Classification of Elements- The Periodic Table

Dobereiner's Triad

Dobereiner arranged the element in increasing order of atomic masses. He found that the atomic mass of the middle element was approximately equal to the arithmetic mean (average) of the atomic masses of the other two elements of that triad when they are arranged in their increasing order of atomic mass,
e.g., Li, Na, K. The atomic mass of Li is 9 and K is 39. The average of two atomic number is 23 that is the atomic number of sodium.

Newland's octaves

Newland's Law of Octaves states that when Elements are arranged in increasing order of Atomic Mass, the properties of every eighth Element starting from any Element are a repetition of the properties of the starting Element. Law of Octaves was true only for Elements up to Calcium.

Newlands was one of the first to detect a periodic pattern in the properties of the elements and anticipated later developments of the periodic law. Newland states that the elements having greater atomic masses cannot accommodate into octaves and could not be fit into the octave pattern.

Mendeleev's Periodic Table

Mendeleev's Periodic law: Physical ,and chemical properties of elements are a periodic function of their atomic masses. Periodic table is a chart of elements prepared in such a way that elements with similar properties occur in the same vertical column. It has seven horizontal rows and eight vertical columns.

Achievements of Mendeleev's Periodic Table

1. It was grouping of elements according to the chemical properties.
2. The periodic table had gaps for the undiscovered elements.
3. The prediction of properties of undiscovered elements were correct.
4. It helped in systematic study of elements

Limitation of Mendeleev's classification
Mendeleev's periodic table suffered few defects as follows:
The position of hydrogen was not correctly defined.
In some cases, Mendeleev placed elements according to their similarities in properties and not in increasing order of their atomic masses. Thus, the position of these elements was not justified
Isotopes were not given separate places in the periodic table although Mendeleev's classification is based on the atomic masses.
Some similar elements were grouped separately while some dissimilar elements were grouped tog
Mendeleev could not explain the cause of periodicity in the elements.
The position for lanthanides and actinides were not included in this table.

Location of Metals, Non-metals, Metalloids and Noble Gases are as follows:
Elements on the left and middle are Metals
Elements on the right are Non-metals
Metalloids form the narrow stair-step area between metals and non-metals
The last group (18) to the right are Noble gases

Periodic Table Chart

The Periodic Table organizes the elements according to their similar chemical and physical properties. The Table has rows and columns.The vertical columns in the periodic table represent Groups. The horizontal rows in table represents Periods.

Introduction to s-Block Elements

s-Block Elements outer electronic configuration is ns1 and ns2. Group 1 (Alkali Metals) and Group 2 (Alkali Earth Metal) belongs to s-block elements. The given figure shows elements present in the s-block of the periodic table.

Introduction to p-Block Elements

p-Block elements consist of elements belonging to group 13 to group 18. They include metals, non-metals, metalloids and noble gas. Their outermost electronic configuration is ns2np1−6. At the end of each period there is a noble gas element. Outermost electronic configuration of noble gas is ns2np6. Group 17 elements are called halogens. They have one electron less in their outermost orbital. Group 16 elements are called chalcogens.

Introduction to d-Block Elements

d-Block elements consist of elements belonging to group 3 to group 12. They have valence electrons in the d-shell. Their outermost electronic configuration is $(n-1)d1-10ns0-2$. They exhibit transitional behavior between s and p-block elements. Thus, they are even called as Transition Elements.

Introduction to f-Block Elements

Two series of 14 elements, placed at the bottom of periodic table are f-block elements. Their outermost electronic configuration is $(n-2)f1-14(n-1)d0-1ns2$. Last electron added to these elements is filled in f-orbital. They are even called as inner-transition elements. Lanthanide Series starts from $Ce(Z=58)-Lu(Z=71)$ while Actinides series starts from $Th(Z=90)-Lr(Z=103)$. Elements after uranium are called as Transuranium Elements.

ionization enthalpy (ionization potential)
The ionization energy (IE) is qualitatively defined as the amount of energy required to remove the most loosely bound electron, the valence electron, of an isolated gaseous atom to form a cation.

