3 Classification of Elements and Periodicity in Properties

LESSON AT A GLANCE

- **Mendeleev's Periodic Law.** Physical and chemical properties of elements are periodic function of their atomic masses.
- **Modern Periodic Law.** Physical and chemical properties of the elements are periodic function of their atomic numbers.
- **Groups.** There are 18 groups. These are vertical rows.
- **Periods.** There are 7 periods. These are horizontal rows.
- **Representative Elements.** The *S* and *P* block of elements are known as representative elements.
- **Transition Elements.** They are also called *d*-block elements. They have general electronic configuration $(n-1)d^{1-10}ns^{0-2}$.
- **Inner Transition Elements.** Lanthanoids (the fourteen elements after Lanthanum) and actinides (the fourteen elements after actinium) are called inner transition elements. General electronic configuration is $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$.

  They are also called *f*-block elements.
- **Metals.** Present on the left side of the periodic table. Comprise more than 78% of the known elements.
- **Non-metals.** Mostly located on the right hand side of the periodic table.
- **Metalloids.** Elements which line as the border line between metals and non-metals (e.g., Si, Ge, As) are called metalloids or semimetals.
- **Atomic Radii and Ionic Radii.** increase down the group decrease along the period.

- **Ionization Enthalpy.** Increases along the period and decreases down the group.
- **Noble Gas Elements.** Elements with symmetrical configuration are chemically inert in nature.
- **Electric Nuclear Charge.** Z = Nuclear charge - Screening constant.
- **Electronegativity.** Increases along a period decreases down the group.
- **Chemical Reactivity.** Chemical reactivity is highest at the two extremess of a period and lowest in the centre.
- **Oxides of Elements.** Oxides formed of the Elements on the left are basic and of elements on the right are acidic in nature.

  Oxides of elements in the centre are amphoteric or neutral.

TEXTBOOK QUESTIONS SOLVED

Q1. *What is the basic theme of organisation in the periodic table?*

Ans. The basic theme of the organisation of elements in the periodic table is to study their chemical and physical behaviour in a systematic way. To do so, elements are arranged in groups (*i.e.*, vertical rows) and periods (*i.e.*, horizontal rows) according to the similarities in chemical and physical properties of a very large number of their chemical compounds.

Q2. *Which important property did Mendeleev use to classify the elements in his periodic table and did he stick to that?*

Ans. Mendeleev used atomic weight as the basis of classification of elements in the periodic table. He did stick to it and classify elements into groups and periods.

Q3. *What is the basic difference in approach between the Mendeleev's Periodic Law and the Modern Periodic Law?*

Ans. Mendeleev's Periodic Law states that chemical and physical properties of elements are the periodic function of the **atomic weights** (or atomic masses) whereas the Modern Periodic Law states that the chemical and physical properties of the elements are the periodic function of their **atomic numbers**. The modern periodic law takes care of all the flaws of Mendeleev's periodic law *e.g.*, the position of isotopes.

Q4. *On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements.*

Ans. The sixth period corresponds to sixth shell. The orbitals present in this shell are 6s, 4f, 5p, and 6d. The maximum number of electrons which can be present in these sub-shell is 2 + 14 + 6 + 10 = 32. Since the number of elements in a period corresponds to the number of electrons in the shells, therefore, sixth period should have a maximum of 32 elements.

Q5. *In terms of period and group where would you locate the element with Z = 114?*

Ans. Period = 7^{th}; Group No. = 14; Block = *p*.

Q6. *Write the atomic number of the element present in the third period and seventeenth group of the periodic table.*

Ans. Third period comprises of 3*s* and 3*p* orbitals, respectively. There are two electrons in 3*s* and six in 3*p* orbitals. On the basis of the electronic configuration, $3p^1$, $3p^2$, $3p^3$, $3p^4$, $3p^5$ and $3p^6$, we have the groups, 13, 14, 15, 16, 17 and 18, respectively. The configuration $3p^5$ will correspond to Group 17 of the periodic table. The electronic configuration till $3p^5$ is as: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$.

The total number of electrons here gives the **atomic number** (2 + 2 + 6 + 2 + 5 = 17). Therefore, the element has the atomic number as **17**.

