Answers

Extended or long form of periodic table

1	c	2	d	3	c	4	a	5	c
6	c	7	b	8	b	9	b	10	a
11	d	12	a	13	d	14	d	15	b
16	a	17	d	18	c	19	a	20	b
21	a	22	a	23	b	24	a	25	c
26	a	27	a	28	a	29	b	30	d
31	c	32	c	33	c	34	a	35	b
36	b	37	d	38	c	39	a	40	d
41	a	42	c	43	d	44	d	45	c
46	a	47	b	48	b	49	c	50	c
51	b	52	b	53	c	54	d	55	b
56	c	57	c	58	b	59	a	60	d
61	b	62	b	63	a	64	c	65	a
66	d	67	d	68	a	69	c	70	d
71	c	72	b	73	c	74	c	75	a
76	d	77	a	78	b	79	d	80	c
81	b	82	d	83	b	84	b	85	b
86	d	87	a	88	d	89	c	90	b
91	d	92	b	93	c	94	a	95	d
96	d	97	c	98	b	99	a	100	c
101	c	102	c	103	abcd	104	c	105	d
106	a	107	d	108	a	109	a	110	b
111	b	112	a	113	b	114	a	115	d
116	c	117	c	118	c	119	a	120	c
121	d	122	b	123	a	124	b	125	b
126	b	127	d	128	c	129	b	130	d
131	d								

Atomic and Ionic radii

1	b	2	d	3	c	4	a	5	a
6	d	7	a	8	b	9	b	10	d
11	a	12	d	13	d	14	a	15	a
16	d	17	d	18	d	19	a	20	a
21	b	22	d	23	a	24	a	25	c
26	d	27	c	28	b	29	c	30	c
31	c	32	b	33	d	34	a	35	d
36	a	37	d	38	a	39	b	40	d
41	b	42	a	43	a	44	d	45	a
46	d	47	a	48	c	49	d	50	a
51	b	52	a	53	d	54	a	55	c
56	b	57	c	58	b	59	c	60	b
61	a	62	d	63	c	64	a	65	d
66	d	67	d	68	c	69	c	70	c
71	b	72	a						

Ionisation energy

1	b	2	c	3	d	4	a	5	c
6	a	7	c	8	d	9	a	10	a
11	b	12	c	13	a	14	d	15	b
16	a	17	a	18	c	19	a	20	b
21	d	22	b	23	d	24	c	25	b
26	c	27	b	28	a	29	c	30	b
31	a	32	b	33	a	34	c	35	c
36	c	37	d	38	b	39	c	40	a
41	c	42	b	43	a	44	b	45	b
46	d	47	a	48	b	49	a	50	a
51	a	52	d	53	a	54	a	55	c
56	b	57	a	58	a	59	d	60	a
61	d	62	d	63	a	64	b	65	b
66	c	67	a	68	b	69	a	70	c
71	a	72	b	73	d	74	b	75	d
76	b	77	c	78	c	79	d	80	a
81	d	82	b	83	b	84	a	85	d
86	c	87	a	88	b	89	c		

Electron affinity

1	d	2	c	3	c	4	d	5	a
6	a	7	b	8	b	9	a	10	b
11	c	12	a	13	c	14	b	15	a
16	d	17	b	18	a	19	b	20	c
21	b	22	b	23	d	24	c	25	d
26	a	27	b						

Electronegativity

1	b	2	b	3	d	4	b	5	b
6	a	7	b	8	a	9	c	10	a
11	c	12	b	13	d	14	a	15	c
16	a	17	b	18	c	19	a	20	a
21	d	22	a	23	d	24	b	25	a
26	b	27	c	28	b	29	b	30	a
31	b	32	b	33	a	34	a	35	d

Valency and oxidation state

1	a	2	c	3	c	4	a	5	a
6	d	7	a	8	d	9	b	10	d
11	d	12	c	13	a	14	c	15	d
16	d	17	a	18	c	19	a	20	b
21	a	22	b	23	b	24	b	25	d
26	c	27	b	28	c	29	d	30	b
31	d	32	d	33	b	34	c	35	a
36	b	37	d	38	c	39	d	40	d
41	a	42	d	43	d	44	d	45	b
46	d	47	b	48	c	49	c	50	d
51	d								

Critical Thinking Questions

1	d	2	d	3	d	4	b	5	a
6	b	7	b	8	c	9	a	10	b
11	a	12	a	13	c	14	a	15	d
16	b	17	d	18	c	19	d	20	d
21	b	22	c						

Assertion & Reason

1	d	2	b	3	c	4	c	5	c
6	c	7	d	8	a	9	a	10	b
11	d	12	a	13	c	14	e	15	a
16	c	17	e	18	e	19	b		

Answers and Solutions

Extended or long form of periodic table

2. (d) n/p ratio is a cause of radioactivity.

