high.

O (Z = 8) $1s^2, 2s^2, 2p_x^2, 2p_y^1, 2p_z^1$. It has incomplete electronic configuration and it requires less ionization energy. $\therefore I.E_1$ N > $I.E_1$ O

(ii) C (Z = 6) $1s^2, 2s^2, 2p_x^1, 2p_y^1$. The electron removal from p orbital is very difficult. So carbon has highest first ionization potential. B (Z = 5) $1s^2, 2s^2, 2p^1$. In boron nuclear charge is less than that of carbon, so boron has lowest first ionization potential. $\therefore I.E_1$ C > $I.E_1$ B

But it is reverse in the case of second ionization energy. Because in case of B^+ the electronic configuration is $1s^2, 2s^2$, which is completely filled and it has high ionization energy. But in C^+ the electronic configuration is $1s^2, 2s^2, 2p^1$, one electron removal is easy so it has low ionization energy. $\therefore I.E_2$ B > $I.E_2$ C

(iii) Be (Z = 4) $1s^2, 2s^2$; Mg (Z = 12) $1s^2, 2s^2, 2p^6, 3s^2$ Noble gases has the electronic configuration of $ns^2\, np^6$. All these are completely filled and are more stable. For all these elements Be, Mg and noble gases, addition of electron is unfavourable and so they have zero electron affinity.

Nitrogen (Z = 7) $1s^2, 2s^2, 2p_x^1, 2p_y^1, 2p_z^1$. It has half-filled electronic configuration. So addition of electron is unfavourable and it has very low electron affinity value of 0.02 eV. Phosphorus (Z = 15) $1s^2, 2s^2, 2p^6, 3s^2, 3p_x^1, 3p_y^1, 3p_z^1$. It also has half-filled electronic configuration. Due to the symmetry and more stability, it has very low electron affinity value of 0.80 eV.

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(iv) $F_{(g)} + e^- \rightarrow F_{(g)}^-$ exothermic F (Z = 9) $1s^2, 2s^2, 2p^5$. It is ready to gain one electron to attain the nearest inert gas configuration. By gaining one electron, energy is released, so it is an exothermic reaction.

$O_{(g)} + 2e^- \rightarrow O^{2-}_{(g)}$ endothermic O (Z = 8) $1s^2, 2s^2, 2p_x^1, 2p_y^1, 2p_z^1$. It is the small atom with high electron density. The first electron affinity is negative because energy is released in the process of adding one electron to the neutral oxygen atom. Second electron affinity is always endothermic (positive) because the electron is added to an ion which is already negative, therefore it must overcome the repulsion.

45. What is screening effect?

- In addition to the electrostatic forces of attraction between the nucleus and the electrons, there exists repulsive forces among the electrons.
- 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.
- Thus, the inner shell electrons act 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. [HY-22]

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

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

Where E_{AB}, E_{AA} and E_{BB} are the bond dissociation energies (K_{cal}) of AB, A_2 and B_2 molecules respectively

47. State the trends in the variation of electronegativity in group and periods. [SEP-21, AUG-22]

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.

Variation of Electronegativity in a period: The electronegativity generally increases across a period from left to right. As discussed earlier, 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 also increases in a period.

EVALUATE YOURSELF

1. What is the basic difference in approach between Mendeleev's periodic table and modern periodic table?

The main basic difference between Mendeleev's periodic table and modern periodic table is that first one is constructed on the basic of atomic weight and the latter is constructed on the basic of atomic number.

2. The element with atomic number 120 has not been discovered so far. What would be the IUPAC name and the symbol for this element? Predict the possible electronic configuration of this element.

Atomic number : 120; IUPAC temporary symbol : Unbinilium

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IUPAC temporary symbol : Ubn Possible electronic configuration : [Og] $8s^2$

3. Predict the position of the element in periodic table satisfying the electronic configuration (n-1)d^2, ns^2 where n=5 [QY-22]

Electronic configuration : (n-1)d^2, ns^2

For n = 5, the electronic configuration is, $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^6$, $4d^2$, $5s^2$

Atomic number : **40** ∴ 4^{th} group and 5^{th} period (d block) element is **Zirconium**.

