IA	IIA	IIIA	IVA	VA	VIA	VIIA	o
H 13.59	*Ionization energy increasing order* →						He 24.58
Li 5.39	Be 9.32	B 8.30	C 11.26	N 14.53	O 13.61	F 17.42	Ne 21.58
Na 5.14	Mg 7.64	Al 5.98	Si 8.15	P 10.48	S 10.36	Cl 13.01	Ar 15.75
K 4.38	Ca 6.11	Ga 6.00	Ge 7.88	As 9.81	Se 9.75	Br 11.84	Kr 14.00
Rb 4.18	Sr 5.69	In 5.78	Sn 7.34	Sb 8.64	Te 9.01	I 10.45	Xe 12.13
Cs 3.89	Ba 5.21	Tl 6.10	Pb 7.41	Bi 7.29	Po 8.43	At	Rn 10.75

It is noted that the ionization energy of oxygen is less than that of nitrogen, despite the increase in ionization energy when moving in one period from left to right. The reason for this is that the ($2P^3$) orbital in nitrogen has three electrons distributed singly (semi-saturated orbitals are more stable). However, in the case of oxygen ($2P^4$), the presence of a fourth electron will be added to the semi-saturated orbital, which reduces stability and reduces the energy required to extract the electron.

- ✓ It is noted that the ionization energy and the effective nucleus charge (Z*) increase in the period (from left to right), which means that there is a direct relationship between them, and at the same time there is an inverse relationship between the square of the period number (n) and the ionization energy, which explains the decrease in ionization energy in the same group when descending from top to bottom. The relationship below shows the relationship between ionization energy, the effective charge of the nucleus and the period number (principal quantum number).

$$IE \propto \frac{(Z^*)^2}{n^2}$$

$$I.E = 13.6\ e.v * \frac{(Z^*)^2}{n^2}$$

Z* = Effective nuclear charge.

n = principle quantum number.

13.6 e.v = The ionization energy of the H atom.

This value (-13.6 ev)represents the energy of the electron in the hydrogen atom, and therefore, to remove the electron from the hydrogen atom, it must be exposed to a potential difference (energy) of 13.6 ev.

Ex: Calculate the first ionization energy or (potential) of $_3$Li

Ans: First we must calculate Effective nuclear charge (Z*)

$_3Li: (1S)^2\ (2S)^1$

S = [2×(0.85)] + [0× (0.35)] = 1.7

Z* = Z – S = 3 – 1.7 = 1.3

So $I.E = 13.6\ e.v * [(Z^*)^2 / n^2]$

$= 13.6\ e.v * [(1.3)^2 / (2)^2]$

= 13.6 e.v * [0.4225] = 5.746 e.v

So, **I.E** of $_3$**Li** in KJ/ mol = [5.746] ***** [96.485]

= 554.4 KJ/ mol

Atom			Ionization energy (eV)			Electron
			I_1	I_2	I_3	affinity E_{ea} (eV)
1	H	$1s^1$	13.60			+0.754
2	He	$1s^2$	24.59	54.51		−0.5
3	Li	$[He]2s^1$	5.320	75.63	122.4	+0.618
4	Be	$[He]2s^2$	9.321	18.21	153.85	≤0
5	B	$[He]2s^22p^1$	8.297	25.15	37.93	+0.277
6	C	$[He]2s^22p^2$	11.257	24.38	47.88	+1.263
7	N	$[He]2s^22p^3$	14.53	29.60	47.44	−0.07
8	O	$[He]2s^22p^4$	13.62	35.11	54.93	+1.461
9	F	$[He]2s^22p^5$	17.42	34.97	62.70	+3.399
10	Ne	$[He]2s^22p^6$	21.56	40.96	63.45	−1.2
11	Na	$[Ne]3s^1$	5.138	47.28	71.63	+0.548
12	Mg	$[Ne]3s^2$	7.642	15.03	80.14	≤0
13	Al	$[Ne]3s^23p^1$	5.984	18.83	28.44	+0.441
14	Si	$[Ne]3s^23p^2$	8.151	16.34	33.49	+1.385
15	P	$[Ne]3s^23p^3$	10.485	19.72	30.18	+0.747
16	S	$[Ne]3s^23p^4$	10.360	23.33	34.83	+2.077
17	Cl	$[Ne]3s^23p^5$	12.966	23.80	39.65	+3.617
18	Ar	$[Ne]3s^23p^6$	15.76	27.62	40.71	−1.0
19	K	$[Ar]4s^1$	4.340	31.62	45.71	+0.502
20	Ca	$[Ar]4s^2$	6.111	11.87	50.89	+0.02

Electron Affinity (EA): is defined as the amount of energy released when an electron is attached to a neutral atom in the gaseous state to form a negative ion.. The gaining of an electron is an exothermic process (releasing energy).

$$\mathbf{X(g) + e^- \rightarrow X^- (g) + Energy}$$

The energy released when an electron is added to a neutral atom is a measure of the strength of the bond between this electron and the atom. The greater the energy released, the more strongly the added electron is bound. Since the atomic radius decreases in the period, this means that the added electron will be closer to the nucleus (i.e. will be more strongly bound) as we move from left to right in the period. Therefore, the electron affinity increases in the period from left to right and decreases in the group from top to bottom.

Note1 :that the electron affinity of metals is low and that of non-metals is high.

Note2: Halogens have a high electron affinity, due to their strong desire to gain an electron to fill their outer shell (i.e. until it is filled with eight electrons) and become more stable.

The following table shows the electron affinity values of the elements represented in kj/mole.

H 72							**He** −48
Li 60	**Be** ≤0	**B** 27	**C** 122	**N** −8	**O** 141 −780	**F** 328	**Ne** −116
Na 53	**Mg** ≤0	**Al** 43	**Si** 134	**P** 72	**S** 200 −492	**Cl** 349	**Ar** −96
K 48	**Ca** 2	**Ga** 29	**Ge** 116	**As** 78	**Se** 195	**Br** 325	**Kr** −96
Rb 47	**Sr** 5	**In** 29	**Sn** 116	**Sb** 103	**Te** 190	**I** 295	**Xe** −77

Electronegativity;- Pauling defined electronegativity as **the power of an atom in a molecule to attract a shared pair of electrons (or electron density) to itself**.

*Other definition

Electronegativity: is a measure of an atom's ability to attract electrons from a neighboring atom to which it is bonded.

Note1: All periodic properties are specific to atoms except for negativity, which is specific to molecules.

Note2: The higher the effective charge of the nucleus of an atom within a molecule, the more this atom shows a desire to attract electrons towards it. Therefore, the electronegativity increases in one period (from left to right) and decreases in the group (from top to bottom).

The table below shows the electronegativity values of a group of representative elements.

H 2.1			
C 2.5	N 3.0	O 3.5	F 4.0
Si 1.8	P 2.1	S 2.5	Cl 3.0
Ge 1.8	As 2.0	Se 2.4	Br 2.8