Q7. *Which element do you think would have been named by*

(*i*) *Lawrence Berkeley Laboratory*

(*ii*) *Seaborg's group?*

Ans. (*i*) Lawrencium (*Z* = 103, Lr) and Berkelium (*Z* = 97, Bk), from Lawrence Berkeley Laboratory.

(*ii*) Seaborgium (*Z* = 106, Sg).

Q8. *Why do elements in the same group have similar physical and chemical properties?*

Ans. Because of similar electronic configurations especially that of valence shell.

Q9. *What does atomic radius and ionic radius really mean to you?*

Ans. **Atomic radius.** The distance from the centre of nucleus to the outermost shell of electrons in the atom of any element is called its atomic radius. It refers to both covalent or metallic radius depending on whether the element is a non-metal or a metal.

Ionic radius. The Ionic radii can be estimated by measuring the distances between cations and anoins in ionic crystals.

Q10. *How do atomic radius vary in a period and in a group? How do you explain the variation?*

Ans. Within a group Atomic radius increases down the group.

Reason. This is due to continuous increases in the number of electronic shells or orbit numbers in the structure of atoms of the elements down a group.

Variation across period.

Atomic Radii. From left to right across a period atomic radii generally decreases due to increase in effective nuclear charge from left to right across a period.

Q11. *What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions.*

(*i*) F^- (*ii*) Ar

(*iii*) Mg^{2+} (*iv*) Rb^+

Ans. Isoelectronic species are those species (atoms/ions) which have same number of electrons. Following are the isoelectronic species.

(*i*) F^-(10 electrons, $Z = 9$, one negative charge so $9 + 1 = 10$).
Isoelectronic species: ${}_7N^{3-}$, ${}_8O^{2-}$, ${}_{10}Ne$, ${}_{11}Na^+$, ${}_{12}Mg^{2+}$ and ${}_{13}Al^{3+}$.

(*ii*) Ar(18 electrons). Isoelectronic species are:
${}_{17}Cl^-$, ${}_{19}K^+$, ${}_{20}Ca^{2+}$.

(*iii*) Mg^{2+} (10 electrons). Isoelectronic species as given in F^- case.

(*iv*) Rb^+ (37–1 = 36 electrons). Isoelectronic species are: ${}_{35}Br^-$, ${}_{36}Kr$ and ${}_{38}Sr^{2+}$.

Q12. *Consider the following species:*

$$N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+}, \text{and } Al^{3+}$$

(*a*) *What is common in them?*

(*b*) *Arrange them in the order of increasing ionic radii.*

Ans. (*a*) All of them are iso-electronic in nature and have 10 electrons each.

(*b*) In iso-electronic species, greater the nuclear charge, lesser will be the atomic or ionic radius.

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

Q13. *Explain why cations are smaller and anions larger in radii than their parent atoms?*

Ans. A cation is smaller than the parent atom because it has fewer electrons while its nuclear charge remains the same. The size of anion will be larger than that of parent atom because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge.

Q14. *What is the significance of the terms 'isolated gaseous atom' and 'ground state' while defining the ionisation enthalpy and electron gain enthalpy?*

Ans. In the definitions for both ionisation enthalpy and electron gain enthalpy emphasis is made on the terms **isolated gaseous atom** and that too in its **ground state**. The importance of these terms is as under:

Isolated gaseous atom: In both the cases, the electrons are either removed or added in the atoms. If the atoms are not separated to a large enough distance, they will be influenced by coulombic forces (**attractive force** due to +ve nucleus and –ve electron cloud, and **repulsive force** due to similar charges). In order to avoid these forces, atoms are taken to be in the isolated gaseous state. Therefore, for the sake of comparison, we stick to one definition of isolated gaseous state.

Ground state: The term 'ground state' implies that the atom must be present in the most stable state, *i.e.*, of minimum energy. The atom must not be in the excited state and all the electrons must be accommodated in various shells in accordance with the aufbau principle.