4. Using Slater's rule calculate the effective nuclear charge on a 3p electron in aluminium and chlorine. Explain how these results relate to the atomic radii of the two atoms. [QY-22]

Electronic Configuration of Aluminium $_{13}Al$ $\underbrace{1s^2}_{n-2}$ $\underbrace{2s^2\ 2p^6}_{n-1}$ $\underbrace{3s^2\ 3p^1}_{n}$

Group	No. of electrons	Contribution of each electron to 'S' value	Contribution of a particular group
n	2	0.35	0.70
n-1	8	0.85	6.80
n-2	2	1	2.00
		Total (S):	**9.50**

Effective nuclear charge = Z – S = 13 – 9.5 = 3.5; ∴ $\mathbf{(Z_{eff})_{Al} = 3.5}$

Electronic Configuration of Chlorine $_{17}Cl$ $\underbrace{1s^2}_{n-2}$ $\underbrace{2s^2\ 2p^6}_{n-1}$ $\underbrace{3s^2\ 3p^5}_{n}$

Group	No. of electrons	Contribution of each electron to 'S' value	Contribution of a particular group
n	6	0.35	2.10
n-1	8	0.85	6.80
n-2	2	1	2.00
		Total (S):	**10.90**

Effective nuclear charge = Z – S = 17 – 10.9 = 6.1; ∴ $\mathbf{(Z_{eff})_{Cl} = 6.1}$

$\mathbf{(Z_{eff})_{Cl} > (Z_{eff})_{Al}}$ and hence $r_{Cl} < r_{Al}$

5. A student reported the ionic radii of isoelectronic species X^{3+}, Y^{2+} and Z^- as 136 pm, 64pm and 49 pm respectively. Is that order correct? Comment.

X^{3+}, Y^{2+}, Z^- are isoelectronic.

∴ Effective nuclear charge is in the order $(Z_{eff})_{Z^-} < (Z_{eff})_{Y^{2+}} < (Z_{eff})_{X^{3+}}$

Hence, ionic radii should be in the order $r_{Z^-} > r_{Y^{2+}} > r_{X^{3+}}$

∴ The correct values are,

Species	Ionic radii
Z^-	136
Y^{2+}	64
X^{3+}	49

6. The first ionisation energy (IE_1) and second ionisation energy (IE_2) of elements X, Y and Z are given below.

Element	IE_1 (kJ mol^{-1})	IE_2 (kJ mol^{-1})
X	2370	5250
Y	522	7298
Z	1680	3381

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Which one of the above elements is the most reactive metal, the least reactive metal and a noble gas?

Noble gases: Ionization energy ranging from 2372 KJ mol^{-1} to 1037 KJ mol^{-1}.

For element X, the IE_1, value is in the range of noble gas, moreover for this element for this element both IE_1 and IE_2 are higher and hence X is the noble gas.

For Y, the first ionization energy is low and second ionisation energy is very high and hence Y is most reactive metal.

For Z, both IE_1 and IE_2 are higher and hence it is least reactive.

7. **The electron gain enthalpy of chlorine is 348 kJ mol^{-1}. How much energy in kJ is released when 17.5 g of chlorine is completely converted into Cl^- ions in the gaseous state?**

$Cl_{(g)} + e^- \rightarrow Cl^-_{(g)}$ $\Delta H = 348$ KJ mol^{-1}

For one mole (35.5g) 348KJ is released.

$\therefore$ For 17.75g chlorine energy released $= \frac{348\ KJ}{35.5\ g} \times 17.75\ g$

$\therefore$ The amount of energy released $= \frac{348}{2} =$ **174 KJ**

GOVERNMENT QUESTIONS AND ANSWERS

1. **Atomic number of elements X,Y,Z and A are 4,8,7 and 12 respectively. Arrange them in the decreasing order of their electronegativity. [GMQ-18]**

Y > Z > X > A

2. **Calculate the Effective nuclear charge of helium. [QY-18]**

$Z_{eff} = Z - S = 2 - 0.30$ (for 1s e^- = 0.30) = **1.70**

3. **Define Valency. How is it determined? [MAR-19]**

Define valency of an element may be defined as the combining capacities of elements. The electrons present in the outermost shell are called valence electrons and these electrons determine the valency of an atom.

4. **Calculate the effective nuclear charge experienced by the 4s electron in K atom. [HY-19]**

The electronic configuration of $_{19}K$ atom is $(1s^2)$, $(2s^2, 2p^6)$, $(3s^2, 3p^6)$, $4s^1$

Effective nuclear charge,

$Z_{eff} = Z - S = 19 - [(0.85 \times$ No. of electrons in $(n-1)^{th}$ shell $+ (1.00 \times$ total No. of electrons in the inner shell$)]$

$Z_{eff} = 19 - [(0.85 \times 8) + (1.00 \times 10)] =$ **2.20**

5. **First ionization potential of carbon atom is greater than that of boron atom, where as the reverse is true for second ionization potential give appropriate reason. [GMQ-18] (or) Compare the first ionization energies of Carbon and Boron give reason. [QY-19]**

Electronic configuration of Carbon (Z=6) $1s^2, 2s^2, 2p^2$.

Electronic configuration of Boron (Z=5) $1s^2, 2s^2, 2p^1$.

- The size of carbon atom is smaller than boron. So the valence electron of carbon has greater nuclear charge than that of boron. Hence the first I.E of carbon is greater than that of boron.
- However, the second ionisation enthalpy of boron is 1 electron, Boron has a fully filled orbital $(2s^2)$ than carbon $(2p^1)$.
- Fully filled orbitals have more stability than partially filled orbitals so greater amount of energy will be needed to remove an electron from boron.
- So in this case, the second $I.E_2$ of boron is higher than that of carbon.

