Chapter 3

Classification of Elements and Periodicity in Properties

Solutions (Set-1)

Very Short Answer Type Questions :

1. In the terms of electronic configuration, why the halogens have similar chemical properties?

Sol. Halogens show similar chemical properties because of similar outermost electronic configuration.

2. What is the nature of the oxides formed by the *p*-block elements?

Sol. *p*-block elements are mainly non-metals. So their oxides are acidic in nature.

3. What is meant by periodic classification of elements?

Sol. Periodic classification of elements means to classify the known elements in accordance of their properties in such a manner that their study of properties become easier.

4. Explain the meaning of a positive electron gain enthalpy.

Sol. Positive electron gain enthalpy means that energy is required rather released during addition of an electron to a neutral gaseous atom to form gaseous ion.

5. What are isoelectronic species?

Sol. Isoelectronic species are those chemical species which contain the same number of electrons in them *e.g.*, N^{3-}, O^{2-}, F^{-}, Na^{+}, Mg^{2+}, Al^{3+}.

6. What is the nature of oxides formed by the 1^{st} group of elements and 17^{th} group of elements?

Sol. 1^{st} group of elements forms basic oxides while the 17^{th} group of elements forms acidic oxides.

7. Out of Na and Na^{+}, which one will have smaller size?

Sol. Na^{+} will have smaller size due to higher effective nuclear charge.

8. Which is smaller in size – Li or F and why?

Sol. F is smaller in size than Li because the effective nuclear charge increases moving from left to right across the period, consequently the size of the atom decreases.

9. Out of Na and Mg, which has a higher second ionization enthalpy and why?

Sol. Na has a higher second ionization enthalpy because the electron is to be removed from Na^{+} which has a stable noble gas configuration of Ne[$1s^22s^22p^6$] whereas Mg attains the configuration like that of Na[$1s^22s^22p^63s^1$].

10. Give examples of three cations and three anions which are isoelectronic with Argon (Ar).

Sol. Cations: K^{+}, Ca^{2+}, Sc^{3+}

Anions: Cl^{-}, S^{2-}, P^{3-}

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Short Answer Type Questions :

11. What is Dobereiner triad?

Sol. Dobereiner tried to arrange similar elements in groups of three known as triad, in which the atomic weight of the middle element was approximately the arithmetic mean of the other two. For example

Li – 7

Na – 23

K – 39

The mean of 1st and 3rd *i.e.*, Li and K

$$= \frac{7+39}{2} = \frac{46}{2} = 23$$

12. What are general characteristics of *s*-block elements?

Sol. (i) They are generally soft with low melting point and boiling point.

(ii) They have low ionization enthalpy and high electropositivity.

(iii) They act as strong reducing agent.

(iv) They are highly reactive and reactivity increases down the group

13. Moving down the alkali metal group which element is expected to be least electronegative and why?

Sol. Electronegativity decreases as the size of atom increases. On moving down the group, francium is expected to have the largest size therefore it will be the least electronegative element.

14. Why zero group elements are inert?

Sol. Zero group elements are inert because they have completely filled valence shells. Neither they have the tendency to lose electrons nor to gain or share electrons with other elements.

15. Why argon (atomic mass 39.94) has been placed before potassium (atomic mass 39.10) in the periodic table?

Sol. In the modern periodic table, elements have been placed in order of their increasing atomic numbers, not on the basis of atomic masses. Since the atomic number of argon is 18 and that of potassium is 19, argon is placed before potassium.

16. Why the second ionization enthalpy of an element is higher than the first?

Sol. Because after the removal of one electron from the atom, the number of electrons decreases whereas the nuclear charge remains the same. The cation so formed has a higher effective nuclear charge, as a result the force of attraction by the nucleus on the electrons increases thus increasing the second ionization enthalpy.

17. First ionization enthalpy of nitrogen is higher than the first ionization enthalpy of oxygen. Why?

Sol. Nitrogen has a stable configuration where the electrons in the degenerate *p*-orbitals have equally distributed electrons $(1s^2, 2s^2, 2p_x^1\ 2p_y^1\ 2p_z^1)$ whereas in oxygen the electrons in the degenerate *p*-orbitals have unsymmetrical distributions in which one orbital contains two electrons. The presence of two electrons in one orbital $(1s^2, 2s^2, 2p_x^2\ 2p_y^1\ 2p_z^1)$ increases the interelectronic repulsion thus decreasing the ionization enthalpy.

18. Why sodium is a strong electropositive element?

Sol. The ionization enthalpy of Na is low as it contains only one electron in its valence shell. Hence, it can easily lose one electron to attain a noble gas configuration. Therefore it is a strong electropositive element.

