Classification of Elements and Periodicity in Properties

CHAPTER 3

EXAM DRILL

ANSWERS

1. **(d) :** Ionization enthalpy : C < O < N
Ionic size : $Mg^{2+} < Na^+ < F^-$

2. **(c) :** d-Block, because last electron enters in $5d$-orbital.

OR

(b) : Bohrium has atomic number 107.

3. **(b) :** When a neutral atom gain an electron, energy is released whereas when anion gain an electron, energy is absorbed.

4. **(b) :** There are 18-groups and 7-periods in periodic table.

5. **(d) :** Alkali metals have lowest value of ionisation enthalpy in a period. On moving down the group from top to bottom, the ionisation enthalpy decreases. Hence from the graph, M has least ionisation enthalpy.

6. Helium

7. 11.2

8. Basic oxide

9. Magnesium

10. Untrinilium

11. **(a)**

12. **(d) :** Chlorine has more negative electron gain enthalpy than fluorine. Additional electrons are repelled less effectively by $3p$-electrons in Cl atom than by $2p$-electrons in F atom.

13. **(c) :** Helium is the smallest inert gas. Ne has most positive electron gain enthalpy.

14. **(a)**

15. **(d) :** The correct order is F > Cl > N.

16. This is because Be and Mg have stable electronic configuration of ns^2. Thus, they don't have tendency to take an additional electron. The incoming electron enters the much higher energy p-orbitals of the valence shell. Thus, energy is required to force the electron in their atoms and therefore, electron gain enthalpies are positive.

OR

Elements A and B belongs to the same group because both have same number of valence electrons. A,B, C and D belongs to p-block because last electron enters to p-orbitals. Element E belongs to s-block as last electron enters to s-orbital.

17. (i) Silicon has valency 4 and bromine has valency 1. Therefore the formula of compound is $SiBr_4$.

(ii) Aluminium has valency 3 and oxygen has valency 2, hence the formula of compound is Al_2O_3.

18. Sodium belongs to period-3 with outer electronic configuration $3s^1$ and potassium belongs to period-4 with outer electronic configuration $4s^1$. After loosing valence electron, these ions have configuration $2s^22p^6(Na^+)$ and $3s^23p^6(K^+)$. The addition of principal shell (n = 3) causes increase in size of K^+. Therefore, K^+ is larger in size than Na^+.

19. (i) The elements after uranium (Z = 92) in periodic table are known as transuranic elements.

(ii) Periodicity is caused by the repetition of similar outer electronic configuration of the atoms in the valence shell after certain regular intervals.

20. (i) Increase in shielding effect decreases effective nuclear charge *i.e.*, the force of attraction between valence electron and nucleus, which makes removal of outer electron easier. Thus, ionisation enthalpy decreases with increasing shielding effect.

(ii) The second ionisation enthalpy of group-1 element is very high because, after losing outer electron they attain stable noble gas configuration ($ns^2\ np^6$). Therefore, they do not form dipositive ions.

OR

(i) It helped in systematic study of chemistry of elements correction of atomic masses of various elements and prediction of new elements.

(ii) Keeping his primary aim of arranging the elements of similar properties in the same group, Mendeleev proposed that some of the elements were still undiscovered and therefore, he left several gaps in the periodic table.

21. (i) 'X' has valence electron four and belongs to 3rd period hence, it is silicon (Si). Its electronic configuration is $1s^22s^22p^63s^23p^2$. Its atomic number is 14.

(ii) Noble gases contains all paired electrons and their subshells are completely filled. Hence, their electron gain enthalpy is almost zero.

(iii) Sodium has eleven electrons and eleven protons but number of protons in Mg^+ are twelve, though it has eleven electrons. Due to higher effective nuclear charge in case of Mg^+, removal of electron from it requires more energy.

22. (i) The electronic configuration of S(Z = 16) is [Ne] $3s^2 3p^4$. It can accommodate two additional electrons in its $3p$-orbital. The addition of first electron is an exothermic process.

i.e., $S_{(g)} + e^- \longrightarrow S^-_{(g)}; \Delta_{eg}H_1 = -200 \text{ kJ mol}^{-1}$

Now, second electron has to be added to an anion which is difficult due to same charge repulsions. Therefore, some energy has to be supplied to overcome these repulsion, *i.e.*, the addition of second electron to S^- is an endothermic process.

$S^-_{(g)} + e^- \longrightarrow S^{2-}_{(g)}; \Delta_{eg}H_2 = +590 \text{ kJ mol}^{-1}$

(ii) Na^+, Mg^{2+} and Al^{3+} are isoelectronic ions. Among isoelectronic ions, ionic radius decrease with the increase in the magnitude of nuclear charge. Hence, the order is

$$Na^+ > Mg^{2+} > Al^{3+}$$

OR

$A(Z = 16) \Rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

Since, last electron enters in $3p$-orbital,

$\therefore$ period =3, block = p

and for p-block, group number

= 10 + number of valence electrons = 10 + 6 = 16

$B(Z = 37)\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^1$

Since, last electron enters in $5s$-orbital,

$\therefore$ period = 5, block = s

and for s-block, group number = number of valence electrons = 1

23. The oxidation state of an element is based on its electronic configuration. The various oxidation states of a transition metal element is due to the involvement of $(n-1)d$ and outer ns electrons in bonding.

