VELAMMAL MATRIC HR.SEC SCHOOL-SURAPET

XI-CHEMISTRY

3. PERIODIC CLASSIFICATION OF ELEMENTS

TEXT BOOK QUESTIONS AND ANSWERS:

1.Define modern periodic law.

The physical and chemical properties of the elements are periodic functions of their
Atomic numbers.

2.What are iso electronic ions? Give examples.

The ions of different elements having the same number of electrons ar called
isoelectronic ions.

Eg. Na^+, F^-, Mg^{2+}

3.What is effective nuclear charge?

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

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

4.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. The correct definition is "It is defined as the minimum amount of energy required to remove the most loosely bound electron from the valence shell of isolated neutral gaseous atom in its ground state."

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

$Mg + IE_1 \rightarrow Mg^+ + e^-$

$Mg^+ + IE_2 \rightarrow Mg^{2+} + e^-$

$Mg^{2+} + IE_3 \rightarrow Mg^{3+} + e^-$

The third step will have the highest ionization energy because it is easier to remove electron from a neutral atom than from a positive and dipositive ions. As the positive charge increases, the nuclear attraction also increases. More over Mg^{2+} has stable fully filled configuration hence more energy is required to remove electron from a stable atom. N$IE_1 < IE_2 < IE_3$

6.Define electronegativity.

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.

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

The total numbers of electrons are 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 ionisation energies, always greater than the previous one.

8.Energy of an electron in the ground state of the hydrogen atom is -2.8 x 10^{-18} J. Calculate the ionisation enthalpy of atomic hydrogen in terms of kJ mol^{-1}.

Ionisation energy is the amount of energy required to remove the electron from ground state to excited state.

$E_1 = -2.8 \times 10^{-18}$ J

$E_2 = 0$

$\Delta E = E_2 - E_1$

$= 0 - (-2.8 \times 10^{-18}$ J)

$= 2.8 \times 10^{-18}$ J

IE per mole of H atom = 2.8×10^{-18} J x 6.023×10^{23}

= 13.13×10^{5} J mol^{-1}

= 1313 KJ mol^{-1}.

9.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 atom affects the value of ionisation potential and electron gain enthalpy.
- Half- filled and completely filled electronic configurations are more stable than the partially filled configurations.
- Such atoms will have high ionization energy and very low or zero electron affinity.

Example: Be (z=4) $1s^2 2s^2$ (fully filled)

N (z=7) $1s^2 2s^2 2p^3$ (half filled)

Both have high ionization energy and zero electron affinity.

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

Z=118 [$_{86}$Rn] $5f^{14} 6d^{10} 7s^2 7p^6$

Period number – 7

Group number – 18

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

Fifth period (n=5) has nine orbitals – one 5s , five 4d , and three 6p orbitals.

Each orbital accommodates two electrons.

So totally 18 electrons. Hence there are 18 elements.

12.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?

Elements a and c belongs to group 18

Elements b and d belongs to group 13

13.Give the general electronic configuration of lanthanides and actinides.

Lanthanides:($4f^{1-14}, 5d^{0-1}, 6s^2$)

Actinides:($5f^{0-14}, 6d^{0-2}, 7s^2$)

14.Why halogens act as oxidising agents?

- Oxidising agents are the one which ready to accept electrons.

- Halogens have $ns^2\ np^5$ configuration. They are ready to accept one electron to attain stable noble gas configuration. So they act as oxidizing agents.

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

- Lithium and beryllium form more covalent compounds, unlike the alkali and alkaline- earth metals which predominantly form ionic compounds.
- The elements of the second period have only four orbitals (2s & 2p) in the valence shell and have a maximum co-valence of 4, whereas the other members of the subsequent periods have more orbitals in their valence shell and shows higher valences.
- For example, boron forms BF^{4-} and aluminium forms AlF_6^{3-}

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

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^-

$r(C^+)$, $r(A^-)$ are the radius of the cation and anion.

Pauling also assumed that the radius of the ion having noble gas electronic configuration is inversely proportional to the effective nuclear charge.

$r(C^+)\ \alpha\ 1\ /(Z_{eff})C^+$ -------------(2)

$r(A^-)\ \alpha\ 1\ /(Z_{eff})A^-$ -------------(3)

Where Z_{eff} is the effective nuclear charge

Dividing the equation 2 by 3

$$\frac{r(C+)}{r(A-)} = \frac{Zeff(A-)}{Zeff(C+)}$$

17.Explain the periodic trend of ionisation potential.

In a period,

- The ionisation energy usually increases along a period.
- along a period, the valence electrons are added to the same shell. 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

In a group,

- The ionisation energy decreases down a group.
- As we move down a group, 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

18.Explain the diagonal relationship.

The similarity in properties existing between the diagonally placed elements is called

Diagonal relationship'.