5. (c) Halogens has 7 electrons in his valance shell (ns^2np^5).

6. (c) As alkali metals have tendency to loose e^-.

7. (b) Each period consists of a series of elements whose atom have the same principal quantum no. (n) of the outer most shell i.e. In second period $n=2$, this shell has four orbitals (one 2s and three $2p$) which can have eight electrons, hence second period contains 8 elements from atomic no. 3 to 10.

9. (b) Neils Bohr developed the long form of periodic table on the basis of Mosley's principle.

10. (a) $33-1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$

11. (d) 16-$1s^2 2s^2 2p^6 3s^2 3p^4$ there are $6e^-$ in outer most shell therefore its group is VIth A.

13. (d) Many metals with catalytic properties because

(i) They provide surface area for reaction to occur

(ii) They decreases the ionisation energy.

(iii) They have vacant d -orbitals.

17. (d) Aluminium. As it belongs to p -block element.

18. (c) $Cu_{29} - [Ar]\ 3d^{10}4s^1$.

21. (a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ principal quantum no. is 4 so it belongs to 4th period.

23. (b) Inert gases, these have ns^2np^6 configuration.

24. (a) $1s^2 2s^2 2p^2 -$ there are $4e^-$ in valence shell therefore it goes to IV- group.

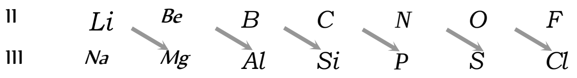
27. (a) $U > Ra > Pb > Hg$

28. (a) *Mg* – *Ba*. Both belongs to II-A group.

29. (b) $Na - Cl$. Both belongs to III period.

30. (d) Elements of second and third period

Diagonal relationship



31. (c) *d*-Block because the last electron enters *d*-subshell.

34. (a) Kr has atomic no. 36 which is a noble gas and all noble gases are included in the *p*-block.

38. (c) *d*-block. As the last e^- enters in *d*-subshell.

40. (d) Due to its vacant *p*-orbital.

41. (a) By obserbing principal quantum number (n), Orbital (s, p, d, f) and equating no. of e^-'s we are able to find the period, block and group of element in periodic table.

42. (c) $33: 1s^2,\ 2s^2,\ 2p^6,\ 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^3$

In its valence shell $5e^-$ are present so it is fifth (A) group element.

43. (d) 38 is the atomic no. of strontium (Sr) which is *s*-block element and all the elements of *s*-block are metals.

44. (d) Hydrogen resembles alkali metals in some properties so it can be placed in the first group of periodic table.

47. (b) Chalcons are oxygen family.

49. (c) Both belongs to VA group.

50. (c) According to Dobernier law of triads the atomic mass of the central element was nearly the arithmetic mean of atomic masses of other two elements.

Cl	*Br*	*I*	Arithmetic mean
31	75	120	$\frac{120+31}{2} = 75.5$

53. (c) Z = 2,8,8,1. ∵ it would donate e^- more easily.

54. (d) Last electron goes to *s*-subshell.

55. (b) Because they belong to same group.

57. (c) Ionic radius will increase as number of shells increases

58. (b) *Al*. Due to diagonal relationship.

60. (d) 2,8,2. ∵ it would donate e^- more easily.

62. (b) A representative element as last e^- enters *p*-orbital.

63. (a) The configuration represents on alkaline earth metals.

65. (a) First group

e.g. $NaCl + H_2O \rightleftharpoons NaOH + HCl$

68. (a) Ionic bond is formed when there is large difference of electronegativities between the atoms.

69. (c) *d*-block - $[Ar]3d^1 4s^2$

70. (d) $Be : 1s^2\, 2s^2$

71. (c) Increasing atomic number. Mosley found that atomic no. was better fundamental property than atomic weight.

72. (b) Lowest ionisation energy due to largest size.

73. (c) Elements on the right side of the periodic table are *p*-block. Mostly non-metals.

74. (c) Screening effect of *d* and *f* block elements is nearly same.

77. (a) Li because of its smallest size.

78. (b) In third group *Na* is a typical element.

85. (b) On equating no. of e^-'s atomic no. is 12 which is for Mg.

86. (d) $17 - 1s^2 2s^2 2p^6 3s^2 3p^5$.

89. (c) Lanthanide's are called rare earth metals.

91. (d) It show similarities with both alkali metals as well as halogens.

