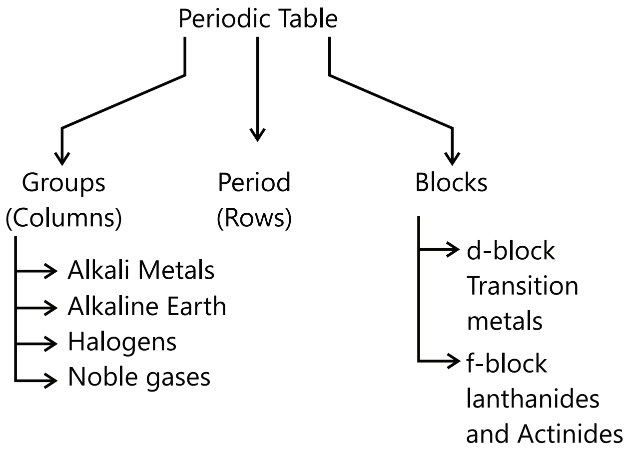
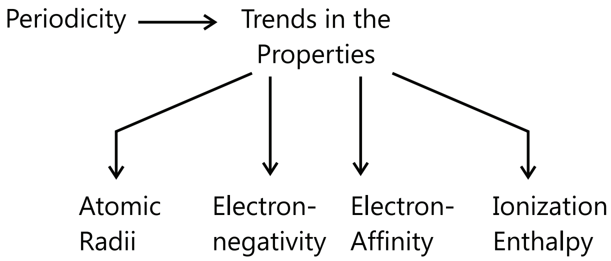
POINTS TO REMEMBER





JEE Main/Boards

Exercise 1

Q.1 The electron affinity of each one of Be and Mg is zero. Give reason.

Q.2 Explain the following

(i) The process $O(g) + e^- \rightarrow O^-(g) + 141\ kJ\ mol^{-1}$ gives a positive electron affinity though

$\Delta E = -141\ kJ\ mol^{-1}$ (exothermic).

(ii) The process $O^-(g) + e^- + 700\ kJ\ mol^{-1} \rightarrow O^{2-}(g)$ gives a negative electron affinity though ,

$\Delta E = +700\ kJ\ mol^{-1}$ (endothermic).

Q.3 There is an irregular trend in the ionization energies of group 13 elements. Explain.

Q.4 Mg^{2+}, O^{2-}, Na^+, F^-, N^{3-} (Arrange in decreasing order of ionic size)

Q.5 Why Ca^{2+} has a smaller ionic radius than K^+.

Q.6 Arrange in decreasing order of atomic size: Na, Cs, Mg, Si, Cl.

Q.7 Why the first ionization energy of carbon atom is greater than that of boron atom whereas, the reverse is true for the second ionization energy.

Q.8 (IE_1) of Be is greater than $(IE)_1$ of B but $(IE)_2$ of B is greater than that of Be. Explain.

Q.9 EA of Cu is 123 $kJ\ mol^{-1}$ but that of Zn is –87 $kJ\ mol^{-1}$. Explain.

Q.10 In the preparation of hydrocarbon by Kolbe's electrolysis, generally RCOONa or RCOOK are taken but not RCOOLi. Explain.

Q.11 The IE do not follow a regular trend in II and III periods with increasing atomic number. Why ?

Q.12 Explain why a few elements such as Be (+0.6), N (+0.3) & He(+0.6) have positive electron gain enthalpies while majority of elements have negative values.

Q.13 Which bond in each pair is more polar

(A) P–Cl or P–Br (B) S–Cl or S–O

(C) N–O or N–F

Q.14 LiCl is hydrated but NaCl is always anhydrous. Explain.

Q.15 Explain physical property of third period?

Q.16 Arrange the following according to the instructions given against each.

(i) S, O, Se, C (Increasing order of atomic radius)

(ii) Ca, Al, O, N (Increasing paramagnetism)

(iii) F, Cl^-, Br^-, I^- (Increasing polarizability)

(iv) Na, Mg and Al (Increasing electropositive character)

(v) Na, P, Si, Al, S (Decreasing order of IE)

(vi) I^-, Br^-, Cl^- (Decreasing order of ionic size)

(vii) C,F, Li, O, Cs (Increasing order of IE_1)

(viii) O, F, Be, C, N (Decreasing order of electron affinity)

(ix) Mg^{2+}, O^{2-}, Na^+, F^-, N^{3-} (Decreasing order of ionic size)

(x) Na^+, I^-, Mg^{2+}, Rb^+, Cl^- (Decreasing order of ionic radii)

Q.17 Explain the following:

(i) Fluorine is the most electronegative and Cesium is the most electropositive element.

(ii) The first ionization energy of carbon atom is greater than that of boron atom whereas, the reverse is true for the second ionization energy.

(iii) Electron affinities of halogens are high.

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

(v) Transition and inner transition metals have variable oxidation states.

(vi) Zero group elements are chemically inert.

(vii) The second ionization potential of an element is higher than its first ionization potential.

(viii) The radius of an anion is greater than the parent atom while that of a cation is lesser than the atom.

Q.18 State giving reasons which one will have a higher value

(i) IE_1 or F or Cl

(ii) EA of O or O^-

(iii) ionic radius of K^+ or Cl^-

Q.19 From among the elements, choose the following: Cl, Br, F, Al, C, Li, Cs & Xe.

(i) The element with highest electron affinity.

(ii) The element with lowest ionization potential.

Q.20 In the ionic compound KF, the K^+ and F^- ions are found to have practically radii, about 1.34 Å each. What do you predict about the relative covalent radii of K and F?

Q.21 Explain applications of electronegativity

Q.22 Define atomic size and ionic size?

Q.23 Which oxide is more basic, MgO or BaO? Why?

Q.24 The basic nature of hydroxides of group 13 (III–A) decreases progressively down the group. Comment.

Q.25 Based on location in P.T., which of the following would you expect to be acidic & which basic.

(i) CsOH (ii) IOH

(iii) $Sr(OH)_2$ (iv) $Se(OH)_2$

(v) FrOH (vi) BrOH

Q.26 Compare the following giving reasons Acidic nature of oxides: CaO, CO, CO_2, N_2O_5, SO_3.

Q.27 Write the drawback of Mendeleev table?

Q.28 Write the postulates of modern periodic table?

Q.29 Explain electron affinity of periodic tables?

Q.30 Explain factors affecting ionization potential?

Q.31 From among the elements, choose the following: Cl, Br, F, Al, C, Li, Cs & Xe.

(i) The element whose oxide is amphoteric.

(ii) The element which has smallest radii.

(iii) The element whose atom has 8 electrons in the outermost shell.

Exercise 2

Single Correct Choice Type

Q.1 When the following five anions are arranged in order of decreasing ionic radius, the correct sequence is:

(A) Se^{2-}, I^{-}, Br^{-}, O^{2-}, F^{-} (B) I^{-}, Se^{2-},O^{2-}, Br^{-}, F^{-}

(C) Se^{2-}, I^{-}, Br^{-}, F^{-} ,O^{2-} (D) I^{-}, Se^{2-},Br^{-}, O^{2-}, F^{-}

Q.2 Which of the following is wrong regarding the stability of the ions Ge,Sn and Pb

(A) $Ge^{2+} < Sn^{2+} < Pb^{2+}$ (B) $Ge^{4+} > Sn^{4+} > Pb^{4+}$

(C) $Sn^{4+} > Sn^{2+}$ (D) $Pb^{2+} < Pb^{4+}$

Q.3 The law of triads is not applicable on

(A) Cl, Br, I (B) Na, K, Rb

(C) S, Se, Te (D) Ca, Sr, Ba

Q.4 The atomic volume was choosen as the basic of periodic classification of elements by

(A) Niels Bohr (B) Mendeleev

(C) Lothar meyer (D) Newlands

Q.5 The majority of gaseous elements in the periodic table are placed

(A) At bottom left hand side

(B) At top right hand side

(C) Below the main table

(D) Along side d block elements

Q.6 The electronic configuration

$1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $3d^{10}$, $4s^2$, $4p^6$, $4d^{10}$, $5s^2$ is for:

(A) f-block element (B) d-block element

(C) p-block element (D) s-block element

Q.7 Fluorine has the highest electronegativity among the ns^2 np^5 group on the Pauling scale, but the electron affinity of fluorine is less than that of chlorine because:

(A) The atomic number of fluorine is less than that of chlorine

(B) Fluorine being the first member of the family behaves in an unusual manner

(C) Chlorine can accommodate an electron better than fluorine by utilizing its vacant 3d orbital

(D) Small size, high electron density and an increased electron repulsion makes addition of an electron to fluorine less favourable than that in the case of chlorine.

