Chapter
Classification of Elements and Periodicity in Properties

Topic-1: Periodic Classification

1 MCQs with One Correct Answer

1. Which of the following has the maximum number of unpaired electrons? **[1996 - 1 Mark]**
   (a) Mg^{2+} (b) Ti^{3+}
   (c) V^{3+} (d) Fe^{2+}
2. The statement that is not correct for the periodic classification of element is **[1992 - 1 Mark]**
   (a) The properties of elements are the periodic functions of their atomic numbers
   (b) Non-metallic elements are lesser in number than metallic elements
   (c) The first ionisation energies of elements along a period do not vary in a regular manner with increase in atomic number
   (d) For transition elements the *d*-subshells are filled with electrons monotonically with increase in atomic number.

6 MCQs with One or More than One Correct Answer

3 The statements that **are true** for the long form of the periodic table are : **[1988 - 1 Mark]**
   (a) it reflects the sequence of filling the electrons in the order of sub-energy level s, p, d and f.
   (b) it helps to predict the stable valency states of the elements
   (c) it reflects trends in physical and chemical properties of the elements
   (d) it helps to predict the relative ionicity of the bond between any two elements.

Topic-2: Periodic Properties

1 MCQs with One Correct Answer

1. Identify the correct order of acidic strengths of CO_2, CuO, CaO, H_2O **[2002S]**
   (a) $CaO < CuO < H_2O < CO_2$
   (b) $H_2O < CuO < CaO < CO_2$
   (c) $CaO < H_2O < CuO < CO_2$
   (d) $H_2O < CO_2 < CaO < CuO$
2. The correct order of radii is **[2000S]**
   (a) $N < Be < B$ (b) $F^- < O^{2-} < N^{3-}$
   (c) $Na < Li < K$ (d) $Fe^{3+} < Fe^{2+} < Fe^{4+}$
3. The correct order of acidic strength is **[2000S]**
   (a) $Cl_2O_7 > SO_2 > P_4O_{10}$ (b) $CO_2 > N_2O_5 > SO_3$
   (c) $Na_2O > MgO > Al_2O_3$ (d) $K_2O > CaO > MgO$
4. Amongst H_2O, H_2S, H_2Se and H_2Te, the one with the highest boiling point is **[2000S]**
   (a) H_2O because of hydrogen bonding
   (b) H_2Te because of higher molecular weight
   (c) H_2S because of hydrogen bonding
   (d) H_2Se because of lower molecular weight
5. Which has most stable +2 oxidation state : **[1995S]**
   (a) Sn (b) Pb (c) Fe (d) Ag
6. Amongst the following elements (whose electronic configurations are given below), the one having the highest ionization energy is : **[1990 - 1 Mark]**
   (a) $[Ne]\ 3s^2 3p^1$ (b) $[Ne]\ 3s^2 3p^3$
   (c) $[Ne]\ 3s^2 3p^2$ (d) $[Ne]\ 3d^{10} 4s^2 4p^3$
7. Which one of the following is the strongest base?
   (a) AsH_3 (b) NH_3 **[1989 - 1 Mark]**
   (c) PH_3 (d) SbH_3
8. Which one of the following is the smallest in size?
   (a) N^{3-} (b) O^{2-} **[1989 - 1 Mark]**
   (c) F^- (d) Na^+
9. The first ionisation potential of Na, Mg, Al and Si are in the order **[1988 - 1 Mark]**
   (a) $Na < Mg > Al < Si$ (b) $Na > Mg > Al > Si$
   (c) $Na < Mg < Al < Si$ (d) $Na > Mg > Al < Si$
10. The electronegativity of the following elements increases in the order **[1987 - 1 Mark]**
    (a) C, N, Si, P (b) N, Si, C, P
    (c) Si, P, C, N (d) P, Si, N, C
11. The first ionisation potential in electron volts of nitrogen and oxygen atoms are respectively given by **[1987 - 1 Mark]**
    (a) 14.6, 13.6 (b) 13.6, 14.6
    (c) 13.6, 13.6 (d) 14.6, 14.6
12. Atomic radii of fluorine and neon in Ångstorm units are respectively given by **[1987 - 1 Mark]**
    (a) 0.72, 1.60 (b) 1.60, 1.60
    (c) 0.72, 0.72 (d) None of these values