Q15. *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 J mol^{-1}.*

Ans. The ionisation enthalpy in kJ mol^{-1}

$$= -(\text{ionisation energy in kJ mol}^{-1})$$

$\therefore$ The ionisation enthalpy

$$= -(-2.18 \times 10^{-18} \times 6.023 \times 10^{23} \times 10^{-3} \text{ kJ mol}^{-1})$$
$$= 1313.014 \text{ kJ mol}^{-1}$$

16. *Among the second period elements the actual ionisation enthalpies are in the order Li < B < Be < C < O < N < F < Ne.*

Explain why

(*i*) *Be has higher $\Delta_i H$ than B*

(*ii*) *O has lower $\Delta_i H$ than N and F?*

Ans. (*i*) In case of Be ($1s^2\ 2s^2$) the outer-most electron is present in 2*s*-orbital while in *B* ($1s^2\ 2s^2\ 2p^1$) it is present in 2*p*-orbital. Since 2*s* - electrons are more strongly attracted by the nucleus than 2*p*-electrons, therefore, lesser amount of energy is required to knock out a 2*p*-electron than a 2*s* - electron. Consequently, $\Delta_i H_1$ of *Be* is higher than that $\Delta_i H_1$ of *B*.

(*ii*) The electronic configuration of

$$N_7 = 1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$$
$$O_8 = 1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$$

We can see that in case of nitrogen 2*p*-orbitals are exactly half filled. Therefore, it is difficult to remove an electron from *N* than from 0. As a result $\Delta_i H$ of *N* is higher than that of 0.

Q17. *How would you explain the fact that the first ionisation enthalpy of sodium is lower than that of magnesium but its second ionisation enthalpy is higher than that of magnesium?*

Ans. Electronic configuration of Na and Mg are

$$\text{Na} = 1s^2\ 2s^2\ 2p^6\ 3s^1$$
$$\text{Mg} = 1s^2\ 2s^2\ 2p^6\ 3s^2$$

First electron in both cases has to be removed from 3*s*-orbital but the nuclear charge of Na (+ 11) is lower than that of Mg (+ 12) therefore first ionization energy of sodium is lower than that of magnesium.

After the loss of first electron, the electronic configuration of

$$\text{Na}^+ = 1s^2\ 2s^2\ 2p^6$$
$$\text{Mg}^+ = 1s^2\ 2s^2\ 2p^6\ 3s^1$$

Here electron is to be removed from inert (neon) gas configuration which is very stable and hence removal of second electron requires more energy in comparison to Mg.

Therefore, second ionization enthalpy of sodium is higher than that of magnesium.

Q18. *What are the various factors due to which the ionisation enthalpy of the main group elements tends to decrease down a group?*

Ans. **Atomic size.** With the increase in atomic size, the number of electron shells increase. Therefore, the force that binds the electrons with the nucleus decreases. The ionization enthalpy thus decreases with the increase in atomic size.

Screening or shielding effect of inner shell electron. With the addition of new shells, the number of inner electron shells which shield the valence electrons increases. As a result, the force of attraction of the nucleus for the valence electrons further decreases and hence the ionization enthalpy decreases.

Q19. *The first ionisation enthalpy values (in kJ mol^{-1}) of Group 13 elements are:*

B	*Al*	*Ga*	*In*	*Tl*
801	*577*	*579*	*558*	*589*

How would you explain this deviation from the general trend?

Ans. The first ionisation enthalpies (kJ mol^{-1}) of Group 13 elements suggests that the normal

B	Al	Ga	In	Tl
801	577	579	558	589

group trend is not followed. From B to Al, the value decreases according to the group trend and afterwards it is erratic, *i.e.*, from Al to Ga it increases very slightly, and upto In it decreases and again increases to Tl. On inspecting the outer electronic configuration of Ga, In and Tl, it is observed that there are $3d^{10}$ electrons in the inner shell of Ga, $4d^{10}$ in In while there are $4f^{14}$ and $5d^{10}$ electrons in the case of Tl, respectively. *d* and *f* electrons shield or screen the nuclear charge poorly thereby increasing the Z_{eff} for the outer electrons in the valence shell. Thus, increase in the value of the ionisation enthalpy in Ga is due to this increase in Z_{eff} due to $3d^{10}$ electrons. Since the Z_{eff} situation is similar in In(due to $4d^{10}$ electron), there is decrease in the value in accordance with the group trend. In the case of Tl, there are 14 4*f*-electrons present in addition to the 10 5*d*-electrons. In this situation, Z_{eff} increases to a much greater extent for the outer electrons. Hence, the group trend is reversed and the ionisation enthalphy increases instead of decreasing.