Q.8 The last electron in each normal elements of period is filled in

(A) The same energy sublevel

(B) The same enrgy level

(C) The same orbital

(D) Relation between I_x and I_z is uncertain

Q.9 The greater stability of the lower oxidation state in heavier p block metals in the consequence of

(A) Electronic transition within p-orbitals

(B) Electronic transition from s to p-orbitals

(C) Inert pair effect

(D) Expansion of octet

Q.10 Oxidation number of p-block elements is [Excluding inert gases]

(A) Equal to group number

(B) Group number +2

(C) Between the range [Group no...(Group no. 8)]

(D) Number of unpaired electrons in the valence shell

Q.11 The correct order of second ionization potential of carbon, nitrogen, oxygen and fluorine is:

(A) C > N > O > F (B) O > N > F > C

(C) O > F > N > C (D) F > O > N > C

Q.12 Which statement is wrong:

(A) 2nd ionization energy shows jump in alkali metals

(B) 2nd electron affinity for halogens is zero

(C) Maximum electron affinity exists for F

(D) Maximum ionization energy exists for He

Q.13 Which of the following is the configuration of second excited state of the element isoelectronic with O_2 or P^- or Cl^+

(A) $[Ne]3s^2 3p_x^1 3p_y^1 3p_z^1$

(B) $[Ne]3s^2$

(C) $[Ne]3s^1 3p_z^1 3p_y^1 3p_z^1 3d^1_{xy} 3d^1_{yz}$

(D) $[Ne] 3s^1 3p_x^1 3p_z^1 3d^1_{xy}$

Q.14 Metallic radii of transition elements

(A) First increase, then decrease periodically

(B) First decrease, then remain almost constant

(C) First increase, then remaining almost constant

(D) First increase, then increase periodically

Q.15 For the formation of a covalent bond, the difference in the value of electro negativities should be:

(A) Equal to or less than 1.7 (B) More than 1.7

(C) 1.7 or more (D) None of these

Q.16 Which one of the following is the correct order of interactions?

(A) Covalent < hydrogen bonding < Vander Waal's < dipole-dipole

(B) Van der Waal's < hydrogen bonding< dipole-dipole < covalent

(C) Van der Waal's < dipole-dipole < hydrogen bonding < covalent

(D) Dipole-dipole < Van der Waal's < hydrogen

Q.17 Properties of the elements of which of the following pairs do not resemble?

(A) Li and Mg (B) Be and Al

(C) Mg and Al (D) B and Si

Q.18 Which of the following statement is not true?

(A) The atoms have no tendency to accept electrons in empty higher energy levels

(B) The atoms have no tendency to accept electrons in empty high energy sublevels

(C) The alkali metals have no tendency to accept electrons

(D) The atoms with exactly half-filled electronic configurations have no tendency to accept electrons.

Q.19 IP_1 and IP_2 of Mg are 178 and 348 kcal mol^{-1}. The energy required for the reaction, $Mg \rightarrow Mg^{2+} + 2e^-$

(A) +170 kcal (B) +526 kcal

(C) $E- 2E_2$ (D) $(E_1 - E_2)/2$

Q.20 Electronic configuration of an element of atomic weight 40 is 2, 8, 8, 2 which of the following statement regarding this element is not correct

(A) It belongs to the second group of the P.T

(B) It has 20 neutrons

(C) The formula of its oxide is MO_2

(D) It has 20 Protons

Q.21 There are four elements P, Q, R and S: their configuration are also given. Show that which element will have highest value of IE_2?

(A) P = $[He]2s^2$ (B) Q = $[He] 2s^2 2p^2$

(C) R = $[He] 2s^2 2p^1$ (D) S = $[He] 2s^1 2p^3$

Q.22 The ionization potential of nitrogen is greater than that of oxygen because

(A) Nitrogen is an inert element

(B) The outermost shell of nitrogen has half-filled orbitals

(C) The radius of nitrogen is more than that of oxygen

(D) The radius of oxygen is more than that of nitrogen

Q.23 Be and Mg have zero value of electron affinities, because

(A) Be and Mg have [He] $2s^2$ and [Ne] $3s^2$ configuration respectively

(B) 2s and 3s orbitals are filled to their capacity

(C) Be and Mg are unable to accept electron

(D) All the above are correct

Q.24 If I_1 and I_3 etc. represent the successive ionization potentials of an atom then the correct order is.

(A) $I_1 > I_2 > I_3$ (B) $I_1 < I_2 > I_3$ (C) $I_1 < I_2 < I_3$ (D) $I_2 > I_1 > I_3$

Q.25 For I.P, which order is wrong

(1) F > O (2) O > N (3) S > P (4) Be > B

Code is –

(A) 1, 2, 3 (B) 2, 3 (C) 1, 4 (D) 1, 2, 4

Q.26 Alkali metals do not form dipositive ions, because

(A) The difference in the first and second I.P. is more than 16 eV

(B) The difference in the first and second I.P. is less than 11 eV

(C) Alkali metals have one electron in their ultimate energy level

(D) Oxidation state of alkali metals is +1

Q.27 Electronic configuration of X^{+2} and Y^{+3} are: X^{+2} = [Ar] $3d^8$, Y^{+3} = [Ar] $3d^3$. What are the atomic number of X^0 and Y^0 respectively.

(A) 28, 24 (B) 28, 25 (C) 28, 26 (D) 28, 27

Q.28 The electronegativity values of C, N, O and F

(A) Increase from carbon to fluorine

(B) Decrease from carbon to fluorine

(C) Increase up to oxygen and minimum at a fluorine

(D) Is minimum at nitrogen and then increase continuously

Q.29 The valency in the II period from left to right

(A) Increases

(B) Decreases

(C) First increases, then decreases

(D) First decreases, then increases

Q.30 Which of the following represents incorrect relation of electronegativity

(A) C > O > N (B) C < O >N

(C) O > C < N (D) O > N > C

Q.31 In a period, elements are arranged in strict sequence of

(A) Decreasing charges in the nucleus

(B) Increasing charges in the nucleus

(C) Constant charges in the nucleus

(D) Equal charges in the nucleus

Q.32 In the periodic table, the metallic character of elements

(A) Decreases across a period and down a group

(B) Decreases across a period and increases down a group

(C) Increases across a period and down a group

(D) Increases across a period and decreases down a group

Q.33 The screening effect of inert electrons of the nucleus causes

(A) A decrease in the ionization potential

(B) An increase in the ionization potential

(C) No effect on the ionization potential

(D) An increase in the nuclei attraction of the electrons

Q.34 The statement that is not correct for the periodic classification of elements is

(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 ionization energies along a period vary in a regular manner with increase in the atomic number

(D) For transition elements, the d-sub-shells are filled with electrons monotonically with an increase in atomic number

Q.35 Match List I with List II and select the correct answer using the codes given below the lists

List I	List II
(i) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$	(p) In
(ii) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$	(q) In
(iii) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6 4d^{10}$	(r) Ca
(iv) $1s^2\ 2p^2\ 2p^6\ 3s^2\ 3d^{10}\ 4s^2\ 4s^2\ 4p^{10} 5s^2 5p^1$	(s) Cu

Codes:	i	ii	iii	iv
(A)	p	q	r	s
(B)	p	r	q	s
(C)	r	s	q	p
(D)	p	s	r	q

Q.36 Match List-I with List-II and select the correct answer from the codes given below the lists

List I	List II
(A) ns^2, np^5	(p) Chromium
(B) $(n–)d^{10}$, ns^1	(q) Copper
(C) $(n–1)d^{10}$, ns^2, np^6	(r) Krypton
(D) $(n–1)d^{10}$, $ns^{2,}\ np^6$	(s) Bromine (n = 4)

Q.37 Match List I with List II and select the correct answer from the codes given below the lists

List I	List II
(A) Highest ionization potential	(p) Technitium
(B) Highest electronegativity	(q) Lithium
(C) Artificial element	(r) Helium
(D) High reducing ability	(s) Fluorine

Q.38 The correct order of relative basic character of NaOH, $Mg(OH)_2$ and $Al(OH)_3$ is–

(A) $Al(OH)_3 > Mg(OH)_2 > NaOH$

(B) $Mg(OH)_2 > NaOH > Al(OH)_3$

(C) $NaOH > Mg(OH)_2 > Al(OH)_3$

(D) $Al(OH)_3 > NaOH < Mg(OH)_2$

Q.39 Match List–I with List–II and select the correct answer from the codes given below the lists

List I	List II
(i) Increasing atomic size	(p) Cl < O < F
(ii) Decreasing atomic radius	(q) Li < Be < B
(iii) Increasing electronegativity	(r) Si < Al < Mg
(iv) Decreasing effective nuclear charge	(s) Na> N > F

Codes:	i	ii	iii	iv
(A)	r	s	p	q
(B)	s	q	r	p
(C)	p	q	r	s
(D)	q	p	s	r

Previous Years Questions

Q.1 Which of the following statement is correct with respect to the property of element with an increase in atomic number in the carbon family (group 14) ***(2004)***

(A) Atomic size decrease

(B) Ionisation energy increase

(C) Metallic character decrease

(D) Stability of +2 oxidation state increase

Q.2 The first ionization potentials in electron volts of nitrogen and oxygen atoms are respectively given by ***(1987)***

(A) 14.6, 13.6 (B) 13.6, 14.6

(C) 13.6, 13.6 (D) 14.6, 14.6

Q.3 The elements which occupy the peaks of ionization energy curve, are ***(2000)***

(A) Na, K, Rb, Cs (B) Na, Mg, Cl, I

(C) Cl, Br, I, F (D) He, Ne, Ar, Kr

Q.4 Arrange F, Cl, O, N in the decreasing order of electronegativity

(A) O > F > N > Cl (B) F > N > Cl > O

(C) Cl > F > N > O (D) F > O > N > Cl

Q.5 Ionic radii of ***(1999)***

(A) $Ti^{4+} < Mn^{7+}$ (B) $^{35}Cl^- < {}^{37}Cl^-$

(C) $K^+ > Cl^-$ (D) $P^{3+} > P^{5+}$

Q.6 In which block, 106th element belongs

(A) s-block (B) p-block

(C) d-block (D) f-block

Q.7 The first $(\Delta_1 H_1)$ and second $(\Delta_1 H_2)$ ionization enthalpies (in kJ mol^{-1}) and the electron gain enthalpy $(\Delta_{eg}H)$ (in kJ mol^{-1}) of the elements I, II, III, IV and V are given below

Element	$\Delta_1 H_1$	$\Delta_1 H_2$	$\Delta_{eg}H$
I	520	7300	–60
II	419	3051	–48
III	1681	3374	–328
IV	1008	1846	–295
V	2372	5251	+48

The most reactive metal and the least reactive non-metal of these are respectively ***(2010)***

(A) I and V (B) V and II

(C) II and V (D) V and II

(E) V and III

Read the assertion and reason carefully to mark the correct option out of the options given below:

(A) If both assertion and reason are true and the reason is the correct explanation of the assertion.