19. Why there are only 14 lanthanides and only 14 actinides?

Sol. In lanthanides and actinides the differentiating electrons enters to (n – 2)f subshell. The maximum number of electrons that can be accommodated in the seven orbitals is 14 electrons. Thus there are only 14 lanthanides $(4f^{1-14})$ an**d only 14 actinides $(5f^{1-14})$.**

20. What are the reasons for the different chemical behaviour of the first member of a group of element in the *s* and *p*-blocks compared to that of subsequent members in the same group?

Sol. The first member of each group in the *s* and *p*-blocks have small size, large charge/radius ratio and high electronegativity which accounts for their anomalous behaviour in comparison to their subsequent members in the same group.

21. Out of metallic radius and covalent radius of an element, which is larger and why?

Sol. The metallic radius of an atom is always larger than the covalent radius of the element because metallic bond is a weak bond in comparison to covalent bond. The atoms held by covalent bond are closer to each other.

22. The ionization energy of lithium is 520 kJ/mol. Calculate the amount of energy required to convert 700 mg of lithium atoms in gaseous state int**o Li^+ ions.**

Sol. Mass of lithium = 700 mg = 700×10^{-3}g = 7×10^{-1} g

Mole of lithium = $\frac{7\times10^{-1}}{7} = 1\times10^{-1}$ mol

Energy required to convert 1×10^{-1} mol atoms of lithium in gaseous state into Li^+ ions

= $520 \times 1 \times 10^{-1}$

= 52 kJ

23. What do you understand by 'Representative elements'? Name the groups whose elements are called representative elements.

Sol. The elements of *s*-block and *p*-block are called representative elements as they represent the properties of their respective groups.

The groups called representative element are 1, 2, 13, 14, 15, 16, 17 and 18.

24. What is the general electronic configuration of *d*-block elements? Why are these known as "Transition elements"?

Sol. The general configuration of *d*-block elements is $(n-1)^{1-10}\ ns^{0-2}$. These are known as transition elements because these form a bridge between the chemically active metals of *s*-block and the less active elements of group 13 and 14.

25. Among the following compounds CsI, CsF, LiF and NaF, which one has the highest cation to anion size ratio?

Sol. Amongst alkali metals cation *i.e.*, Cs, Na, and Li, the largest size of cation is of Cs^+. Among the halides *i.e.*, F and I, smallest size is of F^-. Therefore the ratio of Cs^+ and F^- is the highest.

26. The first and the second ionization enthalpies (kJ/mol) of three elements I, II and III are given below.

Element	$\Delta_i H_1$	$\Delta_i H_2$
I	403	2640
II	549	1060
III	1142	2080

Identify the element which is likely to be

(i) A non-metal

(ii) An alkali metal

(iii) An alkaline earth metal

Sol. (i) Element (III) is a non-metal because it has high first ionization enthalpy in comparison to other elements showing its tendency to gain electron rather than loosing it.

(ii) Element (I) is an alkali metal which has the electronic configuration ns^1. These elements show high tendency to lose electrons to attain a noble gas configuration *i.e.*, they have low ionization enthalpies. Element (I) has the lowest first inonization enthalpy whereas the second ionization enthalpy is very high.

(iii) Element (II) is an alkaline earth metal because its ionization enthalpies are quite low which show the ease of losing electrons.

27. Why He and Be both having similar electronic configuration, Be is not inert?

Sol. Although He ($1s^2$) and Be($1s^2 2s^2$) both having similar outer electronic configuration, *i.e.* ns^2 only. He has the inert gas configuration and hence, chemically inert but Be does not have inert gas configuration and hence, not chemically inert.

28. Ionization enthalpy of Be is greater than that of both Li and B. Explain.

Sol. The electronic configuration of the given elements are :

$_3$Li – $1s^2\,2s^1$

$_4$Be – $1s^2\,2s^2$

$_5$B – $1s^2\,2s^2 2p^1$

The ionization enthalpy of Be is greater than Li because it has completely filled $2s^2$ configuration which imparts greater stability than incompletely filled $2s^1$ configuration in Li. Also Be has a higher nuclear charge.

In case of B, electron is removed from $2p^1$ orbital which is located far away from the nucleus and screened by the 2*s* orbital electrons. So the removal of electron becomes easy. Consequently the ionization enthalpy lowers.

29. Why the zero group (18th group) has been placed at the extreme right of the periodic table?

Sol. This is in accordance with electronic configurations because each period starts with the filling of *s*-subshell and is completed with the filling of *p*-subshell of the same principal shell.