For example, Ti(22) : $3d^2 4s^2$ can show three oxidation states (+2, +3 and +4) in various of its compounds like $TiO_2(+4)$, $Ti_2O_3(+3)$ and $TiO(+2)$.

The non-transition elements mainly the p-block elements can show a number of oxidation states from $+n$ to $(n-8)$ where, n is the number of electrons present in the outermost shell *e.g.*, phosphorus can show –3, + 3 and + 5 oxidation states. Lower oxidation states are ionic as the atom accepts one or more electrons to achieve stable configuration while higher oxidation states are achieved by unpairing the paired electrons and shifting the electrons to vacant d-orbitals.

24. (i) : $H_{(g)} + 13.6 \text{ eV} \longrightarrow H^+_{(g)}$

1 eV = 96.49 kJ mol^{-1}

$\Rightarrow$ 13.6 eV = 13.6 × 96.49 = 1312.3 kJ mol^{-1}

$\therefore$ Energy required to convert 1 g of hydrogen atoms into ions = 1312.3 kJ

Hence, energy required to convert 2.5 g of hydrogen atoms into ions = 1312.3 × 2.5 = 3280.75 kJ

(ii) Metallic character increases as we move from top to bottom in a group. This is because the atomic size increases and removal of electron from outer shell becomes easier *i.e.*, electropositive character increases down the group.

25. As we move from left to right, the acidic nature of oxide increases whereas basic nature of oxide decreases *i.e.*, metallic oxides are basic in nature while non-metallic oxides are acidic in nature. Oxides of the elements in the centre are amphoteric or neutral in nature. For example, Al_2O_3 is amphoteric while CO, NO etc. are neutral in nature.

Na_2O reacts with water and form a strong base. Thus, it is a basic oxide.

$Na_2O + H_2O \longrightarrow 2NaOH$

Cl_2O_7 reacts with water and form a strong acid. Thus, it is an acidic oxide.

$Cl_2O_7 + H_2O \longrightarrow 2HClO_4$

26. (i) Mendeleev arranged elements in horizontal rows and vertical columns of a table in order of their increasing relative atomic weights in such a way that the elements with similar properties occupied the same vertical column or group.

(ii) Nitrogen has positive electron gain enthalpy due to extra stability of half-filled orbitals ($2s^2\ 2p^3$). Oxygen has four electrons in outermost shell ($2s^2\ 2p^4$) and it can easily accept two electrons hence, its electron gain enthalpy is negative. Due to stable configurations, ionisation enthalpy of nitrogen is higher than oxygen. It is easier to remove one electron from $2p^4$ as compared to $2p^3$.

(iii) Newlands's law of octaves seemed to be true only for elements upto calcium.

OR

(i) (a) Due to the presence of fully filled s-orbitals in Mg, it is difficult to pull out electron from stable configuration and therefore, Mg has higher first ionisation energy than Al.

(b) First member of each group of s and p-block elements shows anomalous behaviour due to

– small size
– high ionisation enthalpy
– high electronegativity
– absence of d-orbitals.

(ii) I is least electronegative element among F, Cl, Br and I since electronegative character decreases on moving down a group.

27. (i) (a) **Main characterstics of s-block elements :**

I. They are soft metals and have low melting and boiling points.

II. They are highly electropositive and have low values of ionisation energies.

III. They are very reactive and readily form ionic compounds by losing one or more electrons.

IV. They impart characteristic colours to flame except Be and Mg.

V. Their oxides are basic in nature.

(b) **Main characteristics of *d*-block elements :**

I. They are comparatively hard metals and have high melting and boiling points.

II. They are less electropositive than *s*-block elements and have low value of ionisation energies.

III. Most of them shows paramagnetism.

IV. Most of the transition elements and their compounds act as a catalyst.

V. They have high tendency to form complex compounds and alloys.

(ii) Due to repulsions between ion N^- and the incoming electron, electron gain enthalpy of N^- is positive while that of P is negative.

OR

(i) (a) Potassium (having atomic mass 39.10) has been placed after argon (having atomic mass 39.44) because in modern periodic table elements are arranged in increasing order of their atomic numbers. Since, atomic number of argon is 18 and that of potassium is 19, therefore potassium is placed after argon.

(b) The valence shell electronic configuration of halogen is ns^2np^5. They need only one electron to attain stable noble gas configuration. Thus, they can easily accommodate the incoming electron with the liberation of high amount of energy.

(ii) Let, electronegativity of carbon be χ_C.