(B) If both assertion and reason are true but reason is not the correct explanation of the assertion.

(C) If assertion is true but reason is false.

(D) If the assertion and reason both are false.

(E) If assertion is false but reason is true.

Q.8 Assertion: Positive ions will be wider than parent atoms ***(1999)***

Reason: Nuclear charge pulls them closer.

Q.9 Assertion: More is the electron affinity greater is the reducing character. ***(2000)***

Reason: Reducing character depends on number of electrons gained.

Q.10 Assertion: Ground state configuration of Cr is $3d^5, 4s^1$. ***(2004)***

Reason: A set of half filled orbitals containing one electron each with their spin parallel provides extra stability to the system.

Q.11 Assertion: I.E. Of $_7N$ is more that of $_8O$ as well as $_6C$. ***(2005)***

Reason: This is due to difference in reactivity towards oxygen.

Q.12 Which one of the following constitutes a group of the isoelectronic species? ***(2008)***

(A) C_2^{2-}, O_2^-, CO, NO (B) $NO^+, C_2^{2-}, CN^-, N_2$

(C) $CN^-, N_2, O_2^{2-}, C_2^{2-}$ (D) N_2, O_2^-, NO^+, CO

Q.13 Which one of the following pairs of species have the same bond order? ***(2008)***

(A) CN^- and NO^+ (B) CN^- and CN^+

(C) O_2^- and CN^- (D) NO^+ and CN^+

Q.14 In which of the following arrangements, the sequence is not strictly according to the property written against it? ***(2009)***

(A) $CO_2 < SiO_2 < SnO_2 < PbO_2$: increasing oxidising power

(B) $HF < HCl < HBr < HI$: increasing acid strength

(C) $NH_3 < PH_3 < AsH_3 < SbH_3$: increasing basic strength

(D) $B < C < O < N$: increasing first ionization enthalpy.

Q.15 The set representing the correct order of ionic radius is: ***(2009)***

(A) $Li^+ > Be^{2+} > Na^+ > Mg^{2+}$

(B) $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$

(C) $Li^+ > Na^+ > Mg^{2+} > Be^{2+}$

(D) $Mg^{2+} > Be^{2+} > Li^+ > Na^+$

Q.16 The bond dissociation energy of B – F in BF_3 is 646 kJ mol^{-1} whereas that of C-F in CF_4 is 515 kJ mol^{-1}. The correct reason for higher B-F bond dissociation energy as compared to that of C- F is: ***(2009)***

(A) Smaller size of B-atom as compared to that of C- atom

(B) Stronger σ bond between B and F in BF_3 as compared to that between C and F in CF_4

(C) Significant $p\pi - p\pi$ interaction between B and F in BF_3 whereas there is no possibility of such interaction between C and F in CF_4.

(D) Lower degree of $p\pi - p\pi$ interaction between B and F in BF_3 than that between C and F in CF_4.

Q.17 The correct sequence which shows decreasing order of the ionic radii of the elements is ***(2010)***

(A) $Al^{3+} > Mg^{2+} > Na^+ > F^- > O^{2-}$

(B) $Na^+ > Mg^{2+} > Al^{3+} > O^{2-} > F^-$

(C) $Na^+ > F^- > Mg^{2+} > O^{2-} > Al^{3+}$

(D) $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$

Q.18 Among the following the maximum covalent character is shown by the compound: ***(2011)***

(A) $SnCl_2$ (B) $AlCl_3$

(C) $MgCl_2$ (D) $FeCl_2$

Q.19 Which one of the following order represents the correct sequence of the increasing basic nature of the given oxides? ***(2011)***

(A) $MgO < K_2O < Al_2O_3 < Na_2O$

(B) $Na_2O < K_2O < MgO < Al_2O_3$

(C) $K_2O < Na_2O < Al_2O_3 < MgO$

(D) $Al_2O_3 < MgO < Na_2O < K_2O$

Q.20 The molecule having smallest bond angle is: ***(2012)***

(A) NCl_3 (B) $AsCl_3$ (C) $SbCl_3$ (D) PCl_3

Q.21 Which of the following on thermal decomposition yields a basic as well as an acidic oxide? ***(2012)***

(A) $NaNO_3$ (B) $KClO_3$ (C) $CaCO_3$ (D) NH_4NO_3

Q.22 Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar? ***(2013)***

(A) Ca < S < Ba < Se < Ar

(B) S < Se < Ca < Ba < Ar

(C) Ba < Ca < Se < S < Ar

(D) Ca < Ba < S < Se < Ar

Q.23 The first ionisation potential of Na is 5.1 eV. The value of electron gain enthalpy of Na^+ will be: ***(2013)***

(A) -2.55 eV (B) -5.1 eV

(C) -10.2 eV (D) +2.55 eV

Q.24 The ionic radii (in Å) N^{3-}, O^{2-} and F^- are respectively: ***(2015)***

(A) 1.36, 1.40 and 1.71

(B) 1.36, 1.71 and 1.40

(C) 1.71, 1.40 and 1.36

(D) 1.71, 1.36 and 1.40

JEE Advanced/Boards

Exercise 1

Q.1 Use the following system of naming elements in which first alphabets of the digits are written collectively,

0	1	2	3	4	5	6	7	8	9
nil	uni	bi	tri	quad	pent	hex	sept	oct	enn

to write the three-letter symbols for the elements with atomic number 101 to 109.

Q.2 Arrange the following in increasing order of the property as indicated:

(i) Pb, Pb^{2+} and Pb^{4+}(size)

(ii) Mg, Al, Si and Na(ionisation potential)

(iii) MgO, SrO, Rb_2O, NiO, Cs_2O (basic character)

(iv) Be(OH), $Mg(OH_2$, $Ca(OH)_2$, $Ba(OH)_2$ (basicity)

(v) Cl^-, K^+, Ca^{2+}, Ar (ionization energy)

(vi) As, F, S, Cl (Electronegativity)

(vii) LiCl, LiBr, LiI (Ionic character)

(viii) Li^+, Na^+, K^+, Rb^+, Cs^+ (mobility of hydrated ions)

(ix) Li, Na, K, Rb and Cs (hydrated radii)

(x) Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+} (hydration of ions)

Q.3 Why are inert gases mono-atomic?

Q.4 Why does the third period contains 8 elements and not 18?

Q.5 Arrange the following in order of increasing ionic radius:

(i) Cl^-, P^{3-}, S^{2-}, F^- (ii) $Al^{3+,}$ Mg^{2+}, Na^+, O^{2-}, F^-

(iii) Na^+, Mg^{2+}, K^+.

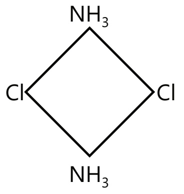
Q.6 If internuclear distance between Cl atoms in Cl_2 is 10 Å & between H atoms is H_2 is 2 Å, then calculate internuclear distance between H and Cl (Electronegativity of H = 2.1 & Cl = 3.0).

Q.7 The As-Cl bond distance in $AsCl_3$ is 2.20 Å. Estimate the SBCR (single bond covalent radius) of As.

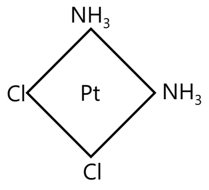
(Assume EN of both to be same and radius of Cl = 0.99 Å.)

Q.8 The Pt–Cl distance has been found to be 2.32 Å in several crystalline compounds. If this value applies to both of the compounds shown in figure, what is Cl–Cl distance in (i) and (ii)

(i)



(ii)



Q.9 K^+ and F^- have identical radius about 1.34 Å. What should be the atomic radius of K and F atoms.

Q.10 The radii of noble gases are greater than that of the radii of precedent halogens.

Q.11 The covalent radii (Å) do not increases regularly from B to Tl (B = 0. 80 Å, Al = 1.25 Å; Ga = 1.25 Å, In = 1.50 Å ; Tl = 1.55 Å) How do you account for this anomalous behavior.

Q.12 Calculate E.N. of chlorine atom on Pauling scale if I.E. of Cl^- is 4eV & of E.A. of Cl^+ is +13.0 eV.

Q.13 Why in isoelectronic species, the ionic / atomic radii decrease with increase in atomic numbers?

Q.14 The ionisation energy of the coinage metals fall in the order Cu > Ag < Au.

