Fundamental Particles

- ❖ We consider quarks to be fundamental, because so far we have been unable to **break them apart .**
- ❖ **Quarks** and **electrons** at **least 1000 times smaller than proton.**

Quarks and electrons are considered "fundamental"

Hadron

Hadron is defined as the subatomic particle made of quarks. Hadrons are the heaviest particles. It is composed of two or more quarks that are held strongly by the electromagnetic force.

Baryons

Baryons are massive particles which are made up of three quarks. Example: proton and neutron .

Mesons

Mesons are composed of only two quarks,it is very unstable.

Ionization energy

Ionization energy is the quantity of energy that an isolated, gaseous atom in the ground electronic state must absorb to discharge an electron, resulting in a cation.

$$H(g) \rightarrow H^{+}(g) + e^{-}$$

This energy is usually expressed in kJ/mol, or the amount of energy it takes for all the atoms in a mole to lose one electron each.

When considering an initially neutral atom, expelling the first electron will require less energy than expelling the second, the second will require less energy than the third, and so on. Each successive electron requires more energy to be released. This is because after the first electron is lost, the overall charge of the atom becomes positive, and the negative forces of the electron will be attracted to the positive charge of the newly formed ion. The more electrons that are lost, the more positive this ion will be, the harder it is to separate the electrons from the atom.

In general, the further away an electron is from the nucleus, the easier it is for it to be expelled. In other words, ionization energy is a function of atomic radius; the larger the radius, the smaller the amount of energy required to remove the electron from the outer most orbital. For example, it would be far easier to take electrons away from the larger element of Ca (Calcium) than it would be from one where the electrons are held tighter to the nucleus, like Cl (Chlorine).

In a chemical reaction, understanding ionization energy is important in order to understand the behavior of whether various atoms make covalent or ionic bonds with each other. For instance, the ionization energy of Sodium (alkali metal) is 496KJ/mol whereas Chlorine's first ionization energy is 1251.1 KJ/mol **.**Due to this difference in their ionization energy, when they chemically combine they make an ionic bond. Elements that reside close to each other in the periodic table or elements that do not have much of a difference in ionization energy make polar covalent or covalent bonds. For example, carbon and oxygen make CO_2 (Carbon dioxide) reside close to each other on a periodic table; they, therefore, form a covalent bond. Carbon and chlorine make CCl_4 (Carbon tetrachloride) another molecule that is covalently bonded.

Periodic Table and Trend of Ionization Energies

As described above, 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. Exceptions to this trend is observed for alkaline earth metals (group 2) and nitrogen group elements (group 15). Typically, group 2 elements have ionization energy greater than group 13 elements and group 15 elements have greater ionization energy than group 16 elements. Groups 2 and 15 have completely and half-filled electronic configuration respectively, thus, it requires more energy to remove an electron from completely filled orbitals than incompletely filled orbitals.

Alkali metals (IA group) have small ionization energies, especially when compared to halogens.

Cesium is said to have the lowest ionization energy and Fluorine is said to have the highest ionization energy (with the exception of Helium and Neon).

ncreasing trend of ionization energy in KJ/mol (exception in case of Boron) from left table(8)				
Be 899	B 800	C 1086	N 1402	O 1314

reasing trend of ionization energies (Kj/mol) from top to bottom (Cs is the exception
Li 520
Na 496
K 419
Rb 408
Cs 376
Fr 398

1st, 2nd, and 3rd Ionization Energies

The symbol I1I1 stands for the ***first ionization energy*** (energy required to take away an electron from a neutral atom) and the symbol I2 stands for the ***second ionization energy*** (energy required to take away an electron from an atom with a +1 charge. Each succeeding ionization energy is larger than the preceding energy. This means that I1<I2<I3

See first, second, and third ionization energies of elements/ions in Table 3.

Table 3: Ionization Energies (kJ/mol)					
2	3	4	5	6	7
5250					
7297	11810				
1757	14845	21000			
2426	3659	25020	32820		
2352	4619	6221	37820	47260	
2855	4576	7473	9442	53250	643
3388	5296	7467	10987	13320	713
3375	6045	8408	11020	15160	178
3963	6130	9361	12180	15240	
4563	6913	9541	13350	16600	201
1450	7731	10545	13627	17995	217

Ionization Energy and Electron Affinity--Similar Trend

Both ionization energy and electron affinity have similar trend in the periodic table. For example, just as ionization energy increases along the periods, electron affinity also increases. Likewise, electron affinity decreases from top to bottom due to the same factor, i.e., shielding effect. Halogens can capture an electron easily as compared to elements in the first and second group. This tendency to capture an electron in a gaseous state is termed as ***electronegativity.*** This tendency also determines one of the chemical differences between Non metallic and metallic elements.