Given, $E_{H-H} = 104.2$ kcal mol^{-1}, $E_{C-C} = 83.1$ kcal mol^{-1}, $E_{C-H} = 98.8$ kcal mol^{-1}, $\chi_H = 2.1$

According to Pauling equation,

$$\chi_C - \chi_H = 0.208\left[E_{C-H} - \sqrt{(E_{C-C} \times E_{H-H})}\right]^{\frac{1}{2}}$$

$$\chi_C - 2.1 = 0.208\left[98.8 - \sqrt{(83.1 \times 104.2)}\right]^{\frac{1}{2}}$$

$$\chi_C - 2.1 = 0.208 \times [98.8 - 93.05]^{1/2}$$

$$\chi_C = 0.208\sqrt{5.75} = 0.5$$

$$\chi_C = 0.5 + 2.1 = 2.6.$$

28. (i) **Electronegativity :** Electronegativity is the tendency of an atom to attract the shared pair of electrons toward itself. On moving from left to right in a period, atomic size decreases and nuclear charge increases. As a result, tendency to attract bonded electron pair increases *i.e.*, electronegativity increases. Down the group atomic size as well as screening effect increases. As a result, tendency to attract the bonded electron pair decreases *i.e.*, electronegativity decreases.

(ii) **Covalent radii :** Variation of covalent radii depends on nuclear charge and number of main energy levels of an atom. On moving from left to right in a period, covalent radii decreases because nuclear charge increases progressively by one unit and additional electron goes to same principal shell. On the other hand, it increases down the group. This is because additional electron enters a new principal shell which results in lesser attraction between nucleus and valence electron.

(iii) **Non-metallic character :** Non-metallic character of an element is the measure of its electronegativity. Since electronegativity increases along a period and decreases down the group, thus non-metallic character of elements also increases along the period and decreases down the group.

(iv) **Ionisation enthalpy :** It is inversely proportional to atomic radii. Larger the atomic radii, lesser will be ionisation enthalpy and vice-versa. Since, atomic radii decreases along the period, thus ionisation enthalpy increases along the period. Down the group, atomic radii increases, therefore ionisation enthalpy decreases down the group.

(v) **Ionic radii :** Ionic radii decreases along a period because nuclear charge increases, which results in strong force of attraction between nucleus and valence shell electron. It increases down the group because effective nuclear charge decreases which results in weak force of attraction between nucleus and valence shell electron.

29. (i) (a) **Isoelectronic ions :** Isoelectronic ions are the ions of different elements which have same number of electrons but differ from one another in magnitude of the nuclear charge. The size of species in isoelectronic series can be compared by comparing their nuclear charges. As nuclear charge increases, the electrons are pulled more and more strongly which in turn decreases the atomic size.

For example, for isoelectronic series N^{3-}, F^-, Na^+, O^{2-}, Mg^{2+}, the increasing order of size is

$$Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$$

Nuclear charges +12 +11 +9 +8 +7

(b) **Valency :** The number of electrons present in valence shell of an element are called valence electrons. These valence electrons determine the valency of an atom. For representative elements (*s*-block and *p*-block elements), valency is generally equal to either number of valence electrons or equal to eight minus the number of valence electrons. *d*-block and *f*-block elements exhibit variable valency. Along the period, it first increases from 1 to 4 and then decreases from 4 to zero. In a group, generally all elements exhibit same valency because all of them have same number of valence electrons.

(ii) The inert gas is krypton. Its atomic number is 36. It belongs to 4th period of the periodic table. Its electronic configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$

Total number of *s*-electrons = 8

Total number of *p*-electrons = 18

Total number of *d*-electrons = 10

OR

(a) The element having electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^1$ has least *I.E.* because it has large atomic radius and just one electron in its valence shell. Also, on loosing one electron it can attain stable noble gas configuration ($1s^22s^22p^6$). Since, it has 11 electrons therefore its atomic number is 11. Thus, the element is sodium.

(b) The element having electronic configuration $1s^22s^22p^5$ is a halogen because it requires only one electron to attain noble gas configuration. As it has 9 electrons that means its atomic number is 9. Hence, it is fluorine.

(c) Element with electronic configuration $1s^22s^2$ is an alkaline earth metal. Since it has 4 electrons, therefore its atomic number is 4 and the element is beryllium.

(d) Element with electronic configuration $1s^22s^22p^6$ belongs to group 18 because it has completely filled valence shell. Since, it has 10 electrons, therefore, its atomic number is 10 and the element is neon.

(e) Element with electronic configuration $1s^22s^22p^63s^23p^1$ has valency 3. Since it has 13 electrons that means atomic number is 13 therefore, the element is aluminium.

30. (i) Na^+ and Ne are isoelectronic as they have same number of electrons. But Na has 11 protons and Ne has 10. Therefore Na^+ has a greater effective nuclear charge and thus, has higher ionisation enthalpy.

(ii) The step Al^{2+} to Al^{3+} will have highest ionisation enthalpy.

(iii) Lithium

(iv) Energy = $1 \times (24.58 + 54.4) \times 96.49$ kJ = 7620.78 kJ

(v) Atom having electronic configuration $1s^22s^22p^4$ will have higher second ionisation enthalpy.