Q.15 The first ionization energy of carbon atom is greater than that of boron where as the reverse is true for second ionization energy. Explain.

Q.16 Define ionization potential?

Q.17 The Ist ionization energy of potassium is less than that of Cu but the reverse is true for IInd ionisation energy.

Q.18 The sums of first and second ionization energies and those of third fourth ionisation energies in (kJ mol^{-1}) of nickel and platinum are

$(IE)_1 + (EI)_2$	$(IE)_3 + (IE)_4$
Ni 2.49	8.80
Pt 2.66	6.70

Based on this information; write

(i) The most stable oxidation states of Ni and Pt

(ii) Name of one of the metals which can more easily form compounds in its +4 oxidation state.

Q.19 The first ionization enthalpy of magnesium is higher than of sodium. On the other hand, the second ionization enthalpy of sodium is very much higher than that of magnesium. Explain.

Q.20 Na^+ has higher value of ionization energy than Ne, though both have same electronic configuration. Explain?

Exercise 2

Single Correct Choice Type

Q.1 Two p-block elements x (outer configuration ns^2 , np^3) and z (outer configuration ns^2 np^4) occupy neighbouring positions in a period. Using this information, which of the following is correct with respect to their ionization potential I_x and I_z.

(A) $I_x > I_z$ (B) I_z and I_x (C) $I_z = I_x$ (D) $I_z > I_x$

Q.2 True position of lanthanides are:

(A) After III B group and in the 6th period

(B) After III B group in the 3rd period

(C) After VI B group in the 3rd period

(D) After III B group in the 4th Period

Q.3 The main cause of diagonal relationship between Be and Al is

(A) Similarity in ionic sizes

(B) Similar ionic potentials

(C) Similar electronegativity

(D) Similar atomic volume

Q.4 Electronic configuration of four elements are: a = $1s^2$; $2s^2$, $2p^1$; b = $1s^2$; $2s^2$, $2p^2$; c = $1s^2$; $2s^2$, $2p^5$; d = $1s^2$; $2s^2$, $2p^6$; $3s^1$. Which one of these would most readily form diatomic molecule?

(A) a (B) b (C) c (D) d

Q.5 EN of the element A is E_1 and IP is E_2. Hence EA will be:

(A) $2E_1 - E_2$ (B) $E_1 - E_2$ (C) $E_1 - 2E_2$ (D) $E_1 + E_2$

Q.6 $\boxed{\uparrow\downarrow}\quad\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$

The above configuration would be of the species

(i) S (ii) Cl^+ (iii) p (iv) Ar^{-2}

Code is –

(A) i, ii (B) iiii, iv (C) i, iv (D) ii, iii

Q.7 Which of the following statements is correct for the addition of an electron to an isolated and gaseous uni-negatively charged oxygen O^- ion?

(A) The addition of electron cannot occur

(B) The addition of electron occurs with evolution of energy

(C) The addition of electron occurs with absorption of energy

(D) None of the above

Q.8 In the transformation Na(s) $\rightarrow$ Na^+(g), the energies involved are

(A) Ionization energy

(B) Sublimation energy

(C) Ionization energy and sublimation energy both

(D) None of the above

Q.9 The process of requiring absorption of energy is

(A) $F \rightarrow F^-$ (B) $Cl \rightarrow Cl^-$

(C) $O^- \rightarrow O^{-2}$ (D) $H \rightarrow H^-$

Q.10 The cynide CN^- & N_2 are isoelectronic. But in contrast to CN– , N_2 is chemically inert because of –

(A) Low bond energy

(B) Absence of bond polarity

(C) Unsymmetrical electron distribution

(D) Presence of more electron in bonding

Previous Year Questions

Q.1 The correct order of radii is ***(2000)***

(A) N < Be < B (B) $F^- < O^{2-} < N^{3-}$

(C) Na < Li < K (D) $Fe^{3+} < Fe^{2+} < Fe^{4+}$

Q.2 The set representing the correct order of first ionization potential is ***(2001)***

(A) K > Na > Li (B) Be > Mg > Ca

(C) B > C > N (D) Ge > Si > C

Q.3 Identify the least stable ion amongst the following ***(2000)***

(A) Li^+ (B) Be^- (C) B^- (D) C^-

Q.4 The statements that is/are true for the long form of the periodic table is/are ***(1988)***

(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 tends in physical and chemical properties of the elements

(D) It helps to predict the relative iconicity of the bond between any two elements

Q.5 Sodium sulphate is soluble in water whereas barium sulphate is sparingly soluble because ***(1989)***

(A) The hydration energy of sodium sulphate is more 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

Q.6 The softness of group IA metals increases down the group with increasing atomic number. ***(1986)***

Q.7 In group IA of alkali metals, the ionization potential decreases down the group. Therefore, lithium is a poor reducing agent. ***(1987)***

Q.8 Arrange the following in order of their ***(1985)***

(i) Decreasing ionic size Mg^{2+}, O^{2-}, Na^+, F^-

(ii) Increasing first ionization energy Mg, Al, Si, Na

(iii) Increasing bond length F_2, N_2 Cl_2, O_2

Q.9 Arrange the following in the order of their increasing size: ***(1986)***

Cl^-, S^{2-}, Ca^{2+}, Ar

Q.10 Compare qualitatively the first and second ionization potentials of copper and zinc. Explain the observation. ***(1996)***

Q.11 Arrange the following ions in order of their increasing radii ***(1997)***

Li^+, Mg^{2+}, K^+, Al^{3+}.

Q.12 The species having bond order different from that in CO is ***(2007)***

(A) NO^- (B) NO^+ (C) CN^- (D) N_2

Q.13 Statement-I: Pb^{4+} compounds are stronger oxidizing agents than Sn^{4+} compounds. ***(2008)***

Statement-II: The higher oxidation states for the group 14 elements are more stable for the heavier members of the group due to 'inert pair effect'.

(A) Statement-I is True, Statement-II is True; Statement-II is correct explanation for Statement-I

(B) Statement-I is True, Statement-II is True; Statement-II is NOT a correct explanation for Statement-I

(C) Statement-I is True, Statement-II is False

(D) Statement-I is False, Statement-II is True

Q.14 Among the following, the number of elements showing only one non-zero oxidation state is: ***(2010)***

O, Cl, F, N, P, Sn, Tl, Na, Ti

Q.15 The increasing order of atomic radii of the following Group 13 elements is ***(2016)***

(A) Al < Ga < In < Ti (B) Ga < Al < Tl < In

(C) Al < In < Ga < Ti (D) Al < Ga < Tl < In

MASTERJEE Essential Questions

JEE Main/Boards

Exercise 1

Q.1 Q.9 Q.10 Q.14
Q.16 Q.17 (D) Q. 21

Exercise 2

Q.13 Q.16 Q.23 Q.26 Q.33
Q.35 Q.37

Previous Years' Questions

Q.7 Q.10

JEE Advanced/Boards

Exercise 1

Q.9 Q.11 Q.14

Exercise 2

Q.5 Q.10

Previous Years' Questions

Q.5 Q.6

Answer Key

JEE Main/Boards

Exercise 2

Single Correct Choice Type

Q.1 D **Q.2** D **Q.3** B **Q.4** C **Q.5** B **Q.6** D **Q.7** D
Q.8 B **Q.9** C **Q.10** C **Q.11** C **Q.12** C **Q.13** C **Q.14** B
Q.15 A **Q.16** C **Q.17** C **Q.18** D **Q.19** B **Q.20** C **Q.21** C
Q.22 B **Q.23** D **Q.24** C **Q.25** B **Q.26** A **Q.27** A **Q.28** A
Q.29 C **Q.30** A **Q.31** B **Q.32** B **Q.33** A **Q.34** C **Q.35** C
Q.36 C **Q.37** A **Q.38** C **Q.39** A

Previous Years' Questions

Q.1 D **Q.2** A **Q.3** D **Q.4** D **Q.5** D **Q.6** C **Q.7** C
Q.8 D **Q.9** E **Q.10** A **Q.11** C **Q.12** B **Q.13** A **Q.14** C
Q.15 B **Q.16** C **Q.17** D **Q.18** B **Q.19** D **Q.20** C **Q.21** C
Q.22 C **Q.23** B **Q.24** C

JEE Advanced/Boards

Exercise 2

Single Correct Choice Type

Q.1 A **Q.2** A **Q.3** B **Q.4** C **Q.5** A **Q.6** A **Q.7** C
Q.8 C **Q.9** C **Q.10** B

Previous Years' Questions

Q.1 B **Q.2** B **Q.3** B **Q.4** B, C, D **Q.5** A, B **Q.6** True **Q.7** False
Q.12 A **Q.13** C **Q.14** 2 **Q.15** B

Solutions

JEE Main/Boards

Exercise 1

Sol 1: As we know, the electronic configuration of Be and Mg are [He] $2s^2$ and [Ne] $3s^2$ respectively. As we see, both of these elements have a fully filled s-orbital & making their configuration very stable compared to other configuration with incompletely filled orbitals. Therefore, they have very less tendency to gain an electron, which makes their electron affinity equal to zero.

Sol 2: The electronic configuration of oxygen is [He] $2s^22p^4$,which is not very stable. Since this is neither a half filled or a full filled configuration, therefore it has a natural tendency to accept an electron which makes it first electron affinity positive and the process exothermic but, after an electron is added, the atom becomes negatively charged and the addition of a second electron is opposed by coulombic repulsion. Hence energy has to be supplied for the addition of second electron. Thus, its second electron affinity is negative and the process is endothermic.