Li & Mg Be & Al B & Si

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

- Na (Z=11) and Mg(Z=12). Higher the nuclear charge greater the force of attraction between the nucleus and outermost electrons. Hence the first ionization energy of magnesium is greater than that of the sodium
- Na^+ has stable fully filled configuration so it is very difficult to remove the second electron from sodium atom whereas Mg^+ has one electron in their 3s orbital. Mg^+ is ready to give its electron to attain stable configuration.
- Hence second ionization energy of sodium is greater than that of magnesium.

20.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 A°

$r(K^+) + r(Cl^-) = d(K^+ - Cl^-) = 3.14$ Å

(1)

K^+ and Cl^- ions have Ar (Z=18) type configuration. The effective nuclear charge for

K^+ and Cl^- can be calculated as follows.

$K^+ = (1s^2)$ $(2s^2\ 2p^6)$ $(3s^2\ 3p^6)$

Innershell (n-1)th shell nth shell

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

$$\frac{r(K+)}{r(cl-)} = \frac{Zeff(cl-)}{Zeff(K+)} = \frac{5.75}{7.75} = 0.74$$

$r(K^+) = 0.74\ r(Cl^-)$ -------------- (2)

Substitute (2) in (1)

$0.74\ r(Cl^-) + r(Cl^-) = 3.14\ A^0$ --------- (3)

$1.74\ r(Cl^-) = 3.14\ A^0$

$r(Cl^-) = 3.14Å / 1.74 = 1.81A^0$

From (2) $r(K^+) = 0.74\ r(Cl^-) = 0.74 \times 1.81A^0$

$= 1.33\ A^0$

$r(K^+) = 1.33A^0$ $r(Cl^-) = 1.81A^0$

21. Explain the following, give appropriate reasons.

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

N (Z=7) $1s^2 2s^2 2p^3$

O (Z=8) $1s^2 2s^2 2p^4$

Since the half- filled electronic configuration in nitrogen is more stable than oxygen,

It requires higher energy to remove an electron from 2p orbital of nitrogen.

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

C(Z=6) B(Z=5)

Higher the nuclear charge, greater the force of attraction between nucleus and outermost electron. So first ionization energy of carbon is greater than boron.

C^+ ($1s^2\ 2s^2\ 2p^1$) B^+ ($1s^2\ 2s^2$)

B^+ has fully filled configuration which is more stable than C^+ which is partially filled. So the second ionization potential of Boron is greater than carbon.

(iii)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.

Be(Z=4)$1s^2\ 2s^2$,Mg (Z=12)$1s^2\ 2s^2\ 2p^6\ 3s^2$(noble gas configuration - $ns^2\ np^6$)

All these have completely filled configuration and are more stable where the addition of electron is not possible. So they have zero electron affinity.

N (Z=7) $1s^2\ 2s^2\ 2p^3$ P (Z=15) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

Both have half-filled configuration. So they are more stable. Hence have low electron affinity.

(iv)The formation of F^- (g) from F(g) is exothermic while that of O^{2-}(g) from O (g) is endothermic.

Fluorine is highly electronegative and gains one electron to attain stable configuration and releases energy; so exothermic

In oxygen the addition of first electron is exothermic but addition of second electron should overcome the high repulsive force so it needs external energy to accept the second electron. Hence endothermic.

22. What is screening effect?

The inner shell electrons act as a shield between the nucleus and the valence electrons.

This effect is called screening effect.

23.Briefly explain the basis for Pauling's scale of electronegativity.

He assigned arbitrary value of electro negativities for hydrogen and fluorine as 2.2 and 4.0 respectively. Based on this the electronegativity values for other elements can be calculated using the following expression

$$(x_A - x_B) = 0.182\ \sqrt{E_{AB}} - (E_{AA}*E_{BB})^{1/2}$$

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

__**In a Period,**__

The electronegativity 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.

In a group,

The electronegativity decreases down a group.

As we move down a group the atomic radius increases and the nuclear attractive force on the valence electron decreases.

ADDITIONAL QUESTIONS AND ANSWERS

(2 MARKS & 3 MARKS)

1. Define Triads.

- Elements like chlorine, Bromine and Iodine have same chemical properties in the group. These are called as Triads.
- In the triads the atomic weight of the middle element nearly equal to the mean of the atomic weight of the remaining two elements.

2. Define Law of Octaves

On arranging the elements in the increasing order of atomic weights, the property of every eighth elements is similar to the property of the first element.

3. Define Mendeleev's Periodic law

The properties of the elements are the periodic functions of the atomic weights.

4. Give the Anomalies of Mendeleev's Periodic table

- Elements with same properties are placed in different groups
- Elements with different properties are placed in same groups.
- Elements with higher atomic weights are placed before the lower atomic weights.

