Chapter – 3

Periodic Classification of Elements

I. Choose the best Answer:

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

Answer:
(d) bibibium

Question 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.

Answer:
(a) AB_2

Question 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

Answer:
(d) f-block elements

Question 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)

Answer:
(a) $I < Br < Cl < F$ (increasing electron gain enthalpy)

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

Answer:
(d) Fluorine

Question 6.
Various successive ionization enthalpies (in kJ mol^{-1}) of an element are given below. The element is …………

IE_1	IE_2	IE_3	IE_4	IE_5
577.5	1,810	2,750	11,580	14,820

(a) phosphorus
(b) sodium
(c) aluminium
(d) silicon table

Answer:
(c) aluminium

Question 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

Answer:
(b) Na < Al < Mg < Si < P

Question 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.

Answer:
(a) Among-st the iso electronic species, smaller the positive charge on cation, smaller is the ionic radius

Question 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

Answer:
(d) Ca < Al < C < O < F

Question 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

Answer:
(c) Cl > F > Br > I

Question 11.
Which one of the following is the least electro negative element?
(a) Bromine
(b) Chlorine
(c) Iodine
(d) Hydrogen

Answer:
(d) Hydrogen.

Solution:
Hydrogen is the least electro negative element. Since electro negativity increases across the period from left to right. Hydrogen is the first element and it has less electro negativity and down the group electro negativity decreases.

Question 12.
The element with positive electron gain enthalpy is ……….
(a) hydrogen
(b) sodium
(c) argon
(d) fluorine

Answer:
(c) argon

Solution:
Argon has completely filled configuration. So addition of the electron is not possible and has positive electron gain enthalpy.

Question 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

Answer:
(a) Y > Z > X > A

Question 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

Answer:
(c) Assertion is true and the reason is false

Question 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) $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^1$

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

Question 16.
Which of the following is second most electro negative element?
(a) Chlorine
(b) Fluorine
(c) Oxygen
(d) Sulphur

Answer:
(a) Chlorine

Question 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}

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

Question 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

Answer:
(a) s > p > d > f

Question 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

Answer:
(d) None of these

Question 20.
The first ionization potential of Na, Mg and Si are 496, 737 and 786 kJ mol^{-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}

Answer:
(b) 575 kJ mol^{-1}

Question 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

Answer:
(a) Decreases in a period and increases along the group

Question 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

Answer:
(a) Generally increases.

Question 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

Answer:
(d) Be and Al

II. Write a brief answer to the following questions.

Question 24.
Define modern periodic law.

Answer:
The physical and chemical properties of the elements are periodic functions of their atomic numbers.

Question 25.
What are isoelectronic ions? Give examples.

Answer:
There are some ions of different elements having the same number of electrons are called isoelectronic ions.
Example:
Na^+, Mg^{2+}, Al^{3+}, F^- , O^{2-} and N^{3-}

Question 26.
What is effective nuclear charge?

Answer:
The net charge experienced by valence electrons in the outermost shell is called the effective nuclear charge. It is approximated by the equation $Z_{eff} = Z - S$, where Z is the atomic number and S is the screening constant which can be calculated using Slater's rules.

Question 27.
Is the definition given below for ionization enthalpy is correct?
"Ionization enthalpy is defined as the energy required to remove the most loosely bound electron from the valence shell of an atom"

Answer:
No. It is not correct. The accurate and absolute definition is as follows:
Ionization 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.

Question 28.
Magnesium loses electrons successively to form Mg^+, Mg^{2+} and Mg^{3+} ions. Which step will have the highest ionization energy and why?

Answer:

$Mg_{(g)}$ (Neutral atom) $\rightarrow Mg^+ + \mathbf{e}\,(I.E_1 = X_1)$

(Dbt: **e** or **e**$^-$)

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

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

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 di positive cation, there will be more number of protons than the electrons and there is more forces of attraction between the nucleus and electron. So the removal of electron in a di positive cation, becomes highly difficult and more energy is required.

Question 29.
Define electro negativity.

Answer:
Electro negativity is the relative tendency of an element present in a covalently bonded molecule, to attract the shared pair of electrons towards itself.

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

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

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 $I.E_1 < I.E_2$. Hence, the second ionization potential is always higher than the first ionization potential.

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

Answer:
Energy of an electron in the ground state of the hydrogen atom = -2.18×10^{-18} J
$H \rightarrow H^+ + e^-$
Energy required to ionize 1 mole of hydrogen atoms, we multiply by the Avogadro constant.
$E = 2.18 \times 10^{-18} \times 6.023 \times 10^{23}$
$= 13.123 \times 10^5\ J\ mol^{-1}$
$I.E = +1312\ K\ J\ mol^{-1}$

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

Answer:
The plot of atomic number vs ionization energy shows that there are two deviations in the trends of ionization energy. It is expected that boron has higher ionization energy than beryllium since it has a higher nuclear charge. However, the actual ionization 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. ($2s^2\ 2p^1$).