Sol 3: We know that the ionisation energy increases while going from left to right in a period. But, there is an exception while going from group 12 to group 13, which means that ionisation energy decreases from group 12 to group 13. This exception is explained by the high stability of group 12 elements due to fully filled orbitals which makes it difficult to remove the outermost electron compared to group 13 where there are no full filled orbitals. Hence, their ionization energy is smaller.

Sol 4: All these ions have the same electronic configuration, i.e. [Ne].So, difference in size will be due to electronic charge only. Now, we know that the ionic size decreases with increasing positive charge, because a loss of electron results in increased effective nuclear charge and hence the remaining electrons are held more firmly bound with the nucleus, hence decreasing ionic size. Also, ionic size increases with increasing ionic charge, because addition of an electron results in an increase in natural repulsion between electrons which decreases effective nuclear charge and leads to increase in ionic size. Therefore, the order of ionic size will be $N^{3-} > O^{2-} > F^- > Na^+ > Mg^{2+}$

Sol 5: Both Ca^{2+} and K^+ have the same electronic configuration, i.e. [Ar] and hence equal number of shells. The ionic radius decreases with increase in positive charge, because the loss of an electron results in an increase in effective nuclear charge and hence a decrease in radius as the remaining electrons are more firmly bound to the nucleus. If 2 electrons are removed, the effective nuclear charge will increase even more and the size of the ion will decrease further. Therefore, the ionic radius of Ca^{2+} is less than K^+

Sol 6: All the elements except Cs are period 3 elements and we know that atomic size decreases in going from left to right in a period due to increased effective nuclear charge, resulting from the increase in number of protons for same number of electronic shells. Also, the atomic radius of Cs will be bigger than the rest because of more number of shells is Cs. Hence order

Cs > Na > Mg > Si > Cl

Sol 7: Electronic configuration of carbon: [He] $2s^22p^2$

Electron configuration of Boron: [He] $2s^22p^1$

First ionization energy of carbon is more than Boron because of increased effective nuclear charge in carbon, making removal of electron difficult. After removal of electron:

Electronic configuration of C^+: [He] $2s^22p^1$

Electronic configuration of B^+: [He] $2s^2$

Now B^+ has a stable full filled orbital making its configuration very stable compared to incompletely filled configuration of C^+. Hence second ionization energy of Boron is more than Carbon

Sol 8: Electronic configuration of Be: [He] $2s^2$

Electronic configuration of B: [He] $2s^2$ $2p^1$ (IE), of Be > (IE) of B due to stable full filled orbital configuration of Be making removal of electron difficult.

After removal of electron

Electronic configuration of Be^+: [He] $2s^1$

Electronic configuration of B^+: [He] $2s^2$

Here, $(IE)_2$ of B > $(IE)_2$ of Be as now B^+ has a more stable configuration than Be^+ due to full filled orbitals which are not present is Be^+

Sol 9: Electronic configuration of Cu: [Ar] $3d^2$ $4s^2$

Electronic configuration of Zn: [Ar] $3d^{10}$ $4s^2$

Zn has a much more stable configuration than Cu because of fully filled d orbital compared to Cu. A stable configuration means that Zn has no tendency to accept an electron which makes its EA negative. Cu on the other hand can attain a stable configuration by adding an electron, hence its EA is positive.

Sol 10: For Kolbe's electrolysis, we want the acid to dissociate into ions easily during the reaction. But since the acidic strength of carbonic acids increases while going from top to bottom in a group, therefore, RCOOLi will be dissociated into ions to a very little extent compared to RCOONa or RCOOK, Hence RCOOLi is not used.

Sol 11: The general trend of IE is that it increases while going from left to right in a period. However IE is more for an element which has a stable configuration due to half-filled or full filled orbitals. Hence, IE of group 7 > IE of group 8 and IE of group 4 > IE of group 5 because group 4 and group 7 have full filled and half-filled orbital respectively.

Sol 12: Electronic configuration of Be: [He] $2s^2$

Electronic configuration of N: [He] $2s^2$ $2p^3$

Electronic configuration of He: [He]

We see, that each of those elements have a full filled or half-filled orbital which makes their configuration very stable and because of this, they have no tendency to accept an electron and hence they have negative E_A and hence positive value of electron gain enthalpy.

Sol 13: Polarity of the bond depends upon the electronegativity difference between the two elements. The more electronegativity difference will be the more polar the bond will be

(A) P – Cl > P – Br as $(EN)_{Cl} > (EN)_{Br}$

$= [(EN)_{Cl} – (EN)_P] > [(EN)_{Br} – (EN)_S]$

(B) S – O > S – Cl as $(EN)_O > (EN)_{Cl}$

$= [(EN)_O – (EN)_S] > [(EN)_{Cl} – (EN)_p]$

(C) N – F > N – O as $(EN)_F > (EN)_O$

$= [(EN)_F – (EN)_N] > [(EN)_F – (EN)_O]$

Sol 14: LiCl is hydrated because hydration energy of Li^+ is much higher than Na^+ due to small size of Li^+ ion. Hence, NaCl is amorphous because of low hydration enthalpy of Na^+.

Sol 15: Atomic Radius: The atomic radius decreases while going from left to right till because of increased effective nuclear charge which pulls the electrons closer to the nucleus and hence decreases atomic radius. Atomic radius of Ar is larger because Vander Waals radius is calculated for Ar as it does not form any strong bonds.

First Ionisation Energy: (IE), generally increases across the period from left to right due to increase in effective nuclear charge and poor screening of outer most electrons. However (IE), of Mg > (IE), of Al or Mg has a stable configuration because of full filled orbital and (IE), of P > (IE), of S as P has half-filled orbital making its configuration stable.

Electron Affinity (EA): Electron affinity increases across the period due to increased nuclear charge and decreasing atomic size. However EA of $S_A > E_A$ of P, due to half-filled orbital stable configuration of P which makes addition of electron difficult.

Electronegativity: It increases across the period due to the decrease in atomic size and hence increased effective nuclear charge.

Sol 16: (i) O < C < S < Se

Since, atomic radius increases while going down in a group O < S < Se. Also since atomic radius decreases across a period O < C .

(ii) Paramagnetism will increase with an increase in number of unpaired electrons.

∴ Ca (0 unpaired electron) Al (1 unpaired) < O (2 unpaired) < N (3 unpaired)

Ca < O < N

(iii) The order of polarisability will depend on the order of ionic size. The larger the ionic size, more will be polarizability.

∴ Order - $F^- < Cl^- < Br^- < I^-$

(iv) Al < Mg < Na

Since electropositive character decreases across a period.

(v) P > S > Si > Al > Na

I.E. increases across the period

I.E. of P > I.E. of S due to stable half-filled orbital configuration of P

(vi) $I^- > Br^- > Cl^-$

Ionic radius increases down the period.

(vii) Cs < Li < C < O < F

IE, increases across the period.

(viii) F > O > C > N > Be

Electron affinity increases across a period $(EA)_C > (EA)_N$ due to stable configuration of N (half-filled orbital)

(ix) $N^{3-} > O^{2-} > F^- > Na^+ > Mg^{2+}$

Ionic size decreases with increasing positive charge.

Positive charge means increase in effective nuclear charge and hence decrease in size negative charge lead to increase in repulsion between electrons thus increases atomic size

(x) $I^- > Rb^+ > Cl^- > Na^+ > Mg^{2+}$

Ionic radii decreases with increase in positive charge and with decrease in number of shells.

Sol 17: (i) We know that electronegativity increases across a period from left to right and decreases while going down in a group. Therefore F is the most electronegative and Cs is the least electronegative element.

(ii) Electronic configuration of C: [He] $2s^22p^2$

Electronic configuration of B: [He] $2s^22p^1$

(IE), of C > (IE) of B because of increased effective number charge in C compared to B

Electronic configuration of C^+: [He] $2s^22p^1$

Electronic configuration of B^+: [He] $2s^2$

$(IE)_2$ of B > $(IE)_2$ of C as B^+ has a stable electronic configuration due to fully filled s orbital which makes removal of electron difficult.

(iii) The outermost electronic configuration of halogens is of the form ns^2np^5. Thus, by adding one electron, they can attain noble gas configuration, which makes addition of electrons very favorable and hence results in high electron affinity.

(iv) The formation of $F^-(g)$ from F(g) enables fluorine to have a stable noble gas configuration. Hence, this process is exothermic.

For formation of $O^{2-}(g)$ from $O^-(g)$, an extra electron needs to be added to an already negative $O^-(g)$ which will be difficult due to coulombic repulsion between electrons making this process endothermic

(v) Transition and inner transition metals have variable oxidation states because ns and (n–1)d orbitals have very similar energy and there is not a huge jump in the amount of energy required to remove the third electron compared with the first and second (which are removed from ns orbitals)

(vi) Zero group elements (noble gases) are chemically inert as these have a very stable electronic configuration of ns^2np^6 and have no tendency to either accept or donate an electron.

(vii) $(IE)_2$ of an element is always greater than $(IE)_1$ as increased amount of energy is needed to remove an electron from an already positively charged ion due to increased nuclear charge.

(viii) Radius of cation is lesser than atom as removal of an electron leads to increase in effective nuclear charge and hence decrease in size.