92. (b) M^- After gaining an e^- the metal attains stable configuration.

95. (d) Due to presence of vacant *d*-orbitals and they show *d-d* transition.

96. (d) Potassium, $K - [Ar]4s^1$.

97. (c) *p*-block; ${}_{31}Ga \rightarrow [Ar]3d^{10} 4s^2 p^1$.

102. (c) *Mg* has only two electrons in the 3*s*-orbital and hence its I.E. is lowest, i.e. it has the maximum tendency to form di-positive ions.

103. (a,b,c,d) It reflects trends in physical and chemical properties of the elements.

104. (c) As last e^- goes to *d*-subshell.

107. (d) First decreases to a minimum and then increases

108. (a) ${}_{25}Mn - 3d^5 4s^2$.

111. (b) Hydrogen, forms hydrides like halides, e.g. *HCl*.

114. (a) Hydration energy increases along the period.

115. (d) In IIA group all elements are metal while in IIIA, IVA and VIIA groups non-metallic elements are also present.

118. (c) Mg, Ba, Ca have ns^2 configuration.

119. (a) Elements of group halogen are : F, Cl, Br I and At.

121. (d) N and P have 3 unpaired electrons in $2p$ and $3p$ respectively; V has 3 unpaired electrons in $3d$.

124. (b) Tungston (W) having highest m.p.

125. (b) These atomic no. gives the configuration ns^2np^5 which are of halogen group or VIIth group.

126. (b) The atomic no. of an element is derived from the no. of proton because during chemical reaction no. of electron undergoes for change

127. (d) Due to identical ionic radii and polarising power

$$\left[\frac{\text{Charge}}{\text{Size}} \text{ ratio of pairs of these elements}\right]$$

Atomic and ionic radii

1. (b) Value of Z for hydrogen =1
Value of Z for helium = 2
Value of n for both is = 1

$$r_H = \frac{0.52 \times 1^2}{1} \quad r_{He^+} = \frac{0.52 \times 1^2}{1}$$

$$\frac{r_H}{r_{He^+}} = 1 : 1 \text{ or } r_{He^+} : r_H = 1 : 1$$

2. (d) The size of an species decreases with increasing nuclear charge because the attraction for the electrons increases. Thus Al^{3+} is smaller in size.

3. (c) As the nuclear charge per electron is maximum in F^-. Therefore it is smallest in size.

4. (a) During the formation of cation the size decreases.

6. (d) Highest the nuclear charge smallest the atomic size as well as radius also.

7. (a) Atomic radius decreases on going from left to right in a period. Thus size of $O > F$. As O^{2-} and F^- are isoelectronic, therefore, size of $O^{2-} > F^-$.

8. (b) As the nuclear charge per e^- is maximum in Mg^{+2}, it has smallest size among Na^+, Mg^{+2}, Cl^- and F^-.

9. (b) S^{2-} and Cl^- both are isoelectronic but nuclear charge of Cl^- is more than S^{2-}. So it has largest size.

10. (d) In completely filled shell inter atomic repulsion is more so have greater size.

12. (d) I^- as it has the biggest size.

13. (d) *Mg*, as we move across the period atomic radius decreases.

14. (a) O^{-2} has the highest value of ionic radii as this can be explained on the basis of $Z/e \left\{\frac{\text{Nucleaus charge}}{\text{No. of electron}}\right\}$

Whereas Z / e ration increases, the size decreases and when Z / e ration decreases the size increases.

15. (a) Continuous increase as no. of shells increases down the group.

16. (d) $Na^+ < F^- < O^{2-} < N^{3-}$
All are isoelectronic, effective number charge is highest for Na^+ so it has smallest size.

17. (d) $I^- > I > I^+$
$54 \quad 53 \quad 52$ atmoic number

19. (a) Continuously decreases as the effective nuclear charge increases.

20. (a) $Mg^{2+} < Na^+ < F^- < Al$

F^- has bigger size than Mg^{2+} and Na^+ due to small nuclear charge.

21. (b) More than F^- as K^+ has more no of shells in atomic state.

22. (d) All are isoelectronic but O^{2-} has lowest charge among them. So it is largest in size.

23. (a) As effective nuclear charge on Na^+ is maximum. It has smallest size.

25. (c) $Be > C > F > Ne$. Atomic size decreases across a period.

26. (d) As the nuclear charge per electron is maximum in P^{5+}. Therefore its size is smallest.

27. (c) $Na^{+} - 10$ electron ; $Li^{-} -$ 4 electron

28. (b) Ionic radius of trivalent lanthanide's almost remains constant with increase in the atomic number.

30. (c) Halogens are most electronegative elements.