Similarly, nitrogen with a $1s^2\ 2s^2\ 2p^3$ electronic configuration has higher ionization energy (1402 $kJ\ mol^{-1}$) than oxygen (1314 $kJ\ mol^{-1}$). Since the half-filled electronic configuration is more stable, it requires higher energy to remove an electron from 2p orbital of nitrogen. Whereas the removal of one 2p electron from oxygen leads to a stable half-filled configuration. This makes it comparatively easier to remove 2p electron from oxygen.

As we move from alkali metals to halogens in a period, generally electron affinity increases. This is due to an increase in the nuclear charge and a decrease in the size of the atoms. However, in the case of elements such as beryllium ($1s^2$ $2s^2$), nitrogen ($1s^2$ $2s^2$ $2p^3$) the addition of extra electrons will disturb their stable electronic configuration and they have almost zero electron affinity.

Noble gases have a stable ns^2 np^6 configuration, and the addition of further electrons is unfavourable and requires energy. Halogens having the general electronic configuration of ns^2 np^5 readily accept an electron to get the stable noble gas electronic configuration {ns^2 np^6) and therefore, in each period the halogen has a high electron affinity.

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

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

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

Answer:
In the fifth period, the filling of valence electrons starts with 5s orbital followed by 4d and 6p orbitals. The filling of 4d orbitals starts with Yttribium and ends with cadmium. There are 10 elements present in the second transition series. The period starts with Rubidium (Rb – $5s^1$) and ends with Xenon (Xe – $5s^2$ $5p^6$).

Question 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?

Answer:

		Element
a: $1s^2, 2s^2, 2p^6$	(Z = 10)	Ne
b: $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$	(Z = 13)	Al
c: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$	(Z = 18)	Ar
d: $1s^2, 2s^2, 2p^1$	(Z = 5)	B

1.
2. In the above elements, Ne and Ar belong to same group (Noble gases – 18th group).
3. Al and B belong to the same group (13^{th} group).

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

Answer:
The general electronic configuration of lanthanides is $4f^{1-14}$ $5d^{0-1}$ $6s^2$ and for Actinides is $5f^{1-14}$ $6d^{0-1}$ $7s^2$.

Question 37.
Why halogens act as oxidizing agents?

Answer:
Halogens act as oxidizing agents. Their electronic configuration is ns^2 np^5. So all the halogens are ready to gain one electron to attain the nearest inert gas configuration. An oxidizing agent is the one which is ready to gain an electron. So all the halogens act as oxidizing agents. Also halogens are highly electro negative with low dissociation energy and high negative electron gain enthalpies. Therefore, the halogens have a high tendency to gain an electron. Hence they act as oxidizing agents.

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

Answer:
The anomalous properties of second-period elements are as follows. The ionization energy of Boron is greater than that of Beryllium due to the fact that Be has completely filled 2s orbital which is more stable than the partially filled valence shell electronic configuration of Boron. Similarly, nitrogen with

a half-filled electronic configuration has higher ionization energy than oxygen because the half-filled electronic configuration is more stable.

Question 39.
Explain the Pauling's method for the determination of ionic radius.

Answer:
1. 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.

2. Ionic radius of uni-univalent crystal can be calculated from the inter-ionic distance between the nuclei of the cation and anion.

3. 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 = distance between the center of the nucleus of cation C+ and the anion A-
r_{C^+} = radius of cation
r_{A^-} = radius of anion.

4. 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.

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

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

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

5. Dividing the equation (2) by (3)

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

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

Question 40.
Explain the periodic trend of ionization potential.

Answer:
Variation along the period:
The ionization energy usually increases along a period with few exceptions. When we move from left to right along a period, the valence electrons are added to the same shell, at the same time protons arc added to the nucleus. This successive increase of nuclear charge increases the electrostatic attractive force on the valence electron and more energy is required to remove the valence electron resulting in high ionization energy.

Consider the variation in ionization energy of second-period elements. The plot of atomic number vs ionization energy is given below. The plot of atomic number vs ionization energy shows that there are two deviations in the trends of ionization energy. It is expected that boron has higher ionization energy than beryllium since it has a higher nuclear charge.

However, the actual ionization 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. ($2s^2$, $2p^1$).

Similarly, nitrogen with $1s^2$, $2s^2$, $2p^3$ electronic configuration has higher ionization energy (1402 kJ mol^{-1}) than oxygen (1314 kJ mol^{-1}). Since the half-filled electronic configuration is more stable, it requires higher energy to remove an electron from the 2p orbital of nitrogen. Whereas the removal of one 2p electron from oxygen leads to a stable half-filled configuration. This makes it comparatively easier to remove 2p electron from oxygen.

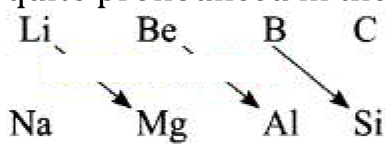
Variation along with the group:
The ionization energy decreases down a group. As we move down a group, the valence electron occupies new shells, the distance between the nucleus and the valence electron increases. So, the nuclear forces of attraction on valence electron decreases, and hence, ionization energy also decreases down a group.

