Class: XI

Chapter 3: Classification of Elements

Chapter Notes

Top Concepts

1. Johann Dobereiner classified elements in group of three elements called triads.

2. In Dobereiner's triad the atomic weight of the middle element is very close to the arithmetic mean of the other two elements.

3. Dobereiner's relationship is referred as Law of triads.

4. Since Dobereiner's Law of triads worked only for few elements, it was dismissed.

5. Chancourtois arranged elements in order of increasing atomic weights and made a cylindrical table of elements.

6. John Newland arranged the elements in the increasing order of atomic weight and noted that the properties of the every eighth element are similar to the first one. This relationship is called as "Law of octaves"

7. Lothar Meyer proposed that on arranging the elements in order of increasing atomic weights similarities appear at a regular interval in physical and chemical properties.

8. According to Mendeleev's periodic law the physical and chemical properties of elements are periodic functions of their atomic weights.

9. Merits of Mendeleev's periodic table:

   - Mendeleev's periodic table was very helpful in remembering and studying the properties of large number of elements
   - Mendeleev's periodic table helped in correcting the atomic masses of some of the elements like gold, beryllium and platinum based on their positions in the periodic table
   - Mendeleev could predict the properties of some undiscovered elements like scandium, gallium and germanium. By this intuition, he had left gaps for the undiscovered elements while arranging elements in his periodic table.

10. Demerits of Merits of Mendeleev’s periodic table:

- Position of hydrogen is not correctly defined in periodic table. It is placed in group I though it resembles both group 1 and 17.
- In certain pairs of elements increasing order of atomic masses was not obeyed. For example argon (Ar, atomic mass 39.9) is placed before potassium (K, atomic mass 39.1)
- Isotopes were not given separate places in the periodic table although Mendeleev's classification is based on the atomic masses.
- Some similar elements are separated and dissimilar elements are grouped together. For example copper and mercury resembled in their properties but had been placed in different groups. On the other hand lithium and copper were placed together although their properties are quite different.
- Mendeleev did not explain the cause of periodicity among the elements.
- Lanthanoids and actinoids were not given a separated position in the table

11. Moseley performed experiments and studied the frequencies of the X-rays emitted from the elements. With these experiments he concluded that atomic number is more fundamental property of an element than its atomic mass.

12. After Moseley’s experimental results Mendeleev’s periodic law was modified to modern periodic law.

13. According to Modern periodic law the physical and chemical properties of the elements are periodic functions of their atomic numbers.

14. Modern periodic table is also referred to as long form of periodic table

15. Horizontal rows in the periodic table are called periods.

16. Vertical columns in the periodic table are called groups.

17. In the modern periodic table there are 7 periods and 18 groups.

18. The period number corresponds to highest principal quantum number of elements.

19. First period contains 2 elements

20. Second and third period contains 8 elements

21. Fourth and fifth period contains 18 elements

22. Sixth period contains 32 elements

23. In the modern periodic table, 14 elements of both sixth and seventh periods i.e. lanthanoids and actinoids respectively are placed separately at the bottom of the periodic table.

24. Elements with atomic number greater than 92 are called transuranic elements.

25. According to IUPAC, until a new element's discovery is proved and its name is officially recognized it is given a temporary name. This nomenclature is based Latin words for their numbers.

26. The interim names of the newly discovered elements are derived by combining together the roots in order of digits which make up the atomic number and ium is added at the end.

27. Notation for the IUPAC nomenclature of elements

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	p
6	hex	h
7	sept	s
8	oct	o
9	enn	e

28. The distribution of electrons into orbitals of an atom is called its electronic configuration.

29. The electrons in an orbital are filled according to n+l rule.

30. The number of elements in each period is twice the number of atomic orbitals available in the energy level that is being filled.

31. On moving down a group in a periodic table the number of shell increases from 1 to 7

32. Elements in the same group have same number of valence electrons

33. Value of the principal quantum number for the valence or outermost shell gives the period.

34. The first period has principal quantum number n=1, contains two elements and corresponds to K-shell.

35. Since K-shell contains only one orbital (1s) it can accommodate two electrons. Thus there are two elements in K-shell.

36. The second period has principal quantum number n=2, contains eight elements and corresponds to L-shell.

37. The 4 orbitals filled in second period are one 2s (with 2 electrons) and three 2p (6 electrons).

38. The third period has principal quantum number n=3, contains eight elements and corresponds to M-shell.

39. The four orbitals filled in third period are one 3s (2 electrons) and three 3p (6 electrons).

40. The fourth period has principal quantum number n=4, contains eighteen elements.

41. The 9 orbitals filled in fourth period are one 4s (2 electrons), five 3d (with 10 electrons) and three 4p (with 6 electrons).

42. Elements from Scandium (Z=21) to Zinc (Z=30) are called 3d transition series of elements or first transition series.

43. The fifth period has principal quantum number n=5, contains eighteen elements.

44. The nine orbitals filled in fifth period are one 5s (2 electrons), five 4d (10 electrons) and three 5p (6 electrons).