33. (d) On moving from left to right in a period value of radius decreases.

34. (a) $Sc^{3+} > Cr^{3+} > Fe^{3+} > Mn^{3+}$ the correct order is $Cr^{+3} > Mn^{+3} > Fe^{+3} > Sc^{+3}$

35. (d) $Na^{+} > Mg^{2+} > Al^{3+} > Si^{4+}$. All are isoelectronic but nuclear charge per electron is greatest for Si^{4+}. So it has smallest size and nuclear charge per electron for Na^{+} is smallest. So it has largest size.

36. (a) $N^{3-} > O^{2-} > F^{-}$. All are isoelectronic but nuclear charge per electron is highest for F^{-} , so it has smallest size.

38. (a) Cation has small size than parent atom and anion has greater size than parent atom.

39. (b) Ionic radii decreases significantly from left to right in a period among representative elements.

40. (d) H^{-} is most stable due to its full filled 1s-orbital.

43. (a) C^{4-} has largest radius due to least nuclear charge per electron.

44. (d) For ionic bond formation low I.E., high electron affinity and high lattice energy is needed.

45. (a) Ionic radii increases in a group.

46. (d) Size of elements decreases across a period.

47. (a) X^{-} ion larger in size than X atoms. Because of low effective nuclear charge on X^{-}, X has a bigger size.

48. (c) Fe, Co, Ni, Cu . Due to shielding of *d*-electrons, the effect of increased nuclear charge due to increase in atomic no. neutralised. Consequently atomic radius remains almost unchanged after chromium.

49. (d) $S^{2-} > Cl^{-} > K^{+} > Ca^{2+}$

52. (a) Covalent radii decreases on going from left to right in periods. However among the transition elements the size do not changes much because the electrons add to the pneultimate *d*-subshells *i.e.* $(n-1)d$ - subshell.

59. (c) During the conversion of neutral atom to cation size decreases because after removal one e^{-} or more

(i) Nuclear charge per electron increases.

(ii) Outermost shell is completely removed.

60. (b) Atomic radius increases as no. of shells increases.

62. (d) Chloride ion and potassium ion are isoelectronic, isoelectronic ions are those ions having same number of electrons.

$K = 2, 8, 8, 1$ $K^{+} = 2, 8, 8$

$Cl = 2, 8, 7$ $Cl = 2, 8, 8$

63. (c) Cs^{+} has the largest ionic radius in the periodic table.

64. (a) Ionic radii increases down the group.

65. (d) Si^{4+} is smallest in size due to their greater +ve charge.

66. (d) Due to having three electrons atomic size increases.

$F^{-} = 9 + 1 = 10$ electrons ; $O^{-2} = 8 + 2 = 10$ electrons

$Al + 3 = 13 - 3 = 10$ electrons ; $N^{3-} = 7 + 3 = 10\,e^{-}$.

Because electrostatic force between nucleus and $\bar{e}$ cloud is least in nitrogen.

67. (d) The trivalent ion having largest size in lanthanide series is lanthanum. This is due to lanthanide contraction.

68. (c) As we know that hydration power decreases on moving down the group hence among alkali metals *Li* has excessive hydration & hence it has low mobility in aqeous solution.

69. (c) Ionic radius in the nth orbit is given by $r_n = \frac{n^2 a}{z}$ or $r_n \propto \frac{1}{Z}$ where *n* is principal equation no., ao. bohr's radius of hydrogen atom and *Z* is the effective nuclear energy.

70. (c) Order of polarising power $Be^{++} > Li^{+} > Na^{+}$

Hence order of covalent character $BeCl_2 > LiCl > NaCl$

71. (b) Higher the $(n+l)$ value higher is the energy associated with orbitals.

72. (a) With the increase in size of cation the size of the hydrated ion decreases hence ionic conductance increases.

Ionisation energy

1. (b) I.E.(II) of Na is higher than that of Mg because in case of Na , the second e^{-} has to be remove from the noble gas core while in case of Mg removal of second e^{-} gives a noble gas core.

Mg has high first ionisation potential than Na because of its stable ns^2 configuration.

7. (c) Ionization potential decreases. Since, atomic size increases.

8. (d) Alkali metals, lower the no. of valence e^{-} , lower is the value of ionization potential.

9. (a) The ionization energy of hydrogen is to high for group of alkali metals, but too low for halogen group.

13. (a) $E_1 < E_2$ because second I.E. is greater than first I.E.

15. (b) Due to high stability of half-filled orbitals.

16. (a) In Cu it has completely filled *d*-orbital so highest energy is absorbed when it convert in Cu^{+} ion.