13. The element with the highest first ionization potential is **[1982 - 1 Mark]**
(a) boron (b) carbon
(c) nitrogen (d) oxygen

14. The correct order of second ionisation potential of carbon, nitrogen, oxygen and fluorine is **[1981 - 1 Mark]**
(a) C > N > O > F (b) O > N > F > C
(c) O > F > N > C (d) F > O > N > C

2 Integer Value Answer

15. The 1st, 2nd, and the 3rd ionization enthalpies, I_1, I_2, and I_3, of four atoms with atomic numbers n, $n+1$, $n+2$, and $n+3$, where $n < 10$, are tabulated below. What is the value of n? **[Adv. 2020]**

Atomic number	Ionization Enthalpy (kJ/mol)		
	I_1	I_2	I_3
n	1681	3374	6050
$n+1$	2081	3952	6122
$n+2$	496	4562	6910
$n+3$	738	1451	7733

16. Among the following, the number of elements showing only one non-zero oxidation state is :
O, Cl, F, N, P, Sn, Tl, Na, Ti **[2010]**

4 Fill in the Blanks

17. On Mulliken scale, the average of ionization potential and electron affinity is known as ..... **[1985 - 1 Mark]**

18. The energy released when an electron is added to a neutral gaseous atom is called ..... of the atom. **[1982 - 1 Mark]**

5 True / False

19. The basic nature of the hydroxides of group 13 (Gr. III B) decreases progressively down the group. **[1993 - 1 Mark]**

20. The decreasing order of electron affinity of F, Cl, Br is F > Cl > Br. **[1993 - 1 Mark]**

21. In group IA, of alkali metals, the ionisation potential decreases on moving down the group. Therefore, lithium is a poor reducing agent. **[1987 - 1 Mark]**

6 MCQs with One or More than One Correct Answer

22. The option(s) with only amphoteric oxides is(are) **[Adv. 2017]**
(a) Cr_2O_3, BeO, SnO, SnO_2 (b) Cr_2O_3, CrO, SnO, PbO
(c) NO, B_2O_3, PbO, SnO_2 (d) ZnO, Al_2O_3, PbO, PbO_2

23. Ionic radii of **[1999 - 3 Marks]**
(a) $Ti^{4+} < Mn^{7+}$ (b) $^{35}Cl^- < {}^{37}Cl^-$
(c) $K^+ > Cl^-$ (d) $P^{3+} > P^{5+}$

24. Sodium sulphate is soluble in water whereas barium sulphate is sparingly soluble because : **[1989 - 1 Mark]**
(a) the hydration energy of sodium sulphate is more than its lattice energy
(b) the lattice energy of barium sulphate is more than its hydration energy
(c) the lattice energy has no role to play in solubility
(d) the hydration energy of sodium sulphate is less than its lattice energy.

9 Assertion and Reason Statement Type Questions

Each question contains **STATEMENT-1 (Assertion)** and **STATEMENT-2 (Reason)**. Each question has 4 choices (a), (b), (c) and (d) out of which ONLY ONE is correct. Mark your answer as
(a) If both Statement -1 and Statement -2 are correct, and Statement -2 is the correct explanation of the Statement -2.
(b) If both Statement -1 and Statement -2 are correct, but Statement -2 is not the correct explanation of the Statement -1.
(c) If Statement -1 is correct but Statement -2 is incorrect.
(d) If Statement -1 is incorrect but Statement -2 is correct.

25. Read the following statement and explanation and answer as per the options given below :
Assertion : The first ionization energy of Be is greater than that of B. **[2000S]**
Reason : $2p$ orbital is lower in energy than $2s$.