Trend in ionization energy down the group
Ionization energy decreases down the group due increase in the atomic size (addition of new shell).

Trend in Ionization energy
Ionization energies are dependent upon the atomic radius. Since going from right to left on the periodic table, the atomic radius increases, and the ionization energy increases from left to right in the periods and up the groups.
As we move down the group, size increases so the outermost electrons are very far away from the nucleus so the electrons are loosely bounded by the nucleus so it is easy to remove it.
Moving across a period, atomic size decreases, so the outermost electrons are nearer to the nucleus. So more force of attraction holds the electrons so more energy is needed to remove the electrons.

Successive ionization enthalpy
Ionization energy is the energy required to remove an electron from the outermost shell of an isolated gaseous atom. When the first electron or the most loosely bound electron is removed, the amount of energy required is less than the energy required to remove the electron in the next successive shell. This ionization energy goes on increasing with the number of electrons removed. So the number of electrons removed from the successive no of shells and the energy involved is called successive ionization energy.

Shielding effect
Shielding effect can be defined as a reduction in the effective nuclear charge on the electron cloud, due to a difference in the attraction forces of the electrons on the nucleus. It is also referred to as the screening effect or atomic shielding.

Electron affinity
The energy released when an electron is added to a neutral gaseous atom is known as electron affinity. The unit for electron affinity is kilo joules per mole.
It depends mainly on two factors. They are atomic size and nuclear charge.

Electron gain enthalpy
The electron affinity increases across a period while it decreases down a group.
The zero group elements have the lowest electron affinity values. Halogens posses highest electron affinity in the periodic table. In halogens chlorine posses highest electron affinity in the periodic table

Trend in electron gain enthalpy
The electron affinity increases across a period while it decreases down a group.
The zero group elements have the lowest electron affinity values.

Electron Affinity
The amount of energy released when an atom in the gaseous state accepts an electron to form an anion. Factors which affect electron affinity are: Atomic size and Nuclear charge. As atomic radii increases, electron affinity increases. As nuclear charge increases, electron affinity increases. It decreases down a group and increases across a period.

Electronegativity
Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons.

Factors which affect electronegativity
Factors Affecting Electronegativity
Nuclear charge. The higher the nuclear charge, more will be the electronegativity value of an element, since the nucleus will be able to attract or pull more electrons towards itself.
Atomic size
Screening effect or shielding effect.

Trends in electronegativity across a period from left to right
Trend-wise, as one moves from left to right across a period in the periodic table, the electronegativity increases.

Electronegativity
The tendency of an atom to attract electrons to itself when combined in a compound is known as electronegativity. The factors which affect the electronegativity are atomic size and nuclear charge. As atomic radii increases, electronegativity decreases. As nuclear charge increases, electronegativity increases. Electronegativity decreases down a group and increases across a per
Definition
Pauling scale
Linus Pauling was the original scientist to describe the phenomena of electronegativity. The best way to describe his method is to look at a hypothetical molecule that we will call XY. By comparing the measured X-Y bond energy with the theoretical X-Y bond energy (computed as the average of the X-X bond energy and the Y-Y bond energy), we can describe the relative affinities of these two atoms with respect to each other.

Δ Bond Energies = (X-Y)measured - (X-Y)expected

if the electonegativities of X and y are the same, then we would expect the measured bond energy to equal the theoretical (expected) bond energy and therefore the bond energies would be zero. If the electronegativities of these atoms are not the same, we would see a polar molecule where one atom would start to pull electron density toward itself, causing it to become partially negative.By doing some careful experiments and calculations, Pauling came up with a slightly more sophisticated equation for the relative electronegativities of two atoms in a molecule:

EN(X) - EN(Y) = 0.102 (Δ1/2).1

In that equation, the factor 0.102 is simply a conversion factor between kJ and eV to keep the units consistent with bond energies.

Applications of electronegativity
The concept of electronegativity is very useful in predicting metallic, non-metallic characters of elements and polarity of bonds.