18. (c) The energy required to remove an electron from outermost orbit of an isolated gaseous atom is called I.E. Now carbon has $4e^{-}$ in outermost shell. Thus it has 4 ionization energies.

19. (a) Since, stable half filled configuration.

21. (d) First I.P. of $Be > B$ because of stable ns^2 configuration.

22. (b) $K^{+} \rightarrow K^{2+} + e^{-}$. Since e^{-} is to be removed from stable configuration.

24. (c) Since the IV, I.E. is very high. Thus electron is to be removed from stable configuration.

25. (b) Li and Cs belong to 1st group but Cs has larger size, hence low nuclear attraction force, thus low ionization energy.

26. (c) Li belongs to 1st group. There is $1e^{-}$ in outermost shell. Thus low I.E.

27. (b) Increases from left to right. Since, the size decreases.

28. (a) As the e^- is to be removed from stable configuration.

29. (c) Since e^- is to be removed from exactly half filled *p*-orbital.

31. (a) Ionisation potential increases across the period.

32. (b) $E = \frac{E_0}{n^2} = \frac{-54.4}{4} = -13.6\,eV$

34. (c) Due to stable half-filled orbitals.

35. (c) Greater than the first ionization energy because after removal one e^-, effective nuclear charge increases.

36. (c) Rare gases as the e^- is to removed from stable electron configuration.

37. (d) Since it is a noble gas.

38. (b) The first *I.P.* is maximum for hydrogen due to its small size.

41. (c) Due to his fullfilled configuration.

42. (b) 1st I.P. decreases down the group.

43. (a) 1st I.P. increases from left to right in a period.

45. (b) First I.P. for *C* is 11.3, for *N* is 14.5 and for *O* is 13.6

47. (a) *Li* has least I.P about 5.4.

48. (b) I.E. increases across the period.

50. (a) *He* has highest ionisation energy due to it full fill 1*s*-orbital.

51. (a) *s*-electrons are strongly bonded to the nucleus. So large amount of energy is required to remove an e^-.

52. (d) $Mg > Al > Na$. This is due to the presence of fully filled *s*-orbital in *Mg*.

55. (c) The 1st I.P. for hydrogen is 13.6 $volts$

56. (b) Alkali metals are strong reducing agents

58. (a) Due to the large size of group IA elements, the outermost electron is far from the nucleus and can easily be removed. their ionisation energies or ionisation potentials are relatively low.

	Li	*Na*	*K*	*Rb*	*Cs*
Ionisation potential (eV)	5.4	5.1	4.3	4.2	3.9

60. (a) $N > O > Be > B$ 1st ionisation energy of $N > O$ because of half filled p-orbital.

61. (d) $M^{2+} \rightarrow M^{3+}$ After the removal of $2e^-$ the nuclear charge per e^- increases, due to which high energy is required to remove $3e^-$.

63. (a) I.E. increases from left to right in a period.

64. (b) More because of stable configuration of *Mg*.

65. (b) *He* and *Xe* belongs to same group but *He* has higher ionisation energy because of small size.

66. (c) In second transition electron is to be removed from half filled orbital.

68. (b) As it belongs to IA group and has maximum size.

69. (a) Since, they have larger size as compared to other.

70. (c) The second I.E. is greater than first I.E. similarly second E.A. is greater than first E.A. the energy is to be supplied to force the second e^- into the anion.

71. (a) Increases as the atomic size decreases and hence effective nuclear charge increases.

72. (b) *B, Be, C, N* as I.E. increases across the period.

73. (d) Ionization potential is least for alkali metals and it decreases down the group.

74. (b) It has maximum ionization energy due to half filled orbitals.

75. (d) It has maximum no. of e^- in outermost shell. So it has maximum I.E.

76. (b) Ionization potential increases as we go from left to right in a period, while it decreases as we come down a group.

Be	*B*	*Li*	*Na*
9.3	8.3	5.4	5.1

77. (c) Half filled *p*-orbitals possess extra stability.

78. (c) Ionization potential decreases down the group.

79. (d) Li^+ and Mg^{+2} ions have similar polarising power or ionic potential and therefore have similar properties. This type of relationship of the first element of a group with the second element of the next group is known as diagonal relationship.

80. (a) The addition of second electron in an atom or ion is always endothermic.

81. (d) We know that ionisation potential gradually decreases of moving down the group while atomic size increases as we move down the group. Hence larger the atomic size, smaller is ionisation potential.

82. (b) Fluorine has highest E^o red {Equal to +2.9 *V*} due to which it can easily accept an electron & hence it is the best oxidising agent.