Diagram 3: showing increasing trend of electron affinity from left to right **(9)**.

B 27	**C** 123.4	**N -7**	**O** 142.5	**F** 331.4

Diagram 4: showing decreasing pattern of electron affinities of elements from top to bottom

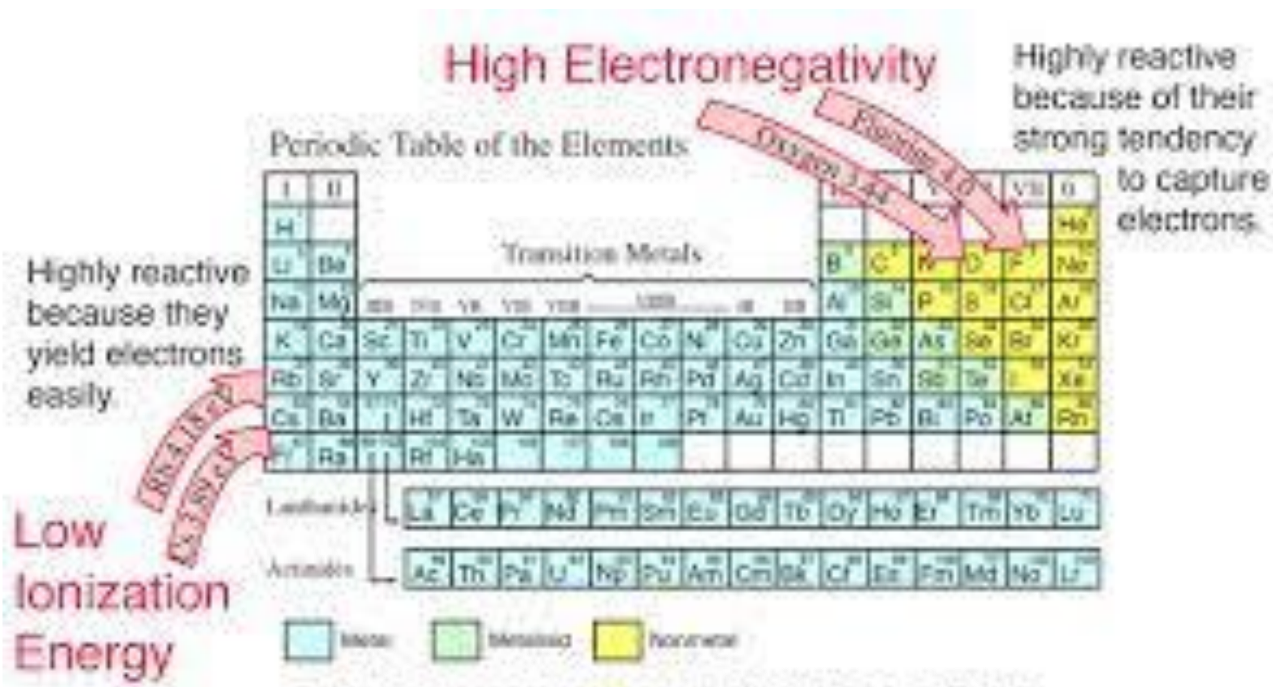
H 73.5
Li 60.4
Na 53.2
K 48.9
Rb 47.4
Cs 46.0
Fr 44.5

As indicated above, the elements to the right side of periodic table (diagram 3) have tendency to receive the electron while the one at the left are more electropositive. Also, from left to right, the metallic characteristics of elements decrease **(4).**

Prediction of Covalent and Ionic Bonds

The difference of electronegativity or ionization energies between two reacting elements determine the fate of the type of bond. For example, there is a big difference of ionization energies and electronegativity between Na and. Cl. Therefore, sodium completely removes the electron from its outermost orbital and chlorine completely accepts the electron, and as a result we have an ***ionic bond*** **(4)**. However, in cases where there is no difference in electronegativity, the sharing of electrons produces a ***covalent bond.*** For example, electronegativity of

Hydrogen is 2.1 and the combination of two Hydrogen atoms will definitely make a covalent bond (by sharing of electrons). The combination of Hydrogen and Fluorine (electronegativity=3.96) will produce a ***polar covalent bond*** because they have small differences between electronegativity .



The highly reactive corners of the Periodic Table

Electron affinity

The **electron affinity** of an atom or molecule is defined as the amount of energy *released* when an electron is attached to a neutral atom or molecule in the gaseous state to form an anion.

$$X(g) + e^{-} \rightarrow X^{-}(g) + \text{energy}$$

Electronegativity

The tendency of an atom in a molecule to attract the shared pair of electrons towards itself is known as **electronegativity**.