10 Subjective Problems

26. Arrange the following in :
(*i*) Arrange the following ions in order of their increasing radii : Li^+, Mg^{2+}, K^+, Al^{3+}. **[1997 - 1 Mark]**
(*ii*) Increasing order of basic character : MgO, SrO, K_2O, NiO, Cs_2O **[1991 - 1 Mark]**
(*iii*) Increasing order of ionic size : N^{3-}, Na^+, F^-, O^{2-}, Mg^{2+} **[1991 - 1 Mark]**
(*iv*) Increasing size : Cl^-, S^{2-}, Ca^{2+}, Ar **[1986 - 1 Mark]**
(*v*) Increasing first ionization potential : Mg, Al, Si, Na **[1985 - 1 Mark]**
(*vi*) Increasing acidic property : ZnO, Na_2O_2, P_2O_5, MgO **[1985 - 1 Mark]**
(*vii*) Decreasing ionic size : Mg^{2+}, O^{2-}, Na^+, F^- **[1985 - 1 Mark]**

27. The first ionization energy of carbon atom is greater than that of boron atom whereas, the reverse is true for the second ionization energy. **[1989 - 2 Marks]**

Answer Key

Topic-1 : Periodic Classification

1. (d) 2. (d) 3. (a, c, d)

Topic-2 : Periodic Properties

1. (a) 2. (b) 3. (a) 4. (a) 5. (b) 6. (b) 7. (b) 8. (d) 9. (a) 10. (c)
11. (a) 12. (a) 13. (c) 14. (c) 15. (9) 16. (2) 17. (Electronegativity)
18. (Electron affinity) 19. (False) 20. (False) 21. (False) 22. (a, d) 23. (d) 24. (a, b)
25. (c)

Hints & Solutions

Topic-1: Periodic Classification

1. **(d)** The electronic configuration of the given ions are as follows.

$_{12}Mg^{2+} = 1s^2$, (No unpaired electron)

$_{22}Ti^{3+} = 1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^1$ (One unpaired electron)

$_{23}V^{3+} = 1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^2$ (Two unpaired electrons)

$_{26}Fe^{2+} = 1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^6$ (Four unpaired electrons)

2. **(d)** The electrons are not filled in *d*-subshell monotonically with increase in atomic number, among transition elements. It breaks at chromium and copper.

3. **(a, c, d)** Periodic table does not help to predict the stable valency states of the elements.

Topic-2: Periodic Properties

1. **(a)** Non-metallic oxides are acidic and metallic oxides are basic. Thus, the acidic order is

$CaO < CuO < H_2O < CO_2$.

2. **(b)** Effective nuclear charge (*i.e.* Z/e ratio) decreases from F^- to N^{3-}. Hence the radii follows the order:

$F^- < O^{2-} < N^{3-}$. Z/e for $F^- = 9/10 = 0.9$, for $O^{2-} = 8/10 = .8$, for $N^{3-} = 7/10 = 0.7$

3. **(a)** Non-metallic oxides are acidic and acidic character decreases with increase in metallic character.

4. **(a)**

(i) Hydrogen bonding increases the boiling point.

(ii) Hydrogen bonds are formed in compounds having F or O or N with hydrogen.

S, Se, Te cannot undergo hydrogen bond formation because of their larger size and lower electronegativity values.

5. **(b)**

(i) Ion having half filled or full filled orbital have extra stability.

(ii) Larger the size of cation more will be its stability Pb^{2+} ($5d^{10}$ $6s^2$), has the most stable +2 oxidation state because here the *d*-orbital is completely filled and is more stable than Fe^{2+} ($3d^6$). Again Ag^+ ($4d^{10}$) is more stable as here again the *d*-orbital is completely filled and Ag^{2+} is not easily obtained. Pb^{2+} is more stable compared to Sn^{2+} ($4d^{10}$ $5s^2$) because of its large size.