5. Define Periodicity

The repetition of the Physical and the chemical properties at regular intervals are
called periodicity

6. How Lavoisier classified the elements?

Acid making elements = Sulphur

Gas like elements = Oxygen

Metallic element = gold

Earthly elements = Lime

7. Explain the Henry Moseley X-ray spectra experiment

Henry Moseley studied the X-ray spectra of many elements by bombarding them with high energy electrons.

He observed a linear correlation between the atomic number and the frequency of the X-ray emitted.

$$\sqrt{V} = a(Z - b)$$

V= frequency of the X-ray Z = Atomic number

a and b = constants

- The plot of $\sqrt{V}$ against Z gives a straight line.

From the frequency the X-ray we can determine the atomic number of the known element

8. Explain the different types of elements in the Periodic table.

There are 4 types of elements

a)s-block elements

- The elements of group 1 and group 2 are called as s-block elements
- The last valence electron enters into the ns orbital.
- Group 1 elements are called as Alkali metals and group 2 are called as alkaline earth metals
- They are highly reactive
- They are soft metals with low boiling and melting point.

b) p-block elements

- The elements of group 13 to 18 are called p-block elements
- Their general electron configuration is $ns^2 np^{1-6}$
- Group 16 is called as Chalcogens.
- Group 17 is called as Halogens.
- Group 18 is called as Noble gases.
- They form covalent compounds.
- They have high electron affinity and Ionization energy values.

C) block elements

- The elements of group 3 to 12 are called d-block elements.
- Their general electron configuration is $ns^2 (n-1)d^{1-10}$
- They have high boiling and melting points.
- They are good conductor of heat and electricity.
- They are used as catalyst.

d) f-Block elements

- There are 2 series of f-block elements
- Lanthanides = $4f^{1-14} 5d^{0-1} 6s^2$
- Actinides = $5f^{0-14} 6d^{0-2} 7s^2$
- They have high melting points
- Most of the compounds are coloured
- They show variable oxidation states

9. Define atomic radius and explain the variation in the periodic table

The distance between the center of the nucleus and the outer most shell containing the valence electron called as atomic radius.

Along the Period: It Decrease along the period.

Reason:

- As we move along the period the valence electrons are added to the same shell.
- So the Nuclear charge increases,
- And the attractive force between the valence electron and the nucleus increases

Down the Group: It Increases along the group.

Reason:

- As we move down the group the valence electrons are added into new shells.
- As a result the distance between the nucleus and the valence electrons increases.

10. Define ionic radius

The distance between the center of the nucleus of a ion and the outermost shell containing the valence electron called as ionic radius.

11. Define metallic radius

The half of the distance between two adjacent metal atoms in a closely packed metallic crystal is called as metallic radius.

12. Define covalent radius.

The half of the inter nuclear distance between two identical atoms linked by a single covalent bond is called as Covalent radius

13. Define Valence State.

- Valence state is the number of electrons present in the valence shell
- Along the group the Valence state remains a constant
- Along the period the Valence State increases.

14. Define second ionization energy

Second Ionization energy is the amount of energy required to remove an electron from an unipositive cation.

15. Define electron affinity and explain the variation in the periodic table.

Electron affinity is defined as the amount of energy released when anelectron is added to the valence shell of an atom.

Down the Group: It decreases along the group.

Reason:

- As we move down the group the nuclear charge decreases.
- The atomic size increases.
- The Shielding effect of the inner electrons increase.

Along the period: It increases along the period

Reason:

- As we move along the period the nuclear charge increases
- The atomic size decreases.
- The attraction between the valence electron and the nucleus increases

16. Why Beryllium has high Ionisation energy than Boron ?

- Beryllium is having high ionization energy than Boron
- Beryllium has Stable Fully filled ns^2 electronic configuration
- The force of attraction between the nucleus and the outermost electron is very high in Beryllium
- So it is difficult to remove the outermost electrons of Beryllium
- Be (Z = 4) $1s^2 2s^2$ B (Z = 5) $1s^2 2s^2 2p^1$

17. Why Nitrogen has high Ionisation energy than oxygen?

Nitrogen has Stable Half- filled electronic configuration

The force of attraction between the nucleus and the outermost electron is very high in Nitrogen

So it is difficult to remove the 2p electrons of Nitrogen

N(Z = 7) $1s^2 2s^2 2p^3$

O(Z = 8) $1s^2 2s^2 2p^4$

18. Why the 17th group (Halogens) have high electron affinity ?

- Halogens have a unstable np^5 electronic configuration.
- By gaining one electron it becomes a Stable Fully filled np^6electronic configuration
- Hence it accepts one electron and become a stable Noble gas configuration

19. Why the electron affinity of Oxygen and Fluorine is lower than sulphur and chorine?

- Oxygen and Fluorine is smaller in size.
- Oxygen and Fluorine have high electron density.
- In oxygen and Fluorine the valence electrons enter into the 2p orbital.
- But in Sulphur and chlorine the valence electrons enter into the 3p orbital.

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