30. Why lanthanoids and actinoids are placed in separate rows at the bottom of the periodic table?

Sol. These are separately placed at the bottom of the periodic table for convenience. If they are placed within the body of the periodic table, the periodic table will become very long and cumbersome.

Long Answer Type Questions :

31. Describe the main features of the long form of the periodic table. In what respects is it superior to Mandeleev's table?

Sol. The main features of the present or long form of the periodic table which makes it superior to Mendeleev's periodic table are:

(i) The positioning of elements in the periodic table is done on the basis of atomic numbers whereas in Mendeleev's table it is on the basis of atomic mass.

(ii) It is easy to remember and reproduce all the elements more easily in sequence of atomic numbers due to differentiating the elements in different *s*, *p*, *d* and *f*-blocks based on the outer electronic configuration, while Mendeleev's table lacked these things.

(iii) The elements are classified into normal elements belonging to 1, 2, 13-17 groups and transition elements belonging to 3-12 groups on the basis of their outer electronic configuration. The elements are also classified as active metals placed in groups 1 and 2, heavy metals placed in groups 3-12 and non-metals placed in the upper right corner (groups 13 -18) of this periodic table, whereas all these clear classifications were lacking in Mendeleev's table.

32. Among the elements B, Al, C and Si

(a) Which has the highest first ionization enthalpy?

(b) Which has the most negative electron gain enthalpy?

(c) Which has the largest atomic radius?

(d) Which has the most metallic character?

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Sol. If we arrange the elements in periods and groups it will become easy to identify the characteristics.

Group →	XIII	XIV
Period 2	B	C
Period 3	Al	Si

(a) C has the highest ionization enthalpy because as we move across the period the effective nuclear charge increases which makes difficult to remove the electron. While moving down the ionization enthalpy decrease due to increase in the distance from the nucleus of the valence electron.

(b) C will have the most negative electron gain enthalpy because more is the effective nuclear charge, smaller is the size of atom and greater is the tendency to gain electron.

(c) Largest atomic radius is of Al because as we move down the group there is an increase in a shell which makes the size larger of the atom and moving from left to right, the size of the atom decreases. So Al will have a greater size than Si.

(d) As we go from left to right the ionization enthalpy increases *i.e.*, metallic character decreases. So in the second period *B* is more metallic and Al in the 3rd period is more metallic. When we move from top to bottom in a group the metallic character increases. Hence, Al is the most metallic element in the given four elements.

33. Give the general valence shell electronic configuration of transition elements and give some characteristics of these. Also tell the block to which they belong and why?

Sol. The general valence shell configuration of transition elements is $(n-1)d^{1-10}\ ns^{0-2}$. The transition elements belong to *d*-block because the last electron enters the penultimate energy level.

Some characteristics of *d*-block elements are :

(i) They all are metals.

(ii) They show variable oxidation states.

(iii) They form coloured complexes.

(iv) Most of them possesses catalytic properties.

(v) They form alloys.

(vi) They are good conductor of heat and electricity.

34. Arrange the isoelectronic species O^{2-}, F^- , Na^+ and Mg^{2+} in order of their

(a) Increasing effective nuclear charge

(b) Increasing ionic radius

(c) Increasing ionization enthalpy

Give reasons for their arrangement also.

Sol. (a) As the positive charge increases on the atom. The effective nuclear charge increases, thus increasing the force of attraction of the nucleus on the valence electrons. The order of species becomes

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

(b) As the effective nuclear charge increases, the size of the ion decreases accordingly due to the increased force of attraction by the nucleus on the valence electrons. So the size of the ions can be arranged accordingly in increasing order as follows:

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

(c) The increased effective nuclear charge increases the hold of nucleus on electrons means higher amount of energy will be required to remove electron from the valence shell. So the order of ionization enthalpy becomes as

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

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

Sol.

Metals	**Non-metals**
(i) They have a strong tendency to lose electrons to form cation.	They have a strong tendency to gain electrons to form anions.
(ii) Due to the tendency to lose electrons they act as strong reducing agents	Due to their tendency to gain electrons, they act as strong oxidizing agents.
(iii) They have low ionization enthalpies due to the ease of removal of electrons.	They have high ionization enthalpies due to their greater tendency to accept electron rather donation.
(iv) They have low electronegativity	They have high electronegativity.
(v) They have low negative electron gain enthalpies.	They have high negative electron gain enthalpies.
(vi) They form basic oxides.	They form acidic oxides.
(vii) Except Li and Be they predominantly form ionic compounds.	They form covalent compounds.