83. (b) The ionisation energy of tin {*Sn*} is less than that of lead (*Pb*). It is due to the poor shielding of *d*- and *f*-electron in *Pb* due to which it feels greater attraction from nucleus.

84. (a) The order of screening effect in a given shell are in order $s > p > d > f$.

85. (d) The ionisation energy of Li, Be, B and *C* is 520, 899, 801, 1086 *kJ / mol* respectively hence, carbon has highest IE_1.

86. (c) Isoelectronic species are those which have same no. of electrons.

$K^+ = 19 - 1 = 18$; $Ca^{+2} = 20 - 2 = 18$

$Sc^{+3} = 21 - 3 = 18$; $Cl^- = 17 + 1 = 18$

87. (a) We know that atomic no. of fluorine (*F*), chlorine (*Cl*), Bromine (*Br*) and Iodine (*I*) are 9, 17, 35 and 53 respectively. Therefore correct reactivity of halogens is $F > Cl > Br > I$

88. (b) Ionisation potential generally increases when we more in a period from left to right but IE_1 of N_2 is greater than that of O_2. It is due to the more stable (half-filled orbitals) configurations of *N*.

89. (c) Nitrogen has more ionisation potential than carbon & oxygen because, if outermost orbit is half filled so it is more stable & order is $C < N > O$

Electron affinity

3. (c) $\underset{140.9}{O} > \underset{122.3}{C} > \underset{83}{B} > \underset{0}{N}$ Value of electron affinity increases on going from left to right in periods but the value of electron affinity of Vth A elements is less than that of IVth A element, this is due to half filled *p*-orbitals presence.

4. (d) Halogens have maximum electron affinity due to their smaller size.

5. (a) Zero, because of the stable electronic configuration the noble gases do not show any force of attraction towards the incoming electron.

8. (b) Energy released when an electron is added to an isolated atom in gaseous state.

9. (a) Electron affinity value of *Cl* is greater the *F* and then decreases down the group.

10. (b) Electron affinity increases across the period.

13. (c) Electron affinity of chorine is maximum.

14. (b) The formation of ionic bond depends upon easy formation of cation and anion. therefore the ionisation energy value of the metal atom should be low, so that it can easily form cation. on the other hand, the electron affinity value of the non-metal atom should be high so that it can easily form anion.

15. (a) Because it can easily accept an e^-.

18. (a) Halogens have the highest e^- affinity.

19. (b) In IB group all elements are metals.

22. (b) Flourine although have highest electronegativity due to its very small size, effective inter electronic repulsions are observed which brings down its electron affinity.

23. (d) The bond dissociation energy of *F-F* bond is very low. The weak *F-F* bond makes fluorine the strongest oxidising halogen.

24. (c) Atomic radius increase from top to bottom in a group while decrease from left to right in a period on the other hand electron affinity shows severe trends i.e. decrease from top to bottom in a group and increase from left to right in a period.

25. (d) It is a fact.

26. (a) Electron affinity of *Cl* is greater than fluorine so the order are as $F < Cl > Br > I$

27. (b) Halogens have very high electron affinity. It may be rated that the electron affinity of fluorine is unexpectedly low $(< Cl)$. This may perhaps be due to small size of *F* atom. The value of electron gain enthalpies for Cl, F, S and *O* are respectively 349, 333, 200 & 142 *KJ* / *mol* hence correct order is $Cl > F > S > O$

Electronegativity

2. (b) Decrease as atomic size increases.

4. (b) Electropositive nature increases down the group and decreases across the period.

5. (b) An atom with high electronegativity has high I.P.

6. (a) If electronegativity difference is greater than 1.7 bond is ionic, if less than 1.7, the bond is covalent.

7. (b) Due to decrease in hydration energy of cation and lattice energy remains almost unchanged.

8. (a) *F*, because of its smallest size.

9. (c) Because of small size and high nuclear charge.

10. (a) Electronegativity decreases down the group.

11. (c) Halogens are most electronegative.

12. (b) Electronegativity decreases down the group.

13. (d) Because of smallest size.

14. (a) Electronegativity decreases down the group.

16. (a) Electronegativity increases since the size decreases.

17. (b) Electropositive character decreases across the period as metallic character decreases.

18. (c) *Si, P, S*. As across the period electronegativity increases.

19. (a) Both electronegativity and electron affinity increases. This is because decrease in the size and increase in the nuclear charge. But electronegativity increases continuously.

20. (a) Electropositive nature increases down the group.

21. (d) Electropositive nature increases down the group.

23. (d) The electronegative character increases as the size decreases.

24. (b) Electronegativity increases across a period.

