CHM 101 (GENERAL CHEMISTRY 1)

PERIODIC TABLE AND PERIODICITY

1. **Introduction**: The most important step in developing a periodic classification of elements was taken in 1869, when Mendeleeff studied the relationship between the atomic weights of elements and their properties with special emphasis on their valencies. He concluded that the properties of elements are in periodic dependence on their atomic weights. A conclusion that Newlands (1864) noted but never developed. Proof that one quantitative property of atoms really was periodic was provided by Lothar Meyer in 1870.
2. **The periodic table**: There are 18 groups and 7 periods in the periodic table. There are also diagonal relationships as follows

   Li Be B C N

   Mg Al Si P

   The periodic table also comprises of four blocks, s-block (valence electrons in the s orbital); p-block (valence electrons in the s and p orbitals); d-block (valence electrons in the s and orbitals) and f-block (valence electrons in the s and f orbitals). The s and p blocks are the main group elements, while the d and f blocks are the transition elements.

3. **Writing electronic configurations using the noble gases**: Conventionally, electronic configurations are written in a simplified manner using the noble gas cores of the respective preceding periods. For example the elements in period 4 from K to Br have their electronic configurations using the argon core; Zn= $[Ar]3d^{10}4s^{2}$ etc.
4. **Position of metals and non-metals in the periodic table**: Metals are characterized by several distinctive properties e.g. high electrical and thermal conductivities, metallic lustre, malleability and ductility. Whereas, non-metals usually do not exhibit such behaviour. Like all generalizations in chemistry however, the borderline between the two classes of elements is far from rigid. Perhaps the best distinguishing feature between metals and non-

metals is the tendency of the former to form predominantly basic oxides, whereas, the latter favour predominantly acidic oxides. The greatest proportion of elements of elements are classified as metals – groups 1, 2, 3, d and f block elements. Metallic properties are more pronounced in the lower left-hand corner of the periodic table; non-metallic properties are more obvious in the upper right-hand corner of the table.

5. **Trends across a period and down a group**: There is a gradual change from metallic to non-metallic character in passing from left to right across a particular period:

Na_2O MgO Al_2O_3 $(SiO_2)x$ P_4O_{10} SO_3 Cl_2O_7
← Basic → Amphoteric ← Acidic →
However, metallic character increases down a particular group:

N_2O_3 P_4O_6 As_4O_6 Sb_4O_6 Bi_2O_3
Decreasingly acidic → Amphoteric Essentially basic

6. **The differences between the first and second members of a group**: There are often significant differences between the first and second members of a group in the periodic table. For instance lithium differs more from sodium than does sodium from potassium. The first members form many compounds that the other heavier members don't.
7. **Valencies of elements**: The main group elements generally exhibit their group valency or 8 minus group valencies. Whereas, transition metals exhibit variable valencies corresponding theoretically to the number of valence electrons they possess.
8. **The position of hydrogen**: Hydrogen is the first element in the periodic table and is unique in the sense that the properties of hydrogen cannot be correlated with any of the main groups in the periodic table, and hydrogen is best considered on its own. Hydrogen cannot be placed with the alkali metals because it will not readily lose its only one electron. Though one electron short from the noble gas structure it will not typically form a negative ion by gaining an electron like the halogens (Group 17).
9. **The position of the noble gases:** They occupy Group 18/0, a position that implies zero valency. The heavier members, krypton and xenon form well established chemical compounds (especially fluorides and oxides). Helium

and argon have so far resisted attempts to induce them to participate in chemical reactions. Argon, krypton and xenon are also known to form clathrate compounds. The gases are trapped in cavities in the crystal lattice of other compounds. Though the gases are trapped, they do not form bonds. In an aqueous solution of quinol (1,4-dihydoxybenzene) is crystallized under pressure of 10-40 atmospheres of Ar, Kr or Xe, the gas becomes trapped in cavities of about 4 Å diameter in the β-quinol structure. When the clathrate is dissolved, the hydrogen bonded arrangement of β-quinol breaks down and the noble gas escapes. Moreover, the position they occupy in the periodic table is interesting in that they immediately follow a very reactive non-metal (halogen) and precede a very reactive metal (alkali metal).

10. **Transition elements**: Dense metals, high melting points and display of variable valencies with marked catalytic activities. Many of their salts are coloured both as solids and in solution. There are three rows, 1st, 2nd and 3rd rows. The respective ten groups are called triads or congeners. The f-block consists of the lanthanides and actinides.
11. **Electronegativity**: In 1931, Pauling defined the electronegativity of an atom as the tendency of the atom to attract electrons to itself when combined in a compound. Generally, small atoms attract electrons more strongly than larger ones, and hence small atoms are more electronegative.

Pauling's electronegativity coefficients

Li (1.0)	Be (1.5)	B (2.0)	C (2.5)	N (3.0)	O (3.5)	F (4.0)
Na (0.9)						Cl (3.0)
K (0.8)						Br (2.8)
Rb (0.8)						I (2.5)
Cs (0.7)						

Electronegativity increases across a particular period but decreases down a particular group.

12. **Ionization energy**: The energy required to remove an electron completely from an atom of an element is known as the first ionization energy. The second ionization energy is similarly the energy needed to remove the second electron from a singly charged ion and so on.

$K \rightarrow K^+ + e^-$ 1^{st} I.E = 418.4 kJ/mol

$K^+ \rightarrow K^{2+} + e^-$ 2^{nd} I.E = 3068 kJ/mol

Ionization energies are determined spectroscopically.

13. **Electron Affinity**: The energy released when an extra electron is added to a neutral gaseous atom is termed the electron affinity. Generally, only one electron is added forming a uni-negative ion. Since energy is evolved it has a negative sign. They cannot be determined directly, but are obtained indirectly from the Born-Haber cycle.

$Cl \rightarrow Cl^-$ E.A = -348 kJ/mol