6. **(b)** Ionisation energy increases with increasing atomic number in a period, while it decreases on moving down a group. IE of element with electronic configuration (d) is lowest because of its biggest size. Among the remaining three elements of the same period (3rd), IE of element with electronic configuration (b) is the highest due to greater stability of the exactly half-filled 3 *p*-subshell.

7. **(b)** Nitrogen, being smallest in size, can give up its lone pair of electrons most easily.

8. **(d)**

For isoelectronic ions, ionic size $\propto \frac{1}{\text{atomic number}}$

Species	No. of e^-	At. No.
N^{-3}	10	7
O^{-2}	10	8
F^-	10	9
Na^+	10	11

$\therefore$ Na^+ is smallest in size.

9. **(a)** First ionisation potential increases from left to right in a period.

IE_1 of Mg is higher than that of Na because of increased nuclear charge and also than that of Al because in Mg, a 3 *s*-electron has to be removed while in Al, it is the 3 *p*-electron. The IE_1 of Si is, however, higher than those of Mg and Al because of its increased nuclear charge. Thus, the overall order is : Na < Mg > Al < Si.

10. **(c)** Electronegativity increases on moving from left to right in a period and decreases on moving from top to bottom in a group.

Si and P are placed in the 3rd period while C and N are placed in the 2nd period. Elements in 2nd period have higher electronegativities than those in the 3rd period. Since, N has smaller size and higher nuclear charge than C, its electronegativity is higher than that of C. Similarly, the electronegativity of P is higher than that of Si. Thus, the overall order is : Si, P, C, N.

11. **(a)** Ionisation potential of nitrogen is more than that of oxygen. This is because nitrogen has more stable **half-filled** *p*-orbitals. ($N = 1s^2, 2s^2, 2p^3$, $O = 1s^2, 2s^2, 2p^4$)

12. **(a)**

(i) Noble gases do not have covalent radii. They have only van der Waals radii.

(ii) Covalent radii is always smaller than corresponding van der Waals radii.

13. **(c)** Amongst B, C, N and O; N has the highest first ionization energy, because of its half filled 2*p* orbital which is more stable.

14. **(c)**

$C - 1s^2\,2s^2\,p^2$	$C^+ - 1s^2\,2s^2\,p^1$
$N - 1s^2\,2s^2\,p^3$	$N^+ - 1s^2\,2s^2\,p^2$
$O - 1s^2\,2s^2\,p^4$	$O^+ - 1s^2\,2s^2\,p^3$
$F - 1s^2\,2s^2\,p^5$	$F^+ - 1s^2\,2s^2\,p^4$

As for IE_2, the electron in all the cases is to be removed from $2p$ orbital, so it must follow the order :

$C < N < O < F$ (*i.e.* increase across a period)

But in case of O^+, the $2p$ orbital is half-filled and is more stable as compared to others. So, the order becomes : $O > F > N > C$

15. **(9)** By observing the values of ionization enthalpy for atomic number $(n+2)$, it is observed that $I_2 >> I_1$. Which shows that number of valence shell electrons is 1 for atomic number $(n+2)$. Therefore element with atomic number $(n+2)$ should be an alkali metal.

For atomic number $(n+3)$, $I_3 >> I_2$, which shows that it will be an alkaline earth metal. All the observations suggests that atomic number $(n+1)$ should be a noble gas and atomic number (n) should belong to halogen family. Since $n < 10$; hence $n = 9$.

16. **(2)** Fluorine generally shows 0 and –1 oxidation states while sodium shows 0 and +1 oxidation state.

17. **Electronegativity**

18. **Electron affinity**

19. **False :** On moving down the group 13 (III) A the basic nature of hydroxides increases. The basic nature increases as the element becomes more electropositive or acquires more metallic character when we move down a group.

20. **False :** Halogens have high electron affinities which decrease as we move down the group. However, fluorine has lower value of E.A. than chlorine which is due to its small size and more repulsion between the electron added and electrons already present. Hence, the order is : $Cl > Fe > Br$.