Note: The above explanation also holds good for the similar trend in the atomic sizes of these elements.

Q20. *Which of the following pairs of elements would have a more negative electron gain enthalpy?*

(*i*) *O or F*

(*ii*) *F or Cl*

Ans. (*i*) The first electron gain enthalpy for F would be more negative than that for O atom. In fact, the values are: F = –328 kJ mol^{-1} and O = –141 kJ mol^{-1}. These values suggest that the formation of F^- ion is energetically more favourable as compared to O^- ion. This is due to the fact that F^- ion attains the stable noble gas electronic configuration while O^- ion is still short of one electron. Further, in going from O to F, in the same period, the size for F atom decreases and its nuclear charge increases. The combined effect of these two factors helps in accommodating the incoming electron.

(*ii*) The electron gain enthaply for Cl is more negative as compared to F (F = –328 kJ mol^{-1}, Cl = –349 kJ mol^{-1}). This irregularity is attributed to the very high inter-electronic repulsive forces that are in operation in the first member of the group, *i.e.*, F than the larger sized Cl atom.

Q21. *Would you expect the second electron gain enthalpy of O as positive, more negative or less negative than the first? Justify your answer.*

Ans. For oxygen atom:

$$O\ (g) + e^- \rightarrow O^-\ (g)\ (\Delta\ \text{eg}\ H = -\ 141\ \text{kJ mol}^{-1})$$

$$O^-\ (g) + e^- \rightarrow O^{2-}\ (g)\ (\Delta\ \text{eg}\ H = +\ 780\ \text{kJ mol}^{-1})$$

The first electron gain enthalpy of oxygen is negative because energy is released when a gaseous atom accepts an electron to form monovalent anion. The second electron gain enthalpy is positive because energy is needed to overcome the force of repulsion between monovalent anion and second incoming electron.

Q 22. *What is the basic difference between the terms electron gain enthalpy and electronegativity?*

Ans. The basic difference lies in the fact that the electron gain enthalpy is the property of an isolated atom in the gaseous state while electronegativity is the property of an atom in the combined state. Electronegativity refers to the tendency of an atom to pull an electron pair towards itself in a covalent bond. Electron gain enthalpy involves in the formation of –ve ions (anions) in the gaseous state.

Q23. *How would you react to the statement that the electronegativity of N on Pauling scale is 3.0 in all the nitrogen compounds?*

Ans. The statement that 'electronegativity of N in all its compounds is 3.0' in all nitrogen compounds is not entirely correct. Factors like hybridisation, type of substituents and oxidation number of the atom govern the electronegativity. For example, the electronegativity of N in $:NH_3$ (N is sp^3 hybridised) would be different from N in NO where N is *sp* hybridised. *s* character in *sp* hybrid orbitals is more than that in sp^3 orbitals. Electronegativity increases with an increase in *s* character of a hybrid orbital, *i.e.*, $sp > sp^2 > sp^3$.

Note: To see the effect of substituents on electronegativity of an element, the concept of *group electronegativity* is used. For example, instead of talking of electronegativity of *only* one atom, we consider the group as a whole. In this regard, CF_3 group is more electronegative than CH_3 group.

Q24. *Describe the theory associated with the radius of an atom as it*

(*a*) *gains an electron*

(*b*) *loses an electron.*

Ans. (*a*) Gain of an electron leads to the formation of an anion. The size of an anion will be larger than that of the parent atom because the addition of one or more electrons would result in increased repulsion among electrons and decrease in effective nuclear charge.

Thus the ionic radius of fluoride ion (F^-) is 136 pm whereas atomic radius of Fluorine (*F*) is only 64 pm.

(*b*) Loss of an electron from an atom results in the formation of a cation. A cation is smaller than its parent atom because it has fomer electrons while its nuclear charge remains the same. For example, the atomic radius of sodium (Na) is 186 pm atomic radius of sodium ion (Na^+) = 95 pm.