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

Sol. The elements of group 1 have only one electron in its valence shell and thus have a strong tendency to lose this electron. By the loss of this one electron, the element can acquire the stable noble gas configuration. The tendency to lose electron in turn depends upon the ionization enthalpy. Since the ionization enthalpy decreases down the group, therefore, the reactivity of the elements of group 1 increases in the order : Li < Na < K < Rb < Cs.

In contrast, the elements of the 17^{th} group have seven electrons in the valence shell and requires one electron to complete its octet and acquire a stable noble gas configuration. So these elements of this group show a high tendency to accept one electron. This tendency to gain electrons can be linked to the electron gain enthalpy. Since electron gain enthalpy becomes less and less negative as we move down the group from Cl to I, therefore, reactivity increases from Cl to I. Fluorine however have the highest reactivity due to low bond dissociation energy (amount of energy required to break the bond).

37. Explain the following

(a) Electronegativity of elements increases on moving from left to right in the periodic table.

(b) Ionization enthalpy decreases in a group from top to bottom.

Sol. (a) As we move across the period from left to right, the effective nuclear charge increases and the atomic radius decreases. As a result, the tendency of the atom of an element to attract the shared pair of electrons towards itself increases and hence, the electronegativity of the element increases.

(b) On moving down the group from top to bottom, the atomic size increases gradually due to the addition of new principal energy shell at each succeeding element. As a result the distance of the valence electrons from the nucleus increases.

Consequently the force of attraction of the nucleus for the valence electrons decrease and the ionization enthalpy should decrease with the addition of a new shell. The shielding effect or the screening effect increases, thus decreasing the force of attraction of the nucleus on the valence electrons, decreasing the ionization enthalpy.

38. The formation of Cl^- from Cl is exothermic but formation of O^{2-} from O^- is endothermic. Explain

Sol. The formation of Cl^- from Cl is an exothermic reaction because the electron gain enthalpy is negative. Cl having electronic configuration $1s^22s^22p^63s^23p^5$, requires one electron to acquire a stable noble gas configuration like that of argon.

$$Cl(g) \xrightarrow{+1e^-} Cl^-(g);\ \Delta_{eg}H_1 = -ve$$

The formation of O^{2-} involves two processes

$$O^-(g) \xrightarrow{+1e^-} O^-(g);\ \Delta_{eg}H_1 = -ve$$

$$O^-(g) \xrightarrow{+1e^-} O^{2-}(g);\ \Delta_{eg}H_2 = +ve$$

The cause for the negative electron gain enthalpy in the first step forming O^-(g) ion from O(g) is the tendency to gain electron due to high effective nuclear charge and small size. But after gaining one electron though the nuclear charge remains the same, the number of electrons increases in O^-. The effective nuclear charge decreases. When the second electron is added to O^- it requires energy *i.e.*, electron gain enthalpy becomes positive because the incoming second electron starts facing strong repulsion from the inner electrons thus dominating the nuclear force of attraction on the valence electron.

39. (i) Why Na cannot exhibit +2 oxidation state?

(ii) Why do noble gases have the largest size in their respective periods?

Sol. (i) Na cannot exhibit +2 oxidation state because if we look at the electronic configuration $1s^22s^22p^63s^1$, it contains only one electron in its valence shell. It is easy to remove one electron from its valence shell *i.e.*, it has low ionization enthalpy. But after loosing one electron it acquires a stable noble gas configuration with the complete octet. The second ionization enthalpy drastically increases making it unfeasible to lose the second electron. So Na can exhibit only +1 oxidation state and not +2 oxidation state.

(ii) Noble gases have the largest atomic size in their respective periods because noble gases being inert do not combine with other elements. So they have a weak interparticle force of attraction known as van der Waals forces and the bond formed is known as van der Waals bond.

40. Predict the formula of the stable binary compounds that would be formed by the following pairs of elements

(i) Silicon and oxygen

(ii) Aluminium and bromine

(iii) Calcium and iodine

(iv) Element 114 and fluorine

(v) Element 120 and oxygen

Sol. (i) Silicon belongs to 14th group. Its valence is 4. The valence of oxygen is 2. Thus, the binary compound between silicon and oxygen would have formula SiO_2.

(ii) Aluminium belongs to 13th group. Its valence is 3 the valence of bromine is 1. Thus, the binary compound between aluminium and bromine would have formula $AlBr_3$.

(iii) Calcium belongs to 2nd group. Its valence is 2. The valence of iodine is 1. Thus, the compound between calcium and iodine would have formula CaI_2.

(iv) The element (M) with atomic number 114 belongs to 14th group. It exhibits a valence of 2 (due to inert pair of effect) most prominantly than 4. So the compound formed would have the formula MF_2.