As we move down a group, the number of inner-shell electrons increases which in turn increases the repulsive force exerted by them on the valence electrons, i.e., the increased shielding effect caused by the inner electrons decreases the attractive force acting on the valence electron by the nucleus. Therefore, the ionization energy decreases.

Question 41.
Explain the diagonal relationship.

Answer:

- On moving diagonally across the periodic table, the second and the 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".

Question 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?

Answer:
The first ionization enthalpy of sodium is lower than that of magnesium.
$Na(1s^2, 2s^2, 2p^6, 3s^1) + IE_1 \rightarrow Na^+ (1s^2, 2s^2, 2p^6) + e$
$Mg(1s^2, 2s^2, 2p^6, 3s^2) + IE_1 \rightarrow Mg^+(1s^2, 2s^2, 2p^6, 3s^1) + e$

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.

Question 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

Answer:
Given

$$d_{K^+-Cl^-} = 3.14 \text{ Å}$$

$$r_{K^+} = ? \qquad r_{Cl^-} = ?$$

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

We know that,

$$\frac{r_{K^+}}{r_{Cl^-}} = \frac{(Z_{eff})_{Cl^-}}{(Z_{eff})_{K^+}}$$

$(Z_{eff})_{Cl^-} = Z - S$
$= 17 - [(0.35 \times 7) + (0.85 \times 8) + (1 \times 2)]$
$= 17 - 11.25 = 5.75$
$(Z_{eff})_{K^+} = Z - S$
$= 19 - [(0.35 \times 7) + (0.85 \times 8) + (1 \times 2)]$
$= 19 - 11.25 = 7.75$

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

$r_{(K^+)} = 0.74\, r_{(Cl^-)}$
Substitute the value of $r_{(K^+)}$ in equation (1)
$0.74\, r_{(Cl^-)} + r_{(Cl^-)} = 3.14$ Å

$$1.74\, r_{(Cl^-)} = \frac{3.14\text{Å}}{1.74} = 1.81 \text{ Å}.$$

Question 44.
Explain the following, give appropriate reasons.

1. Ionization potential of N is greater than that of O.
2. First ionization potential of C-atom is greater than that of B-atom, where as the reverse is true for second ionization potential.
3. The electron affinity values of Be, Mg and noble gases are zero and those of N (0.02 eV) and P (0.80 eV) are very low
4. The formation of F^- (g) from F(g) is exothermic while that of O^{2-} (g) from O (g) is endothermic.

Answer:

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

2. 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.
$I.E_1\ C > I.E_1\ B$

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.
$I.E_2\ B > I.E_2\ C$

3. 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 unfavorable 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 unfavorable 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.

4. $F_{(g)} + e^- \rightarrow F_{(g)}^-$ 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.

Question 45.
What is screening effect? Briefly give the basis for Pauling's scale of electro negativity. Screening effect:

Answer:
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 the shielding effect.

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. Electronegativity is not a measurable quantity. In the Pauling scale, an arbitrary values of electronegativities for hydrogen and fluorine are assigned as 2.2 and 4.0 respectively. Based on this, the electronegativity values for other elements can be calculated using the following expression,

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

where E_{AB}, E_{AA}, and E_{BB} are the bond dissociation energies of AB, A_2, and B_2molecules respectively.

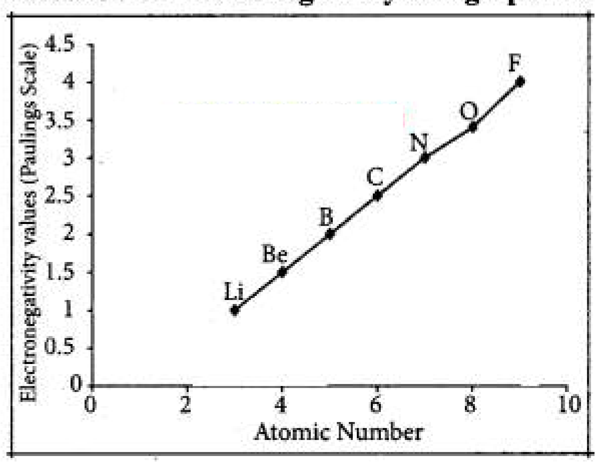
The electronegativity of any given element is not a constant and its value depends on the element to which it is covalently bound. The electronegativity values play an important role in predicting the nature of the bond.

Question 46.
State the trends in the variation of electro negativity in period and group.

Answer:
Variation of electron negativity in a period:
The electro negativity increases across a period from left to right. Since the

atomic radius decreases in a period, the attraction between the valence electron and the nucleus increases. Hence the tendency to attract shared pair of electrons increases. Therefore, electro negativity increases in a period.

Variation of Electronegativity along I period



Variation of electro negativity in a group:

The electro negativity 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 electro negativity decreases in a group.