25. (a) $Li_3 - 1s^2 2s^1$ donates $1e^-$ easily.

28. (b) Electronegativity decreases down the group as atomic radius increases.

30. (a) Electronegativity increases across the period because size decreases.

31. (b) Alkali metals are most electropositive and moreover, electropositive character increases down the group.

32. (b) Electronegativity increases when moves towards period & decrease when moves toward group.

33. (a) Electronegativity is the property of a bonded atom. The relative tendency on an atom to attract the shared pair of electron toward itself is called electronegativity.

34. (a) Due to Raving small in size and electron defficient in nature it has highest polarising ability we can use Fazan's rule to understand it further.

35. (d) With decrease in size from *Al* to *S* the basic nature of oxide decrease and acidic nature increases.

$$Al_2O_3 < SiO_2 < P_2O_3 < SO_2$$

Al_2O_3 is amphoteric, SiO_2 is slightly acidic whereas P_2O_3 and SO_2 are the anhydrides of acids H_3PO_3 and H_2SO_3.

Valency and Oxidation state

1. (a) Examples of neutral oxides are CO, H_2O, N_2O. These oxides are neutral towards litmus paper.

5. (a) Basic or metallic character of pentaoxides in VA group increases down the group. Hence acidity decreases.

7. (a) Na_2O, MgO, Al_2O_3, CuO. More the metallic character higher the e^- donating tendency. Therefore lower the I.E. more the basic nature of oxide.

8. (d) As it can donate e^- easily due to low comparative attraction by the nucleus to the valence e^-.

9. (b) Because of the non-metallic character increases.

10. (d) Oxidizing power increases in a group.

14. (c) HF is least acidic due to the small size of fluorine.

16. (d) $Co - [Ar]3d^7 4s^2$, it has 3 unpaired e^- so it is a paramagnetic.

17. (a) Transition elements due to presence of vacant *d*-orbitals.

18. (c) Its valency is 2. So it will form *MO* type compound.

19. (a) Oxides of alkali metals are most basic.

21. (a) Fluorine is the most easily reduced in halogens.

22. (b) Across the period non-metallic character increases. Hence basic nature of oxide decreases.

23. (b) Fluorine is more reactive than chlorine, bromine and iodine.

24. (b) Both are coinage metals

$$3d^{10}\,4s^1 - Cu\ ; \quad 4d^{10}\,5s^1 - Ag$$

26. (c) *Li, Na, K*, contains only one e^- in outer most orbit.

27. (b) Valency is according to valence shall configuration which here is $1s^2, 2s^2, 2p^3$, *ie.* 5

28. (c) *Fe* belongs to first transition series.

29. (d) Reactivity of alkaline earth metals increases down the group.

31. (d) Tendency to gain e^- and oxidising power are related. Among halogens F is the directly most powerful oxidising agent.

32. (d) Electronic configuration of outermost shell of group-17 or halogens are $ns^2\ np^5$.

33. (b) On passing from left to right in a period acidic character of the normal oxides of the element goes on increasing with increase in electronegativity.

36. (b) Gold is found in native state.

37. (d) The elements which having same number of electrons in the valence shell are placed in the same group of periodic table.

38. (c) Alkali metals have the configuration $(n-1)s^2 p^6, ns^1$

41. (a) As going down the group size increases, an liberation of H^+ ion becomes easy. So the order of acidity is :

$$HI > HBr > HCl > HF$$

44. (d) Valence shell configuration for IIA group elements is : ns^2

45. (b) A_2B_3

$A \xrightarrow{-3e^-} A^{+3}$; $B \xrightarrow{+2e^-} B^{-2}$

47. (b) Lower the value of I.P. of an element, the greater will be the basic character of the element.

48. (c) *N*, *O* and *F* have strong tendency to attract the shared pair of electrons i.e. by gaining electrons to form anions.

49. (c) B_2O_3 , Al_2O_3 are amphoteric oxides.

50. (d) He has the atomic number 2 so it does not have octet.

51. (c) Beryllium has the valency of +2 while aluminum exhibits its valency as +3

Critical Thinking Questions

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.

2. (d) Both $Be(OH)_2$ and $Zn(OH)_2$ are amphoteric in nature.

3. (d) CaO is basic; CO_2 is acidic; SiO_2 is weakly acidic. SnO_2 is amphoteric.

4. (b) In $BeCl_2$ has the highest melting point due to ionic bond.

5. (a) Correct order of electron affinity is $Se < S < O$. In a group electron affinity decreases with increase in atomic number.

6. (b) The correct increasing order of I.E. is,

$Cu < Ag < Au$.