(v) The element (X) with atomic number 120 belongs to group 2. Its valence would be 2. The valence of oxygen is 2. Thus, the formula of the compound between this element and oxygen would have formula XO.

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41. (i) Show by a chemical reaction that MgO is a basic oxide while P_4O_{10} is an acidic oxide.

(ii) Which period elements have the ability to form $p\pi$-$p\pi$ multiple bonds? Give reason for it.

Sol. (i) MgO reacts with aqueous HCl to form salt and water thus exhibiting its basic characters

$$MgO(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2O(l)$$

P_4O_{10} reacts with water to form phosphoric acid which turns the blue litmus red.

$$P_4O_{10}(s) + 6H_2O(l) \longrightarrow \underset{\text{Phosphoric acid}}{4H_3PO_4(aq)}$$

(ii) The first member of each group of p-block elements has a great tendency to form $p\pi$-$p\pi$ multiple bonds to itself and to the other second period elements. For example, carbon forms bonds of the type, C = C, C ≡ C, C = O, C ≡ N etc. This property of these elements is due to their small size. The higher members of the group have little tendency to form $p\pi$-$p\pi$ bonds.

42. (i) What is the basic difference between the terms electron gain enthalpy and electronegativity?

(ii) What do you understand by shielding effect or screening effect?

Sol. (i) Electron gain enthalpy is the property of an isolated atom of an element in gaseous state whereas electronegativity is the property of an atom of the element when it is bonded to some other atom. The value of electron gain enthalpy can be experimentally determined while that of electronegativity cannot be determined experimentally.

(ii) In multi-electron atoms, the electrons present in the outermost shell do not experience the complete nuclear charge because of the repulsive interaction of the intervening electrons. Thus, the outermost electrons are shielded or screened from the nucleus by the inner electrons. This is known as screening effect.

If the number of electrons in the inner shells is large, the screening effect will be large.

43. Select from each group the species which has the smallest radius stating appropriate reason.

(i) O, O^-, O^{2-}

(ii) K^+, Sr^{2+}, Ar

(iii) Si, P, Cl

Sol. (i) O has the smallest radius because the size of anion is always larger than the parent atom.

Here O^- and O^{2-} are anions of O.

(ii) K^+ has the smallest radius. In K^+ and Ar, the outermost shell is third whereas in Sr^{2+} it is fourth. Out of K^+ and Ar, K^+ has smaller size because it has greater nuclear charge.

(iii) Cl has the smallest radius because Si, P, Cl belong to the same period and as we move across the period the effective nuclear charge increases hence, the size decreases.

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Chapter 3

Classification of Elements and Periodicity in Properties

Solutions (Set-2)

[Periodic Table and Classification]

1. In Lother Meyer's curve, the element on the peak of curve will be

(1) F (2) Na (3) Mg (4) Ne

Sol. Answer (2)

Due to large atomic volume.

2. According to Mendeleev's periodic law, physical and chemical properties are function of

(1) Atomic number (2) Atomic weight (3) Atomic volume (4) Number of neutrons

Sol. Answer (2)

Classification is based on Mendeleev's periodic law.

3. Uub is the symbol for the element with atomic number

(1) 102 (2) 108 (3) 110 (4) 112

Sol. Answer (4)

Uub

U = un = 1

u = un = 1

b = bium = 2

Hence element has atomic number 112.

4. With which block $_{30}Zn$ belongs?

(1) *s* (2) *p* (3) *d* (4) *f*

Sol. Answer (3)

Last electron enters in 'd' sub-shell.

Zn belongs with 12th group.

5. The alkali metal which is radioactive is

(1) Fr (2) Al (3) Li (4) Mg

Sol. Answer (1)

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6. The element which belongs with chalcogen family is

(1) N (2) P (3) S (4) Cl

Sol. Answer (3)

16^{th} group is chalcogen.

7. If the atomic number of an element is 33, it will be placed in the periodic table in the

(1) 1st group (2) 3rd group (3) 15th group (4) 17th group

Sol. Answer (3)

$33 = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3 = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^3$

Above is a p-block element.

group number = 10 + number of valence electrons

= 10 + 5 = 15

8. An element has electronic configuration [Xe] $4f^7$, $5d^1$, $6s^2$. It belongs to ______ block of the periodic table.

(1) *s* (2) *p* (3) *d* (4) *f*

Sol. Answer (4)

[Xe] $4f^7\ 5d^1\ 6s^2$

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 5d^1\ 4f^7$

Last electron enters the *f* sub shell, hence, it belongs to *f* block.