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6. How does the ionisation energy vary across a period and down a group? [HY-19, QY-23]

Two properties are important in determining ionisation energies:

- Nuclear charge
- Shielding by other electrons
- In partially filled shells, electrons shield each other very imperfectly, so across the period (from left to right) as the nuclear increases, ionisation energies markedly increase.
- On the other hand, down a group, the increased nuclear charge is effectively shielded by the filled electronic shells. Nucleus/valence electron attraction becomes attenuated and ionization energies decrease.

7. Ionisation potential of Nitrogen is greater than that of Oxygen. Explain. [QY,CRT-22]

An atom is highly stable when its valence orbital is completely filled or half-filled and in nitrogen, the valence p orbital is exactly half-filled. So, p orbital in nitrogen is more stable than in oxygen, which has one electron more than the half-filled configuration. So, the ionization energy of nitrogen is more stable than that of oxygen.

- N electronic configuration is $1s^2$, $2s^2$, $2p^3$ (half-filled)
- electronic configuration is $1s^2$, $2s^2$, $2p^4$ (partially filled)

Therefore, the ionisation of nitrogen will be higher than that of oxygen.

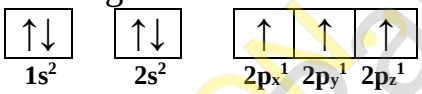
8. Define electron affinity. [MAY-22]

It is defined as the amount of energy released (required in the case 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. It is expressed in kJ mol^{-1} $A + e^- \rightarrow A^- + EA$

9. Write the electronic configuration and orbital diagram for Nitrogen. [MAY-22]

Electronic configuration is $1s^2$, $2s^2$, $2p^3$

Orbital diagram for Nitrogen



10.What are f-block elements? [AUG-22]

The lanthanides ($4f^{1-14}$, $5d^{0-1}$, $6s^2$) and the actinides ($5f^{0-14}$, $6d^{0-2}$, $7s^2$) are called f-block elements. These elements are metallic in nature and have high melting points. Their compounds are mostly coloured. These elements also show variable oxidation states.

11.Compare the ionisation energy of Beryllium and Boron. [AUG-22]

- It is expected that boron has higher ionisation energy than beryllium since it has higher nuclear charge.
- However, the actual ionisation energies of beryllium and boron are 899 and 800 KJ mol^{-1} respectively contrary to the expectation.
- It is due to the fact that beryllium with completely filled 2s orbital, is more stable than partially filled valence shell electronic configuration of boron. ($2s^2$, $2p^1$)
- The electronic configuration of beryllium (Z=4) in its ground state is $1s^2$, $2s^2$ and that of boron is (Z=5) $1s^2$, $2s^2$, $2p^1$.

12.Define effective nuclear charge. Arranges s, p, d and f orbitals in the descending order of their effective nuclear charge. [QY-19]

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

$$Z_{eff} = Z - S$$

$$s > p > d > f$$

UNIT – III – PERIODIC CLASSIFICATION OF ELEMENTS

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s-orbital is very close to each other therefore it will experience the maximum Effective nuclear charge.

13.State and explain Dobereiner's "Triad". [MAR-19]

Some elements such as chlorine, bromine and iodine with similar chemical properties into the group of three elements called as triads. In triads, the atomic weight of the middle element nearly equal to the arithmetic mean of the atomic weights of the remaining two elements. However, only a limited number of elements can be grouped as triads.

S. No.	Elements in the Triad	Atomic weight of middle element	Average atomic weight of the remaining elements
1	Li, Na, K	23	$\frac{7+39}{2} = 23$
2	Cl, Br, I	80	$\frac{35.5+127}{2} = 81.25$
3	Ca, Sr, Ba	88	$\frac{40+137}{2} = 88.5$

14.Define atomic radius. Explain the variation of atomic radius in group and periods. [SEP-21]

Atomic radius of an atom is defined as the distance between the centre of its nucleus and the outermost shell containing the valence electron.

It is not possible to measure the radius of an isolated atom directly. Except for noble gases, usually atomic radius is referred to as covalent radius or metallic radius depending upon the nature of bonding between the concerned atoms.

Variation in Groups: In the periodic table, the atomic radius of elements increases down the group. As we move down a group, new shells are opened to accommodate the newly added valence electrons. As a result, the distance between the centre of the nucleus and the outermost shell containing the valence electron increases. Hence, the atomic radius increases.

Variation in Periods: Atomic radius tends to decrease in a period. As we move from left to right along a period, the valence electrons are added to the same shell. The simultaneous addition of protons to the nucleus, increases the nuclear charge, as well as the electrostatic attractive force between the valence electrons and the nucleus. Therefore atomic radius decreases along a period.