7. (b) $B < C < N < O$; When we move from *B* to *O* in a periodic table the first ionisation enthalpy increase due to the attraction of nucleous towords the outer most of electron.

8. (c) Both Fe and Be are metal but Be has stable configuration so it is difficult to release e^- from it. So it has less metallic character than Fe.

9. (a) The basic nature of oxide decreases across the period as metallic character decreases. Therefore acidic nature of oxide increases.

10. (b) Halogens have very high electron affinities. It may be noted that the electron affinity of fluorine is unexpectedly low $(< Cl)$. This may perhaps be due to the small size of the F atom. The values of electron gain enthalpies for *Cl,F, S* and *O* are respectively 349, 333, 200 and 142 *kJ/mole* hence correct order is $Cl>F>S>O$.

11. (a) Increasing order of electronegativity is $Bi<P<S<Cl$.

12. (a) In a group, the ionisation potential decreases from top to bottom. In the alkali group, the ionisation potential decreases from *Li* to *Cs*.

	Li	*Na*	*K*	*Rb*	*Cs*
eV	5.3	5.1	4.3	4.2	3.9

13. (c) Because for removing second electron, it has to be taken out from stable configuration that needs a large amount of energy.

14. (a) First I.E. of $N >$ First I.E. of *O*.

15. (d) All the nobal gases occupy the peaks of I.E. curve.

16. (b) Correct order of ionic size is $Sn > Ce > Lu > Yb$.

17. (d) $1s^2 2s^2 p^6 3s^2$ - In III transition e^- is to be removed from stable configuration.

18. (c) $1s^2 2s^2 2p^6 3s^1$. It belongs to 1A group which has least ionization potential and it decreases down the group

19. (d) Electronegativity increases on going from left to right in a period. Thus electronegativity of $F > O > N \simeq Cl$.

20. (d) Nuclear charge per electron is greater in P^{5+}. Therefore, its size is smaller.

21. (b) The electron affinities of some of the elements of second period (i.e., N, O, F etc.) are however, lower than the corresponding elements (i.e., *P, S, Cl*, etc.) of the third period. This is due to the reason that the elements of second period have the smallest atomic size amongst the elements in their respective groups. As a result, there are considerable electron-electron repulsion within the atom itself and hence the additional electron is not accepted with the same ease as is the case with the remaining elements in the same group.

22. (c) Element belongs to *d*-block is unnilhexium $(Unh)_{106}$.

Assertion and Reason

1. (d) Positive ions will be smaller than parent atoms.

3. (c) Calcium has a higher nuclear charge than sodium.

4. (c) 2*s* orbital has lower energy than 2*p*.

5. (c) *Cl* is more electronegative than *Li*. Although the difference is not much. Therefore the electron pair moves equally to both an thus forming a covalent compound.

6. (c) The lower value of electron affinity of *F* is due to electron-electron repulsion in 2-*p* orbitals of *F*-atom is stronger.

7. (d) All noble gases have stable configuration. Therefore, they can not take any electron means that they have no affinity for electrons. High electron affinity shows that electron is strongly bonded to the atom. Here both assertion and reason are false.

8. (a) The first ionization energy of *Be* is greater than Boron because it is difficult to remove electron from *Be* in comparison to boron. It is also true that the 2P orbitals have lower energy than 2 s-orbitals. Both assertion and reason are true and reason is correct explanation.

9. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

10. (b) Both assertion and reason are true but reason is not the correct explanation of assertion.

Actinoids are more complicated due to the opssibility of large number oxidation states.

11. (d) Both assertion and reason are false.

Ionization enthalpies are always positive. Energy is always absorbed when electrons are removed from an atom.

12. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

The phenomenon is which the Penultimate Shell $(n-1)$ electrons act as screen or shield in between nucleus and valence shell electrons thereby reducing the effective nuclear charge is known as shielding effect.

13. (c) Assertion is true but reason is false.

Ionisation potential decreases with increase in atomic size and also for a given a shell. I.E. is in given order.

$s > p > d > f$

14. (e) Assertion is false but reason is true.

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

15. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

I.E. of *N* is more than that of ${}_8O$ as well as ${}_6C$.

16. (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.

17. (e) Assertion is false but reason is true.

$NO^- = 7 + 8 + 1 = 16\,e^-$ whereas

$CN^- = 6 + 7 + 1 = 14\,e^-$. So both are not isoelectronic.

18. (e) Assertion is false but reason is true.

Outermost electronic configuration of most electropositive elements is ns^1

19. (b) First ionization energy for nitrogen is lower than oxygen due to decrease nuclear charge in nitrogen comparison than oxygen.