9. The statement that is not correct for the periodic classification of element is

(1) The properties of elements are the periodic function of their atomic numbers

(2) Non-metallic elements are less in number than metallic elements

(3) The first ionisation energy of elements along a period decrease with increase in atomic number

(4) For transition elements the *d*-subshells are filled with electrons with increase in atomic numbers

Sol. Answer (3)

Generally IE increases along period with few exceptions.

10. Which of the following elements do not belong to the family indicated?

(1) Cu – Coinage metal (2) Ba – Alkaline earth metal

(3) Zn – Alkaline earth metal (4) Xe – Noble gas

Sol. Answer (3)

Zn is not an alkaline earth metal, IIA group elements are called alkaline earth metals.

11. Which of the following gradation in the properties is false, as we move from left to right in the periodic table?

(1) Metallic to non-metallic character

(2) Oxidising to reducing properties

(3) Metallic solids through network solids to molecular solids

(4) Base forming to acid forming character

Sol. Answer (2)

Factual

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12. An element has 56 nucleons in nucleus and if it is isotonic with $_{30}Y^{60}$. Which group and period does it belong to?

(1) 8^{th} group, 4^{th} period
(2) 14^{th} group, 3^{rd} period
(3) 12^{th} group, 3^{rd} period
(4) 12^{th} group, 4^{th} period

Sol. Answer (1)

n + p = 56

for y, n = 60 – 30 = 30

hence for given element

n = 30

p = 26

$P(26) = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

period = 4, group number = 2 + 6 = 8

[Screening Effect]

13. Which of the following element has highest shielding constant?

(1) Mg (2) Al (3) K (4) Ga

Sol. Answer (4)

Mg $12e^-$

Al $13e^-$

K $19e^-$

Ga $31e^-$

Ga has highest atomic number (i.e. number of electrons), hence, it will have highest shielding constant.

14. Which of the following has maximum number of unpaired electrons?

(1) Mg^{2+} (2) Ti^{3+} (3) V^{3+} (4) Fe^{2+}

Sol. Answer (4)

Fe^{2+} has four unpaired electrons.

15. Which of the following orbitals has higher screening power?

(1) 4*s* (2) 4*p* (3) 4*d* (4) 4*f*

Sol. Answer (1)

s orbital have more screening power.

16. Which of the following has largest size?

(1) H^- (2) He (3) Li^+ (4) Be^{+2}

Sol. Answer (1)

H^- these are isoelectronic species.

17. The screening effect of inner electrons on the nucleus causes

(1) A decrease in the ionisation energy
(2) An increase in the ionisation energy
(3) No effect on the ionisation potential
(4) An increase in the attraction of the nucleus on the outermost electrons

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Sol. Answer (1)

Due to screening effect, repulsion on electron increases which implies that attraction of nucleus on electrons decreases, hence ionization energy decreases.

18. Ionic radii are

(1) Greater than the respective atomic radii of elements in general

(2) Greater than the respective atomic radii of electropositive elements

(3) Greater than the respective atomic radii of electronegative elements

(4) Less than the respective atomic radii of electronegative elements

Sol. Answer (3)

During formation of Anions, electron is added to neutral atom, hence size of anion is bigger, compared to neutral atom.

[Ionisation Energy]

19. The first ionisation potentials of Na, Mg, Al and Si are in the order

(1) Na < Mg > Al < Si (2) Na < Mg < Al > Si

(3) Na > Mg > Al > Si (4) Na > Mg > Al < Si

Sol. Answer (1)

Size; Mg < Na

Hence, ionization energy of Na < Mg

Size; Si < Al

Hence, ionization energy of Si > Al

Mg = $1s^2\ 2s^2\ 2p^6\ 3s^2$

Al = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

Mg has stable configuration, hence its ionization energy will be higher than Al.

$\therefore$ Na < Mg > Al < Si

20. Which among the following elements has the highest value for third ionisation energy?

(1) Mg (2) Al (3) Na (4) Ar

Sol. Answer (1)

Since, Mg belongs to IIA group hence, after removal of $2e^-$, atom will become stable, and hence, removal of 3rd electron will require high energy.

21. The ionisation potential of isotopes of an element will be

(1) Same

(2) Different

(3) Dependent on atomic masses

(4) Dependent on the number of neutrons present in the nucleus

Sol. Answer (2)

Ionisation potential of isotopes are different.

22. The element which has highest IInd I.E.?

(1) Li (2) Be (3) K (4) B

Sol. Answer (1)

IA group elements have highest IInd ionization energy.

23. Which one of the following order is correct?

(1) $I > I^+ > I^-$ (radii)

(2) $I^- > I > I^+$ (radii)

(3) $I^- > I > I^+$ (Ionisation energy)

(4) $I^{+5} < I^+ < I^{+7}$ (Ionisation energy)

Sol. Answer (2)

Anions are larger than the neutral atom, while cations are smaller than the neutral atom.