Radius of anion is bigger than atom because addition of an electron results in increased coulombic repulsion between electrons causing decrease in effective nuclear charge and hence increase in size.

Sol : 18 (a) IE1 of F will be higher than that of Cl.

Reason: As we go down in a group, atomic radius increases and consequently the effective nuclear charge decreases and thus ionization decreases.

(b) EA of O^- is greater than O.

Reason: The second electron affinity of oxygen is particularly high because the electron is being forced into a small, very electron-dense space.

(c) Ionic radius of Cl^- will be greater than K^+.

Reason: Both are isoelectronic species thus ionic radius decreases with increase in nuclear charge.

Sol 19: (i) F because EA increases across the period and decreases down the group

(ii) Cs because IE decreases down the group and increases across the period.

Sol 20 Covalent radius of K > covalent radius of F

Since, size of K^+ < size of K and size of F^- > size of F

Sol 21: (i) Nomenclature:

Name of more electronegative element is written at the end and 'ide' is sufficed to it

Name of less electronegative element is written first in the formula. Eg. I, Br Iodine Bromine

(ii) If difference is EN > 1.7 ionic bond

If difference is EN < 1.7 covalent bond

(iii) Metallic and non-metallic nature

Low EN: metallic

high EN: nonmetallic

(iv) Percentage ionic character in covalent bonds =

$16(X_A - X_B) + 3.5\,(X_A - X_B)^2$

(v) Bond length

When electronegativity difference increases, bond length decreases

$d_{A-B} = r_A + r_B - 0.09\,|x_A - x_B|$

(vi) Bond strength and stability

It increases with increase in electronegativity difference

Eg. H–F > H–Cl > H–Br > H–I

(vii) Acidic strength of hydrides

It decreases with increase in electronegativity difference

H – I > H – Br > H – Cl > H – F

(ix) Acidic nature of oxides

If $X_0 - X_A \geq 2.3$ basic oxides

$X_0 - X_A < 2.3$ acidic oxides

(x) Nature of hydroxide

In AOH, if $(EN)_A > 1.7$, acidic

$(EN)_A < 1.7$, basic

Sol 22: Atomic size: It is the distance between the centre of the nucleus and the outermost shell containing electrons.

Ionic size: It is the distance between the centre of the nucleus and the outermost shell containing electrons in an ion.

Sol 23: BaO is more basic because $(EN)_{Ba} < (EN)_{Mg}$

Sol 24 This statement is wrong since the electronegativity decreases while going down in a group. Therefore, basic strength should increase.

Sol 25: (i) CsOH: basic ($(EN)_{Cs} < 1.7$)

(ii) IOH: acidic $(EN)_I > 1.7$)

(iii) $Sr(OH)_2$: basic ($(EN)_{Sr} < 1.7$)

(iv) $Se(OH)_2$: acidic ($(EN)_{Se} > 1.7$)

(v) FrOH: basic ($(EN)_{Fr} < 1.7$)

(vi) BrOH: acidic ($(EN)_{Br} > 1.7$)

Sol 26: Acidic nature: $N_2O_5 > SO_3 > CO_2 <> CO > CaO$

Acidic strength decreases with increase in electronegativity difference. Also, CO is neutral and CaO is basic.

Sol 27: (i) Anomalous position of hydrogen.

(ii) Isotopes were placed in same group despite having different properties.

(iii) Isobars are placed in different groups.

(iv) Metals and non-metals both placed together in same group.

(v) Some similar elements placed differently.

(vi) Position of lanthanides and actinides is not properly specified.

Sol 28: (i) The nuclear charge on an atom is equal to the atomic number.

(ii) The properties of elements are the periodic functions of their atomic number.

Sol 29: Electron affinity (EA) is the amount of energy released when a neutral gaseous atom accepts an electron to form gaseous anion

$X(g) + E \rightarrow X^-(g) + EA$

EA increase across the period in periodic table due to increased effective nuclear charge with the exception of elements having stable configurations (due to half or full filled orbitals) for which EA is high than adjacent elements. EA decreases down the group due to decrease in effective nuclear charge.

Sol 30: (i) **Number of shells:** IE decreases with increase in number of shells.

(ii) **Effective Nuclear Charge:** IE increases with increase in effective nuclear charge

(iii) **Shielding effect:** The more will be the shielding effect. Less will be the IE.

(iv) The elements which have half-filled or fully filled orbitals have higher IE due to their stable electronic configuration.

(v) **Penetrative Power:** In any atom, the s orbital is nearer to the nucleus than p, d or f orbitals. Therefore, greater energy is required to remove an electron from s orbital. Decreasing order of IE is:

s > p > d > f

Sol 31: (i) Al, due to its intermediate electronegativity.

(ii) F,due to highest effective nuclear charge and least number of shells.

(iii) Xe, It is a group zero element (a noble gas).

Exercise 2

Single Correct Choice Type

Sol 1: (D) I^- will have the largest ionic radius increase further, we know that ionic radius will down the group and decrease across a period. And for isoelectronic species, the ionic radius will increase with increase in negative charge

Sol 2: (D) In group 13, the stability of +2 oxidation state will increase and stability of +4 oxidation state will decrease down the group due to inert pair effect, which is negligible in Ge and most significant in Pb.

Sol 3: (B) We check the atomic radius of all the elements in the four options.

For option (B), we find that the atomic mass of the middle elements is not equal to the average of atomic mass of other 2 elements.

Sol 4: (C) Lothar Meyer used atomic volume for plotting Lothar Mayer's volume curves.

Sol 5: (B) Majority of low atomic number non metallic elements are gases which are placed at top right hand side.

Sol 6: (D) Since, here s orbital is filled last, it is a s-block element

Sol 7: (D) All the reasons given are correct.

Sol 8: (B) The last electron in each period is filled in np orbital which is the same shell, i.e. same energy level.

Sol 9: (C) This is a consequence of inert pair effect, according to which in heavier p-block elements, the s orbitals become very close to nucleus and thus become relatively inert and are not removed.

Sol 10: (C) p-block elements can display a range of oxidation states due to availability of electrons in both of s and p orbitals and also possibility of attaining noble gas configuration by adding electrons.

Sol 11: (C) We note that order of (IE) is F > N > O > C

Here (IE) of N > (IE) of O because N has a stable half-filled orbital configuration and also IE increases across a period.

But after removal of an electron O will have a stable, half-filled orbital configuration and therefore, its IE will be more than F^+. Hence order O > F > N > C

Sol 12: (C) Electron affinity is higher for Cl due to its low electron density.

$(IE)_2$ of alkali metals shows a jump because they have stable noble gas configuration in +1 oxidation state

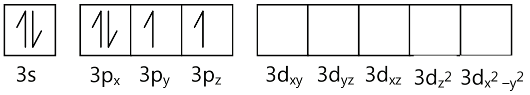
$(IE)_2$ of halogens is zero because halogens have noble gas configuration in –1 oxidation state

He has maximum ionisation energy due to its small size and fully filled orbital configuration.

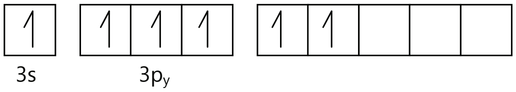
Sol 13: (C) The element isoelectronic with O_2 is sulphur (s).

Its configuration in group state [Ne] $3s^2 3p^4 3d^0$

Outer shell configuration in ground state



Outer shell configuration in second excited state (all electrons get unpaired



Sol 14: (B) Metallic radii of transition elements first decreases due to increase in effective nuclear charge, but after since the electrons are filled in (n – 1)d orbital, after a few elements, the nuclear charge on ns electrons remains almost constant due to extra stability.

Sol 15: (A) If difference in electronegativity > 1.7, ionic bond

Sol 16: (C) In covalent bond, there is actual bond formation therefore they are the strongest of allowed by hydrogen bonding. Vander waal's interactions are least strong because they are developed due to momentary dipole interactions.

Sol 17: (C) The properties of other three pairs resemble due to diagonal relationship between them (identical charge 1 size ratio)

Sol 18: (D) The added electron will have no tendency to go to the higher energy level or sublevel as this will further increase the energy of the system. Alkali metals have no tendency to accept electrons due to low effective nuclear charge.

Atoms with exactly half filled configuration will have low (but not zero) tendency to accept electrons.

Sol 19: (B) Energy required = $IP_1 + IP_2$

$= (178 + 3481\,\text{kcalmol}^{-1}) = 526\,\text{kcalmol}^{-1}$

Sol 20: (C) Configuration $= 1s^2 2s^2 2b^2 3s^2 3p^6 4s^2$

Since last filled shell is 4 and it contains 2 electrons, 4th period and 2nd group

Number of neutrons= Atomic weight – Atomic Number = 40 -20 = 20

For this element this the most stable oxidation state will be +2 as in +2 oxidation state. It will attain a therefore formula of metal oxide = Mo

Sol 21: (C) (Given answer = a =wrong)

For R, after removal of an electron, a fully filled stable configuration is achieved which increases its $(IE)_2$ above all other 3 elements.

Sol 22: (B) Nitrogen has a configuration of $[He]2s^2\ 2p^3$, which is stable because of the half-filled p- orbital and thus, it has more IE than Oxygen.

Sol 23: (D) All the given reasons are correct.