Q25. *Would you expect the first ionisation enthalpies for two isotopes of the same element to be the same or different? Justify your answer.*

Ans. The first ionisation enthalpy values of the two isotopes of the same element would be the same since they differ only in the number of neutrons, the number of protons and electrons remains the same. This way, the Z_{eff} for the electron to be removed remains the same.

Q26. *What are the major differences between metals and non-metals?*

Ans. The characteristic property that distinguishes metals and non-metals is conduction of electric current. Metals conduct electricity while non-metals are non-conductors (except graphite). Metals are malleable and ductile while solid non-metals are neither malleable nor ductile. Metals form cations while non-metals anions. Metals have low ionisation and low electron gain enthalpies and thus are good reducing agents. Metals form basic oxides while non-metals form acidic oxides.

Q27. *Use the periodic table to answer the following questions:*

(*a*) *Identify an element with five electrons in the outer subshell.*

(*b*) *Identify an element that would tend to lose two electrons.*

(*c*) *Identify an element that would tend to gain two electrons.*

(*d*) *Identify the group having metal, non-metal, liquid as well as gas at the room temperature.*

Ans. (*a*) Halogens, Group 17 elements with outer electronic configuration as np^5 (from $ns^2\ np^5$). For example F, Cl, Br and I.

(*b*) Group 2 elements with ns^2 electronic configuration, *e.g.*, Mg, Ca, Sr and Ba.

(*c*) Group 16 elements—oxygen group with $ns^2\ np^4$ electronic configuration, *e.g.*, O and S.

(*d*) It is Group 17—the halogens.

F_2, Cl_2 and Br_2 are non-metals. F_2 and Cl_2 are gases at room temperature, Br_2 is a liquid at room temperature and I_2 is solid at room temperature and shows metallic properties.

Q28. *The increasing order of reactivity among Group 1 elements is Li < Na < K < Rb < Cs whereas that among Group 17 elements is F > Cl > Br > I. Explain.*

Ans. The order of reactivity among Group 1 members Li < Na < K < Rb < Cs is due to their low ionisation enthalpies which decrease down the group.

The reactivity order among Group 17 members F < Cl < Br < I is due to their high values of electron gain enthalpies and also high values of standard electrode potentials (E^0_{red} values) which together facilitate the formation of anions, *i.e.*,

$$X_2(g) + 2e^- \longrightarrow 2X^-(g)$$

Formation of anions and also the reactivity of the parent halogens decreases down the group as both the factors show decreasing trend.

Q29. *Write the general outer electronic configuration of s-, p-, d- and f-block elements.*

Ans. (*i*) *s*-Block elements: ns^{1-2}; $n = 2–7$

(*ii*) *p*-Block elements: $ns^2\ np^{1-6}$; $n = 2–6$

(*iii*) *d*-Block elements: $(n-1)d^{1-10}\ ns^{1-2}$; $n = 4–7$

(*iv*) *f*-Block elements: $(n-2)f^{1-14}\ (n-1)d^{0-1}\ ns^2$; $n = 6–7$

Q30. *Assign the position of the element having outer electronic configuration*

(*i*) ns^2np^4 *for* $n = 3$

(*ii*) $(n-1)d^2ns^2$ *for* $n = 4$ *and*

(*iii*) $(n-2)f^7\ (n-1)d^1ns^2$ *for* $n = 6$, *in the periodic table.*

Ans. (*i*) For $ns^2\ np^4$ when $n = 3$

The complete electronic configuration is

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$ or [Ne] $3s^2\ 3p^4$

Total number of electrons = 16

From the information given, one can infer that for the given element as:

Period = 3rd since $n = 3$

From the outer electronic configuration, $ns^2\ np^4$, it is a ***p*-block element** with 16 as its atomic number

Group number = 10 + No. of valence electrons

= 10 + 6 = 16

= Oxygen group

The element is sulphur (S).

(*ii*) $(n-1)d^2\ ns^2$ when $n = 4$

The element is in 4th **period** since $n = 4$ and also it is a ***d*-block element** as the configuration corresponds to *d*-block elements.