24. Enthalpy change in the following process is

$A + e^- \rightarrow A^-$ $\Delta H = -X$ kJ/mole

Which of the following processes has enthalpy change = X kJ/mole?

(1) $A^{-2} \longrightarrow A^- + 1e^-$ (2) $A + e^- \longrightarrow A^-$

(3) $A^- \longrightarrow A + e^-$ (4) $A^+ + e^- \longrightarrow A$

Sol. Answer (3)

$A + e^- \rightarrow A^-$; $\Delta H = -X$

$A^- \rightarrow A + e^-$; $\Delta H = +X$

As direction of reaction is reversed, sign of heat exchanged is also reversed.

25. Which of the following processes will release energy equal to ionization energy?

(1) $M_{(g)} \xrightarrow{+e^-} M^{(-)}_{(g)}$ (2) $M_{(s)} \xrightarrow{-e} M^{(+)}_{(s)}$ (3) $M^+_{(g)} \xrightarrow{+e^-} M_{(g)}$ (4) $M_{(g)} \xrightarrow{-e^-} M^{(-)}_{(g)}$

Sol. Answer (3)

The process is opposite of Ionization energy.

26. The ionization potential of lithium is 520 kJ/mole. The energy required to convert 70 mg of lithium atoms in gaseous state into Li^+ ions is

(1) 5.2 kJ (2) 52 kJ (3) 520 kJ (4) 52 J

Sol. Answer (1)

Number of moles of Li = $\dfrac{70 \times 10^{-3}}{7} = 10 \times 10^{-3}$ = 10^{-2} mole

$\because$ 1 mole Li requires 520 kJ energy.

$\therefore$ 10^{-2} mole Li will require = $10^{-2} \times 520$ kJ = 5.2 kJ energy

27. Which of the following configuration is associated with the biggest jump between first and second ionization energy?

(1) $1s^2 2s^2 2p^5$ (2) $1s^2 2s^2 2p^6 3s^1$ (3) $1s^2 2s^2 2p^4$ (4) $1s^2 2s^1$

Sol. Answer (4)

IA group element (ns^1) have biggest jump between Ist and IInd ionization energy.

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28. A sudden large jump between the values of second and third ionisation energies of an element would be associated with the electronic configuration

(1) $1s^2, 2s^2p^6, 3s^1$ (2) $1s^2, 2s^2p^6, 3s^2p^1$ (3) $1s^2, 2s^2p^6, 3s^2p^2$ (4) $1s^2, 2s^2p^6, 3s^2$

Sol. Answer (4)

For IIA group elements i.e., elements containing $2e^-$ in outermost shell, there is a sudden jump between values of 2nd and 3rd ionization energy (because in 3rd ionization we have to remove electron from a stable configuration)

29. The element which has highest 2nd ionisation energy is

(1) Na (2) Mg (3) Ca (4) Ar

Sol. Answer (1)

Electron will be removed from completely filled second orbit.

[Electronegativity and Electron Affinity]

30. An atom with high electronegativity generally has

(1) Low electron affinity (2) Small atomic number

(3) Large atomic radius (4) High ionisation potential

Sol. Answer (4)

Higher electronegativity implies higher ionization potential.

31. Which of the following has highest electron affinity?

(1) Na (2) Li (3) K (4) Rb

Sol. Answer (2)

32. The first electron affinity values of 'O', S & Se are given correctly as

(1) O > S > Se (2) S > Se > O (3) Se > O > S (4) Se > S > O

Sol. Answer (2)

Element	Electron gain enthalpy	Electron affinity
O	–144	144
S	–200	200
Se	–195	195

∴ S > Se > O

33. The element which has highest electron affinity?

(1) Oxygen (2) Sulphur (3) Nitrogen (4) Phosphorus

Sol. Answer (2)

Resultant of size factor and electronic configuration factor.

34. Electron gain enthalpy will be positive in

(1) O^{-2} is formed from O^{-1} (2) O^{-1} is formed from O

(3) S^{-1} is formed from S (4) Na^- is formed from Na

Sol. Answer (1)

$$O^- + e^- \longrightarrow O^{-2}$$

When e^- will approach O^-, O^- will repel the approaching electron, hence, we have to give some energy therefore, electron gain enthalpy will be +ve (endothermic).