Sol 24: (C) With the removal of an electron, the further removal of electron becomes increasingly difficult due to increase in effective nuclear charge and decrease in size. Hence $I_1 < I_2 < I_3$

Sol 25: (B) Only (B) and (C) order is wrong

F > O, since IP increases across the period

N > O, as N has a stable half filled configuration

P > S, as P has a stable half filled configuration

Be > B, as Be has a stable fully filled configuration.

Sol 26: (A) If difference between $(IP)_1$, and $(IP)_2$ > 16 eV, then element will not form a dipositive ion.

Sol 27: (A) Inconsistency in question, taking x^{-2}, no correct option

No. of electrons in X^{-2} = 18 + 8 = 26

Atomic number of X^0 = 26 – 2 = 24

No. of electrons in Y^{-3} = 18 + 3 = 21

Atomic number of Y = 21 + 3 = 24

Sol 28: (A) Electronegativity increases across a period.

Sol 29: (C) The valency first increases as the number of unpaired electrons increases, but then starts decreasing as electrons begin to get paired.

Sol 30: (A) Electronegativity of carbon is least.

Sol 31: (B) Elements are arranged on the basis of increasing atomic number (i.e. increased nuclear charge)

Sol 32: (B) Metallic character decreases across the period and increases down the group

Sol 33: (A) Due to screening effect of inner electrons, the effective nuclear charge on the valence electrons decreases which lowers the IP.

Sol 34: (C) IE does not vary in a regular manner as IE is higher for elements which have stable half-filled or fully filled configurations compared to their adjacent elements, which violates the general trend that IE increases across a period. All other statements are correct.

Sol 35: (C) Just check the electronic configuration for each elements

In : $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^1$ (iv)

Pd : $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10}$ (iii)

Ca : $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ (i)

Cu : $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$ (ii)

Sol 36: (C) Just check the configuration of each element, taking n = 4

Chromium : $[Ar]\ 3d^5 4s^1$

Copper : $[Ar]\ 3d^{10} 4s^1$

Krypton : $[Ar]\ 3d^{10} 4s^2 4p^6$

Bromine : $[Ar]\ 3d^{10} 4s^2 4p^5$

Sol 37: (A) Based on facts

Helium has highest ionization potential.

Fluorine has highest electronegativity.

Technitium is an artificial element.

Lithium has a very high reducing ability.

Sol 38: (C) Basic character decreases across a period.

Sol 39: (A) atomic radius, and Atomic size, decreases across a period electronegativity increases across a period. Effective nuclear charge increases across a period

Previous Years' Questions

Sol 1: (D) As we go down the group, inertness of ns^2 pair increase. Hence, tendency to exhibit +2 oxidation state increases and that of +4 oxidation state decreases.

Sol 2: (A) First I.E. of N > First I.E. of O.

Sol 3: (D) All the noble gases occupy the peaks of I.E. curve.

Sol 4: (D) Electronegativity increases on going from left to right in a period. Thus electronegativity of F > O > N > Cl.

Sol 5: (D) Nuclear charge per electron is greater in P^{5+}. Therefore its size is smaller.

Sol 6: (C) Element belongs to d-block is unnilhexium $(Unh)_{106}$.

Sol 7: (C) I represents Li

II represents K

III represents Br

IV represents I

V represents He

So, amongst these, II represents most reactive metal and V represents least reactive non-metal.

Sol 8: (D) Positive ions will be smaller than parent atoms.

Sol 9: (E) Assertion is false but reason is true.

More is the electron affinity, greater is the oxidizing character.

Sol 10: (A) Both assertion and reason are true and reason is the correct explanation of assertion.

I.E. of N is more than that of $_xO$ as well as $_yC$.

Sol 11: (C) Assertion is true but reason is false.

N is half-filled ($1s^2\ 2s^2 2p^3$) and therefore more stable and hence energy required to lose electron is greater.

Sol 12: (B) NO^+, C_2^{2-}, CN^- and N_2. All have fourteen electrons.

Sol 13: (A) Both are isoelectronic and have same bond order.

Sol 14: (C) Correct basic strength is

$$NH_3 > PH_3 > AsH_3 > BiH_3$$

Sol 15: (B) Follow the periodic trends

Sol 16: (C) Option itself is the reason

Sol 17: (D) For isoelectronic species higher the $\frac{Z}{e}$ ratio, smaller the ionic radius

$$\frac{Z}{e} \text{ for } O^{2-} = \frac{8}{10} = 0.8$$

$$F^- = \frac{9}{10} = 0.9$$

$$Na^+ = \frac{11}{10} = 1.1$$

$$Mg^{2+} = \frac{12}{10} = 1.2$$

$$Al^{3+} = \frac{13}{10} = 1.3$$

Sol 18: (B) Greater charge and small size of cation cause more polarization and more covalent is that compound.

Sol 19: (D) Across a period metallic strength decreases & down the group it increases.

Sol 20: (C) As the size of central atom increases lone pair bond pair repulsions increases so, bond angle decreases.

Sol 21: (C) $CaCO_3 \rightarrow \underset{\text{Basic}}{CaO} + \underset{\text{Acidic}}{CO_2}$

Sol 22: (C) Order of increasing

ΔH_{IE_1} : Ba < Ca < Se < S < Ar

Ba < Ca ; Se < S : On moving top to bottom in a group, size increases. So ionisation energy decreases.

Ar: Maximum value of ionisation energy, since it is an inert gas.

Sol 23: (B) $Na \rightarrow Na^{+} + e^{-}$ Ist I.E.

$Na^{+} + e^{-} \rightarrow Na$ Electron gain enthalpy of Na^{+}

Because reaction is reverse so then.

$\Delta H_{eg} = -5.1\ eV$.

Sol 24: (C) These are isoelectronic species.

As negative charge increases, ionic radius increases.

JEE Advanced/Boards

Exercise 1

Sol 1: 101: Unnilunium

102: Unnilbium

103: Unniltrium

104: Unnilquadium

105: Unnilpentium

106: Unnilhexium

107: Unnilseptium

108: Unniloctium

109: Unnilennium

Sol 2: (i) $Pb^{4+} < Pb^{2+} < Pb$

Size decreases with increasing positive charge.

(ii) Na < Al < Mg < Si

IP increases across a period.

More for Mg because of its stable half filled orbital configuration.

(iii) NiO < MgO < Sr < Rb_2O < Cs_2O

Basic character increases down a group and decreases across a period.

Less for transition metal oxides

(iv) $Be(OH)_2 < Mg(OH)_2 < Ca(OH)_2 < Ba(OH)_2$

Basic character increases down the group.

(v) $Cl^{-} < Ar < K^{+} < Ca^{2+}$

Same configuration in all.

IE is decided by effective nuclear charge

(vi) As < P < S < Cl

EN decreases down the group and increases across the period

(vii) LiI < LiBr < LiCl

Ionic character decreases with decrease in electronegativity difference.

(viii) $Li^{+} < Na^{+} < K^{+} < Rb^{+} < Cs^{+}$

Mobility of ions increases down the group

(ix) Cs < Rb < K < Na < Li

Size of hydrated radii decreases down the group

(x) $Ba^{2+} < Sr^{2+} < Ca^{2+} < Mg^{2+} < Be^{2+}$

Extent of hydration decreases down the group

Sol 3: Inert gases have an electronic configuration of ns^2np^6 which is an extremely stable configuration due to all fully filled orbitals. Hence, they have no tendency to combine with other atoms and change this configuration, so they are monoatomic.

Sol 4: The third period contains 8 elements as the energy of 3d orbital is more than 4s orbital. 3d orbital is filled during the fourth period making the number of elements in third period equal to 8.

Sol 5: (A) $F^{-} < Cl^{-} < S^{2-} < P^{3-}$

Ionic radius increase with increase in number of shells and negative charge

(B) $Al^{3+} < Mg^{2+} < Na^{+} < F^{-} < O^{2-}$

Ionic radius decrease with increasing positive charge

(C) $Mg^{2+} < Na^{+} < K^{+}$

Ionic radius decrease with increase in positive charge and increase with increase in number of shells.

Sol 6: We have,

Internuclear distance

$d_{H-Cl} = r_H + r_{Cl} - 0.09\,|\,r_H - r_{Cl}\,|$

Now, $r_{Cl} = \frac{10}{2}\text{Å} = 5\ \text{Å}$

$r_H = \frac{2}{2}\text{Å} = 1\ \text{Å}$

$x_{Cl} - x_H = 3 - 2.1 = 0.9$

$\therefore d_{H-Cl} = 5 + 1 - 0.09 \times 0.9 = 5.919$

Hence, internuclear distance = 5.919 Å

Sol 7: We have

$d_{As-Cl} = r_{As} + r_{Cl} - 0.09\,|EN_{Cl} - EN_{As}|$

Here, $d_{As-Cl} = 2.20$ Å

$r_{Cl} = 0.99$Å

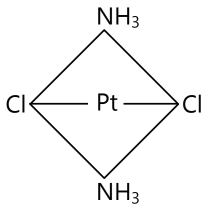
$EN_{Cl} - EN_{e} = 0$

$\therefore r_{As} = d_{As-Cl} - r_{Cl} = 1.21$ Å

Hence, SBCR of As is 1.21 Å

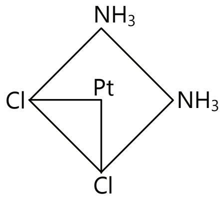
Sol 8: (A) We have Pt–Cl distance = 2.32Å from the figure, we have

Cl–Cl distance = 2 × (Pt –Cl distance) = 4.64 Å



Hence Cl–Cl distance is 4.64 Å

(B) Here, since this is a square planar geometry. Therefore, Cl–Pt–Cl angle will be equal to 90º



Therefore, Cl–Cl distance = $\sqrt{2}$ Pt – Cl distance is 3.28 Å

Hence Cl–Cl distance is 3.28 Å

Sol 9: Since , we know that size of cation is smaller than atomic size and size of anion is larger than atomic size. Therefore atomic radius of K > 1.34 Å and atomic radius of F < 1.34 Å

Sol 10: Noble gases are chemically inert and are therefore monoatomic so, it is not possible to calculate covalent radius of noble gases, instead, for noble gases, Vander waals radius is calculated which is always greater than the radii of precedent halogens (which is covalent radius)

Sol 11: While going from Aluminium to Gallium, 3d electrons are filled which poorly screen the outermost electrons from nuclear charge hence, even though the number of shells increases, an increase in effective nuclear charge will ensure that atomic size of Gallium is equal to the atomic size of Aluminium. Going further down from Ga to In- d-block will be there for all elements.