Complete electronic configuration is $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^2$ or [Ar] $4s^2\,3d^2$.

Total number of electrons = 22

The atomic number is 22.

It is Ti in Group 4.

(*iii*) $(n-2)f^7(n-1)d^1\,ns^2$ when $n = 6$

The configuration suggests that it is a **f-block element** with **4f series,** $(n-2)f$, and belongs to **6^{th} period.** The outer electronic configuration becomes $(6\text{–}2)f^7\,(6\text{–}1)d^1\,6s^2$, *i.e.*, $4f^7\,5d^1\,6s^2$. Its complete electronic configuration is [Xe] $4f^7\,5d^1\,6s^2$. Its atomic number is 54 + 7 + 1 + 2 = 64 and the element is gadolinium (Gd).

Q31. *The first ($\Delta_i H_1$) and the second ($\Delta_i H_2$) ionisation enthalpies (in kJ mol^{-1}) and the ($\Delta_{eg}H$) electron gain enthalpy (in kJ mol^{-1}) of a few elements are given below:*

Elements	ΔH_1	ΔH_2	$\Delta_{eg}H$
I	*520*	*7300*	*–60*
II	*419*	*3051*	*–48*
III	*1681*	*3374*	*–328*
IV	*1008*	*1846*	*–295*
V	*2372*	*5251*	*+48*
VI	*738*	*1451*	*–40*

Which of the above elements is likely to be:

(*a*) *the least reactive element.*

(*b*) *the most reactive metal.*

(*c*) *the most reactive non-metal.*

(*d*) *the least reactive non-metal.*

(*e*) *the metal which can form a stable binary halide of the formula MX_2 (X = halogen).*

(*f*) *the metal which can form a predominantly stable covalent halide of the formula MX (X = halogen)?*

Ans. (*a*) For the least reactive element, its first ionisation enthalpy ($\Delta_{ie}H_1$) must be very high and also electron gain enthalpy ($\Delta_{eg}H$). These conditions are met by **element number** V, therefore, element V is the least reactive. It is likely a Group 18 element.

(*b*) The most reactive metal must have very low $\Delta_{ie}H_1$ but high $\Delta_{ie}H_2$ and at the same time $\Delta_{eg}H$ should be less negative. Therefore, **metal number I** is the most reactive. The metal belongs to Group I of periodic table, *i.e.*, an alkali metal.

(*c*) For a non-metal to be most reactive, it should have high ionisation enthalpy and also its electron gain enthalpy should be highly negative. **Element III** satisfies these requirements and hence it is the most reactive non-metal (it is fluorine).

(*d*) Least reactive non-metals are undoubtedly Group 18 noble gases. These elements have very high ionisation enthalpy and very low (positive) electron gain enthalpy. **Element V** fits in this category.

(*e*) Metal that forms stable binary halide, MX_2, must form M^{2+} ions easily. It can happen when the sum of two ionisation enthalpies ($\Delta_{ie}H_1 + \Delta_{ie}H_2$) is relatively not very high and the electron gain enthalpy should also be less negative. In this regard, **element VI** is the answer and the metal belongs to Group 2 of the periodic table (alkaline earth metal).

(*f*) The metal that forms predominantly stable covalent halide of the type MX must belong to Group 1, *i.e.*, it should be an alkali metal for which the $\Delta_{ie}H_1$ should also be low. To form covalent halide, the metal should occupy the top most position of the group. The metal I (lithium) is the answer, also it has less negative $\Delta_{eg}H$ value.

Q32. *Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements:*

(*a*) *Lithium and oxygen*

(*b*) *Magnesium and nitrogen*

(*c*) *Aluminium and iodine*

(*d*) *Silicon and oxygen*

(*e*) *Phosphorus and fluorine*

(*f*) *Element 71 and fluorine.*

Ans. (*a*) Li_2O

(*b*) Mg_3N_2

(*c*) AlI_3

(*d*) SiO_2, silicon is in Group 14 with 4 valency electrons and the valency of oxygen is 2 in Group 16.

(e) Phosphorous is in Group 15 with a valency of 3 and also 5 while fluorine is in Group 17 with a valency of 1. The formulae would be PF_3 and PF_5.