45. Elements from Yttrium (Z=39) to Cadmium (Z=48) are called 4d transition series of elements or second transition series.

46. The sixth period has principal quantum number n=6, contains 32 elements.

47. The 16 orbitals filled in sixth period are one 6s (2 electrons), seven 4f (14 electrons), five 5d (10 electrons) and three 6p (6 electrons).

48. The orbitals filled in seventh period are 7s,5f,6d and 7p

49. Elements from lanthanum (Z=57), Hafnium (Z=72) to mercury (Z=80) are called 5d transition series of elements or third transition series.

50. Fourteen elements from Cerium (Z=58) to Lutetium (Z=71) are called elements of inner transition series or lanthanoid series.

51. Fourteen elements from Thorium (Z=90) to Lawrencium (Z=103) are called elements of 5f inner transition series or actinoid series.

52. The 4f and 5f series of elements are placed separately in periodic table to provide a theoretical justification for periodicity occurring at regular intervals.

53. The modern periodic table is divided into four main blocks – s -block, p-block, d-block and f-block depending on the type of orbital that are being filled with exception of hydrogen and helium.

54. The elements in which last electron enter the s-orbital of their outermost energy level are called s-block elements.

55. The s-block consists of two groups, Group-1 and Group-2.

56. The elements of Group-1 are called alkali metals and have ns^1 as the general outer electronic configuration.

57. The elements of Group-2 are called alkaline earth metals and have ns^2 as the general outer electronic configuration.

58. The elements in which last electron enter the p-orbital of their outermost energy level are called p-block elements.

59. The p-block elements constitute elements belonging to group 13 to 18.

60. Elements of s-block and p-block are collectively called representative element

61. The outermost electronic configuration of p-block elements varies from ns^2np^1 to ns^2np^6

62. Elements of group 18 having ns^2np^6 configuration are called noble gases.

63. Elements of group 17 are called halogens

64. Elements of group 16 are called chalcogens

65. Number of valence electrons in group =Group number -10 for elements belonging to group 13 to 18

66. Elements in which the last electron enters d-orbitals of penultimate energy level constitute d-block elements.

67. Elements of group 3 to 12 in the centre of periodic table constitute the d-block elements

68. General outer electronic configuration of d-block elements is $(n-1)d^{1-10}$ ns^{1-2}

69. d-block elements constitute transition series elements. The name "transition series" is derived from the fact the d-block elements represent transition in character from reactive metals (belonging to group1 and 2 constituting s-block) on one side of the periodic table to non-metals (belonging to group 13 to 18 constituting p-block) on other side of the periodic table .

70. Elements in which last electron enters f-orbitals are called f-block elements

71. Elements of Lanthanoid series have general outer electronic configuration of $4f^{1-14}$ $5d^{0-1}$ $6s^2$

72. Elements of Actinoid series have general outer electronic configuration of $5f^{1-14}$ $6d^{0-1}$ $7s^2$

73. Elements in lanthanoid and actinoid series are called inner transition series.

74. Metals comprise more than 78 % of all known elements and appear on left hand side of periodic table

75. Non-metals are placed on right hand side of periodic table

76. Metal are characterized by having a tendency to loose electron

77. Non-metals are characterised by having tendency to gain electron

78. In general metallic character increases down the group and decreases along period

79. In general non-metallic increases along a period and increases along group

80. Elements showing properties of both metals and non-metals are called metalloids or semi-metals

81. The recurrence of similar properties of elements after certain regular intervals when they are arranged in order of increasing atomic number is called periodicity.

82. The cause of periodicity of properties of elements is due to the repetition of similar electronic configuration of their atoms in the outermost energy shell after certain regular interval.

83. Covalent radius for a homonuclear molecule is defined as one half of the distance between the centres of nuclei of two similar atoms bonded by single covalent bond.

84. For heteronuclear molecule covalent radius may be defined as the distance between the centre of nucleus of atom and mean position of the shared pair of electrons between the bonded atoms.

85. Metallic radius is defined as the one half of the internuclear distance two neighbouring atoms of a metal in a metallic lattice.

86. For simplicity term atomic radius is used for both covalent and metallic radius depending on whether element is non-metal or a metal.

87. Atomic radius decrease with increase in atomic number on going from left to right in a period.

88. Atomic radius of elements increase from top to bottom in a group.

89. van der Waals radius is half of the distance between two similar atoms in separate molecules in a solid.

90. Ionic radius may be defined as the effective distance from the nucleus of the ion upto which it has an influence in the ionic bond.

91. A cation is smaller than the parent atom.

92. An anion is larger than the parent atom.

93. On moving from top to bottom in a group in a periodic table ionic radius increases.

94. On moving from left to right in a period in a periodic table ionic radius decrease.

95. Atoms or ions which contain same number of electrons are called isoelectronic species.

96. In case of isoelectronic cations, the cation with a greater positive charge will have a smaller radius because of greater attraction of electrons to nucleus.

97. In case of isoelectronic anions, the anions with a greater negative charge will have a larger radius because repulsion of electrons will outweigh the nuclear charge.