Hence, atomic radii will increase as usual.

Sol 12: We have $X_p = \dfrac{X_M}{2.8} = \dfrac{IE + EA}{5.6} = \dfrac{4+13}{5.6}$

$\Rightarrow X_p = 3.03$

Sol 13: In isoelectronic species, the increase in atomic radii means an increase in the effective nuclear charge on outermost shell electrons, as the number of electrons are the same. This increase in effective nuclear charge means that that the valence electrons are more firmly bound to the nucleus, hence it results in a decrease in ionic/atomic radii.

Sol 14: Ionisation energy if Cu > Ag is expected as ionization energy decreases down the group. But after Ag the 4f electrons start getting filled and since F orbital is the least penetrative it offers least screening and hence effective nuclear charge increases while going across lanthanoid series. This phenomona is called lanthanoid contraction and due to this $(IE)_{Ag} < (IE)_{Au}$. As Au has more effective nuclear charge than Ag.

Sol 15: Electronic configuration of C = [He] $2s^2 2p^2$

Electronic configuration of B = [He] $2s^2 2p^1$

(IE)1 of C > $(IE)_1$ of B as effective nuclear charge of B is greater than that for C

Electronic configuration of B^+ = [He] $2s^2$

Electronic configuration of C^+ = [He] $2s^2 2p^1$

$(IE)_2$ of B > $(IE)_2$ of C as B have a very stable fully filled orbital configuration which makes removal of electron extremely difficult.

Sol 16: Ionisation potential is the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom $M(g) + IE \rightarrow M^+(g) + e^-$

Sol 17: Electronic configuration of K = [Ar] $4s^1$

Electronic configuration of Cu = [Ar] $3d^{10}4s^1$

We see that on both K and Cu have the same outer shell of $4s^1$ but for Cu the effective nuclear charge will be more than K, as the 3d orbital do not provide very good screening from nuclear charge, therefore, its (IE) is high.

Electronic configuration of K^+ = [Ar]

Electronic configuration of Cu^+ = [Ar] $3d^{10}$

We note that K^+ has a very stable noble gas configuration compared to Cu whose configuration is not so stable. Hence, $(IE)_2$ of K is higher than $(IE)_2$ of Cu.

Sol 18: (i) For Ni, sum of four ionisation energies = 11.29 kJ/mol

For Pt, sum of four ionisation energies = 9.36 kJ/mol

Therefore, Ni can only exhibit +2 oxidation state.

Hence, most stable oxidation state for Ni: +2

Most stable oxidation state for Pt: +2, +4

(ii) Pt can easily form compounds in +4 oxidation state as $(IE)_1 + (IE)_2 + (IE)_3 + (IE)_4$ for Pt is not so high.

Sol 19: Electronic configuration of Mg: [Ne] $3s^2$

Electronic configuration of Na: [Ne] $3s^1$

Mg has a stable fully filled orbital configuration.

Hence (IE), of Mg is high

Electronic configuration of Mg^+: [Ne] $3s^1$

Electronic configuration of Na^+: [Ne]

Na^+ has an extremely stable noble gas configuration which makes removal of electron very difficult.

Hence, $(IE)_2$ of Na is high

Sol 20: Na^+ has higher value of IE because for Na due to removal of an electron, effective nuclear charge is much more compared to Ne, as there are more protons in Na^+ than Ne for same number of p electrons.

Exercise 2

Single Correct Choice Type

Sol 1 (A) X has a stable half-filled orbital configuration which makes removal of electron difficult and increases its IE.

Sol 2: (A) This is the true position of lanthanoids. They are kept below the periodic table to make it more compact.

Sol 3: (B) Similar ionic potentials is the main cause of diagonal relationship.

Sol 4: (C) Will from a diatomic molecule most readily because it has an outer shell configuration of $2s^2 2p^2$ and it needs just 1 electron to attain noble gas configuration, which it can achieve by sharing an electron with another atom to form a diatomic module.

Sol 5: (A) We know,

$$EN = \frac{EA + IP}{2}$$

$$\Rightarrow \; E_1 = \frac{EA + E_2}{2} \Rightarrow EA = 2E_1 - E_2$$

Sol 6: (A) This is ns^2np^4 configuration which is the configuration of ground state of S and Cl^+.

Sol 7: (C) O^- has no natural tendency to accept electrons because of coulombic repulsion between electrons , thus to add an electron to O^-, we have to give it energy from an external source.

Sol 8: (C) The process is

$$Na(s) \xrightarrow{SE} Na(g) \xrightarrow{IE} Na^+(g)$$

Therefore , total energy =SE +IE

Sol 9: (C) The addition of an extra electron to O^- requires addition of energy because of the coulombic repulsion offered by the extra electron to the incoming electron. For other processes, energy is evolved. As far, all of them, EA is positive.

Sol 10: (B) N_2 has no bond polarity and is a very neutral molecule compared to CN^-, which has a net negative charge and therefore will be much more reactive.

Previous Years' Questions

Sol 1: (B) Among isoelectronic species, greater the negative charge, greater the ionic size hence

$F^- < O^{2-} < N^{3-}$

Sol 2: (B) In a group, ionization energy decreases down the group:

Be > Mg > Ca

Sol 3: (B) Be^- is the least stable ion, $Be(1s^2 2s^2)$ has stable electronic configuration, addition of electron decreases stability.

Sol 4: (B, C, D) Incorrect: Electrons are not filled in sub-energy levels s, p, d and f in the same sequence.

(B) **Correct:** No. of valence shell electrons usually determine the stable valency state of an elements.

(C) **Correct:** Physical and chemical properties of elements are periodic function of atomic number which is the basis of modern, long form, of periodic table.

(D) **Correct:** Relative ionicity of the bond between any two elements is function of electronegativity difference of the bonded atoms which in turn has periodic trend in long form of Periodic table.

Sol 5: (A, B) Correct: For greater solubility, hydration energy must be greater than lattice energy.

(B) **Correct:** Greater lattice energy discourage dissolution of salt.

(C) **Incorrect:** When a salt dissolve, energy is required to break the lattice, which comes from hydration process.

(D) **Incorrect:** Explained in (A).

Sol 6: (True) In a group, size increases from top to bottom.

Sol 7: (False) Ionization potential decreases down the group but this is not the only criteria of reducing power.

Sol 8: (i) Mg^{2+}, O^{2-}, Na^{+} and F^{-} are all isoelectronic, has 10 electrons each. Among isoelectronic species, the order of size is cation < neutral< anion.

Also, between cations, higher the charge, smaller the size and between anions, greater the negative charge, larger the size. Therefore,

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

(ii) First ionization energy increases from left to right in a period. However, exception occurs between group 2 and 13 and group 15 and 16 where trend is reversed on the grounds of stability of completely filled and completely half-filled orbitals. Therefore : Ionization energy (1st) : Na < Al < Mg < Si

(iii) If the atoms are from same period, bond length is related directly to atomic radius. Therefore, Bond length : $N_2 < O_2 < F_2 < Cl_2$

Sol 9: Cl^{-}, S^{2-}, Ca^{2+}, Ar

Size : $Ca^{2+} < Ar < Cl^{-} < S^{2-}$

Explained in (i), question 1.

Sol 10: Zn : $3d^{10}\,4s^2$, Cu : $3d^{10}$, $4s^1$

The first ionization energy is greater for Zn but reverse is true for 2nd ionization energy.

Sol 11: $Li^{+} < Al^{3+} < Mg^{2+} < K^{+}$

Size decreases from left to right in a period and it increases from top to bottom in a group. Variation is more pronounced in group than in period.

Sol 12: (A) NO^{-} (16 electron system)

Bond order = 2.

$NO^{\oplus}$, CN^{-} and N_2 are isoelectronic with CO therefore all have same bond order (= 3)

Hence (A) is correct.

Sol 13: (C) The lower oxidation states for the group 14 elements are more stable for the heavier member of the group due to inert pair effect.

Sol 14: Na, F show only one non-zero oxidation state.

Sol 15: (B) The increasing order of atomic radii of the following Group 13 elements is Ga < Al < Tl < In