(f) Element 71 is an inner transition 4f element, *i.e.*, belongs to lanthanoids. It is lutetium (Lu). Its common valency is 3 and with fluorine will form the compound LuF_3.

Q33. *In the modern periodic table, the period indicates the value of:*

(a) *atomic number*

(b) *atomic mass*

(c) *principal quantum number*

(d) *azimuthal quantum number.*

Ans. Correct answer is (c) since the principal quantum number designates a period.

Q34. *Which of the following statements related to the modern periodic table is incorrect?*

(a) *The p-block has 6 columns, because a maximum of 6 electrons can occupy all the orbitals in a p-shell.*

(b) *The d-block has 8 columns, because a maximum of 8 electrons can occupy all the orbitals in a d-subshell.*

(c) *Each block contains a number of columns equal to the number of electrons that can occupy that subshell.*

(d) *The block indicates value of azimuthal quantum number (l) for the last subshell that received electrons in building up the electronic configuration.*

Ans. Statement (b) is incorrect. The d-block must have 10 groups since the d-subshell can accommodate 10 electrons to the maximum.

Q35. *Anything that influences the valence electrons will affect the chemistry of the element. Which one of the following factors does not affect the valence shell?*

(a) *Valence principal quantum number (n)*

(b) *Nuclear charge (Z)*

(c) *Nuclear mass*

(d) *Number of core electrons.*

Ans. The answer is (c). Nuclear mass consists of neutrons and protons present in nucleus. Protons are responsible for the nuclear charge that can affect the valence electrons while neutrons do not affect the valence electrons.

Q36. *The size of isoelectronic species — F^-, Ne and Na^+ is affected by*

(*a*) *nuclear charge (Z)*

(*b*) *valence principal quantum number (n)*

(*c*) *electron-electron interaction in the outer orbitals*

(*d*) *none of the factors because their size is the same.*

Ans. The answer is (*a*) since the size of isoelectronic species given will vary as:

$$_9F^- > {}_{10}He > {}_{11}Na^+$$

In this series, size determining shell remains the same while nuclear charge is increasing in the order given.

Q37. *Which one of the following statements is incorrect in relation to ionisation enthalpy?*

(*a*) *Ionisation enthalpy increases for each successive electron.*

(*b*) *The greatest increase in ionisation enthalpy is experienced on removal of electron from core noble gas configuration.*

(*c*) *End of valence electrons is marked by a big jump in ionisation enthalpy.*

(*d*) *Removal of electron from orbitals bearing lower n value is easier than from orbital having higher n value.*

Ans. Statement (*d*) is incorrect. Electron with lower value of *n* will be held more strongly than the electron with higher *n* value.

Q38. *Considering the elements B, Al, Mg and K, the correct order of their metallic character is:*

(*a*) *B > Al > Mg > K* (*b*) *Al > Mg > B > K*

(*c*) *Mg > Al > K > B* (*d*) *K > Mg > Al > B*

Ans. The statement (*d*) is correct. The metallic character decreases across a period and increases on moving down a group.

Q39. *Considering the elements B, C, N, F and Si, the correct order of their non-metallic character is:*

(*a*) *B > C > Si > N > F* (*b*) *Si > C > B > N > F*

(*c*) *F > N > C > B > Si* (*d*) *F > N > C > Si > B*

Ans. The correct order of the non-metallic character is given in (*c*), *i.e.*, F > N > C > B > Si.

Non-metallic character increases across a period from left to right and decreases from top to bottom within a group.

Elements B, C, N and F are in the same period while Si is in the next period and below C in its group and hence the sequence.

Q40. *Considering the elements F, Cl, O and N, the correct order of their chemical reactivity in terms of oxidising property is:*

(*a*) $F > Cl > O > N$ (*b*) $F > O > Cl > N$

(*c*) $Cl > F > O > N$ (*d*) $O > F > N > Cl$

Ans. For the relative strength of oxidising and reducing agents, the best guide is standard electrode potentials of the half reactions. After checking the E°_{red} values, the correct order will be (*a*), *i.e.*, F > Cl > O > N.

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