21. **False :** *Ionisation energy decreases on moving down in group IA from* Li *to* Cs, the reducing property should **increase** in the same order, *i.e.,* from Li to Cs which is found to be so, except an *anomaly in lithium which is found to be the strongest reducing agent*; because of its **higher oxidation potential** (E°).

22. **(a, d)** $NO \Rightarrow$ Neutral

$B_2O_3 \Rightarrow$ Acidic

$CrO \Rightarrow$ Basic

All other oxides are amphoteric.

23. **(d)** Higher the (+) charge, smaller will be radii.

24. **(a, b)**

For dissolution, Hydration energy > Lattice energy.

$BaSO_4$ is sparingly soluble in water because its hydration energy is lesser than the lattice energy and thus, ions are not separated from each other. On the contrary in Na_2SO_4, the hydration energy is more than its lattice energy. Thus, ions are separated from each other and pass in solution state.

25. **(c)** The first ionisation energy of Be $(1s^2 2s^2)$ is greater than that of B $(1s^2 2s^2 p^1)$ because removal of electron from $2s$ orbital is tough compare to that from $2p$ orbital. This is because $2s$-orbital is closer to nucleus than the $2p$-orbital. Reason is incorrect as $2p$-orbital is higher in energy.

26. (*i*) $Al^{3+} < Mg^{2+} < Li^+ < K^+$

Al^{3+} and Mg^{2+} are isoelectronic species, so in these, size decreases with increase in atomic number because increase in atomic number increases Z_{eff}.

Due to anomalous behaviour and diagonal relationship, the ionic radius of Li^+ is slightly bigger than Mg^{2+}.

(*ii*) Increasing order of basic character :

$NiO < MgO < SrO < K_2O < Cs_2O$

The basic character of oxides increases when we move down the group. So, $K_2O < Cs_2O$ and $MgO < SrO$.

Further, higher the group number lesser is the basic character. Hence, NiO is the least basic.

(*iii*) Increasing order of ionic size :

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

All the above ions are isoelectronic having 10 electrons each.

In such a case, the greater the nuclear charge, the greater is the attraction for electrons and smaller is the ionic radius. Hence, N^{3-} has the highest and Mg^{2+} has the least ionic size.

(*iv*) $Ca^{2+} < Ar < Cl^- < S^{2-}$; All of these are isoelectronic. In such cases, the greater the nuclear charge, the greater is the attraction for electrons and smaller is ionic size.

$$\text{ionic radius} \propto \frac{1}{\text{effective nuclear charge}}$$

(*v*) The first ionization potential of the 3rd period elements follows the order :

$Na < Al < Mg < Si$

Ionisation energy **increases across a period** but not regularly. Mg $(1s^2, 2s^2p^6, 3s^2)$ is more stable because the electron is to be removed from $3s$ which is difficult as compared to Al $(1s^2, 2s^2p^6, 3s^2p^1)$ where electron is to be removed from $3p$.

(*vi*) $\overset{+1}{Na_2}O_2 < \overset{+2}{Mg}O < \overset{+2}{Zn}O < \overset{+5}{P_2}O_5$

Among oxides, the acidic strength increases with oxidation state. So Na_2O_2 is least acidic and P_2O_5 is most acidic. Further Na_2O_2 and MgO are basic, ZnO is amphoteric and P_2O_5 is acidic.

(*vii*) $O^{2-} > F^- > Na^+ > Mg^{2+}$

All the above ions are isoelectronic having 10 electron each.

In such case, the greater the nuclear charge, the greater is the attraction for electrons and smaller is the ionic radius. Hence, O^{2-} has the highest and Mg^{2+} has the least ionic size.

27. C $(1s^2 2s^2 p^2)$ has high nuclear charge than B $(1s^2 2s^2 p^1)$. Thus, carbon has higher IE value than Boron.

Further, for second ionization energy (IE_2) in C^+ $(1s^2 2s^2 p^1)$ the electron is to be removed from $2p$ which is easy as compared to B^+ $(1s^2 2s^2)$, where it has to be removed from $2s$.