98. A quantitative measure of the tendency of an element to lose electron is given by ionization enthalpy. It represents the energy required to remove an electron from an isolated gaseous atom in ground state.

99. The energy required to remove second most loosely bound electron is called second ionization energy. The value of second ionization enthalpy is higher than first ionization enthalpy because it is more difficult to remove an electron from a positively charged ion than a neutral atom.

100. The effective nuclear charge experienced by the valence electron in an atom will be less than actual nuclear charge on nucleus because of shielding or screening of valence electron from the nucleus by inner core electrons.

101. On moving from left to right along a period in periodic table ionisation enthalpy increases. On moving along a period successive electrons are added to orbitals in same quantum level and shielding of nuclear charge by inner core of electrons does not increase to an extent to compensate for increased attraction of electron to nucleus. Thus increasing nuclear charge outweighs shielding across a period. Eventually more energy is required to remove outermost electron.

102. On moving from top to bottom in a group in periodic table ionisation enthalpy decreases. On moving down a group successive shells and are added and outermost electron move further away from nucleus. Due to electrons present in inner shells shielding of nuclear charge increases. Thus along a group shielding outweighs increasing nuclear charge. Eventually less energy is required to remove an outermost electron.

103. Among various groups in a periodic table, Group 18 elements have highest ionization enthalpy and because of stable electronic configuration

104. When an electron is added to neutral gaseous atom to convert it into a negative ion, the enthalpy change accompanying the process is called electron gain enthalpy.

105. Group 17 elements have high negative electron gain enthalpy because they can attain a stable electronic configuration as of noble gases by accepting an electron.

106. In general electron gain enthalpy becomes more negative from left to right in a period. The effective nuclear charge increases from left to right across a period. Thus it is easier to add to add an electron to a smaller atom because the added electron would be on an average closer to positively charged nucleus.

107. In general electron gain enthalpy becomes less negative as we go from top to bottom in a group. This is because added electron would be farther away nucleus.

108. Elements like O and F have less negative electron gain enthalpy than the succeeding elements like S and Cl respectively of the same group. This is because in case of O and F electron is added to smaller quantum number (n=2) and suffers greater repulsion from electrons present in that level. In case of succeeding elements like S and Cl electron is added to n=3.The added electron occupies large region of space and experiences less repulsion from electrons in that level.

109. Ability of an atom in a chemical compound to attract shared electrons to itself is called electronegativity.

110. The electrons present in the outermost shell are called valence electrons and these electrons determine the valence of atom.

111. Valence of representative element is usually equal to

- The number of electrons in the valence shell
- 8-(the number of electrons in the valence shell)

112. On moving down a group since the number of valence electrons remains the same, all elements exhibit same valence.

113. Oxidation state of an element in a particular compound gives the charge acquired by its atoms on basis of electronegativity consideration from other atoms in the molecule.

114. It is observed that some elements of second period show similarities with elements of third period present diagonally to each other though belonging to different group. This similarity in properties of elements present diagonally is called diagonal relationship.

115. Lithium is diagonally related to Magnesium, beryllium is diagonally related to aluminium and boron is diagonally related to silicon.

116. The anomalous behavior of first element of s and p block elements of each group as compared to other group members is due to following reasons:

- Small size of atom
- Large charge/radius ratio
- High electronegativity
- Non availability of d-orbitals in their valence shell

117. In second period, the first element of each group has 4 valence orbital (2s and 2p) which are available for bonding. Therefore covalence of first member of each group is only 4.

118. Elements of p-block in the second period displays greater ability to form $p\pi - p\pi$ multiple bond to itself (e.g., C=C, $C \equiv C$, N=N, $N \equiv N$) and to other second period elements (e.g. C=O,C=N, $C \equiv N$,N=O) compared to subsequent members of the same group.

119. In a periodic table there is high chemical reactivity at two extreme ends and lowest in centre. Maximum chemical reactivity at extreme left (alkali metals) is exhibited by the easy loss of electrons forming a cation and at extreme right (among halogens) shown by gain of electrons forming anion.

120. The elements which readily loose electrons act as strong reducing agent.

121. The elements which readily accept electrons acts as strong oxidizing agent.

122. Tendency of an element to lose or gain electrons is also related to metallic or non-metallic character.

123. Elements in extreme left of the periodic table have a tendency to lose electron and become positively charged. Hence they show metallic character.

124. Elements in extreme right of the periodic table have a tendency to gain electron. Hence they show non-metallic character.

125. Metallic character decreases along a period on moving from left to right in a periodic table.

126. Non-metallic character increases along a period on moving from left to right in a periodic table.

127. Since elements at extreme left of periodic table show metallic character, oxides formed by them are basic.

128. Since elements at extreme right of periodic table show non-metallic character, oxides formed by them are acidic.

129. Oxides of elements in centre of periodic table are amphoteric or neutral.

130. Transition metals of 3d series are less electropositive than group 1 and 2 metals. This is because their size is small as compared to group 1 and 2 elements accompanied by higher ionization enthalpy as compared to group 1 and 2 elements.