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[Miscellaneous]

35. Which one of the following oxides has highest acidic character?

(1) CO_2 (2) Cl_2O_7 (3) SiO_2 (4) SO_2

Sol. Answer (2)

IV	**VI**	**VII**
CO_2		Cl_2O_7
SiO_2	SO_2	

As we move from left to right across a period, acidic character of oxides of elements increases.

As we move from top to bottom in a group, acidic character of oxides of element decreases.

36. Element 'A' belongs to group 16 and 5th period. Its atomic number is

(1) 34 (2) 50 (3) 52 (4) 85

Sol. Answer (3)

Element is $_{52}Te$.

37. Zr and Hf have nearly same size because

(1) They belong to same group

(2) They belong to same period

(3) Of lanthanoid contraction

(4) Of poor screening of *d* orbitals

Sol. Answer (3)

Zr and Hf have nearly same size due to lanthanide contraction.

38. Arrange the given species in the increasing order of group electronegativity

$\underset{\text{I}}{-NO_2}$, $\underset{\text{II}}{-F}$, $\underset{\text{III}}{-H}$, $\underset{\text{IV}}{-OH}$

(1) III < IV < II < I (2) I < II < IV < III (3) I < III < IV < II (4) III < I < IV < II

Sol. Answer (1)

39. The chemistry of Be is very similar to that of aluminium, because

(1) They belong to same group (2) They belong to same period

(3) Both have nearly the same ionic size (4) The ratio of their charge to size is nearly the same

Sol. Answer (4)

Be and Al show diagonal relationship hence we can say that, for these two elements charge to size ratio is nearly same.

40. If an element A shows two cationic states +2 and +3 and form oxides in such a way that ratio of the element showing +2 and +3 state is 1 : 3 in a compound. Formula of the compound will be

(1) A_8O_{11} (2) A_4O_{11} (3) A_9O_{11} (4) A_5O_{11}

Sol. Answer (1)

Total +ve charge (A_1A_3)

$= 1 \times (+2) + 3 \times (+3)$

$= 2 + 9 = 11$

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For compound to be neutral, O must contain total 11 unit negative charge.

$\because$ –2 unit charge possessed by one oxygen atom

$\therefore$ –11 unit charge possessed by = $\frac{-11\times 1}{-2} = \frac{11}{2}$ oxygen atom

Hence formula = $A_1A_3O_{\frac{11}{2}} = A_4O_{\frac{11}{2}} = A_8O_{11}$

41. Which of the following statement(s) is incorrect?

(1) The electronic configuration of Cr is [Ar] $3d^5 4s^1$ (Atomic No. of Cr = 24)

(2) The magnetic quantum number may have a negative value

(3) In silver atom, 23 electrons have a spin of one type and 24 of the opposite type (Atomic number of Ag = 47)

(4) van der Waal radius of chlorine molecule is less than its covalent radius

Sol. Answer (4)

Cr (24) = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$ (Half filled-full filled rule)

Magnetic quantum number varies from +l to –l through zero for a given value of *l*.

Ag(47) = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^1\ 4d^{10}$ (Half filled-full filled rule)

In Ag except $5s^1$, all subshells are fully filled which contains total 46e^- out of which 23e^- have anticlockwise spin, hence total number of e^- having clockwise spin

= 23 + 1($5s^1$) = 24

Total number of e^- having anticlockwise spin = 23

van der Waal's radii of a molecule is more than its covalent radii.

42. The diagonal similarities are due to similar polarising powers for the elements

The polarising power is directly proportional to

(1) $\frac{\text{ionic charge}}{\text{ionic radius}}$ (2) $\frac{(\text{ionic charge})^2}{\text{ionic radius}}$

(3) $\frac{\text{ionic charge}}{(\text{ionic radius})^2}$ (4) $\frac{\text{ionic charge}}{(\text{ionic radius})^{1/2}}$

Sol. Answer (3)

$$\text{Polarising power} \propto \frac{\text{ionic charge}}{(\text{ionic radius})^2}$$

43. Out of following, which has the highest electronegativity?

(1) H (2) Li (3) Na (4) Be

Sol. Answer (1)

H - 2.1

44. Which element is largest in size?

(1) Co (2) Ni (3) Cu (4) Zn

Sol. Answer (4)

Zn has largest radius.

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45. Hypothetically if one orbital bear three electrons then how many elements are present in 2nd period?

(1) Twelve (2) Eight (3) Six (4) Nine

Sol. Answer (1)

Three possible values of spin quantum number means an orbital can accommodate a maximum of three electrons. Since in second period 2s and 2p subshells are involved so a maximum of 12 (3 + 9) electrons can be accommodated.

46. Which of the following element has lowest value of electron affinity?

(1) Be (2) Cl (3) N (4) B

Sol. Answer (1)

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