ALMUSTAQBAL UNIVERSITY

College of Health and Medical Techniques
Medical Laboratory Techniques Department

Stage : First year students
Subject : General Chemistry 1 - Lecture 2

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(Electron configuration and periodic table)

ELECTRON CONFIGURATION

The properties of the elements are determined by the arrangement of electrons in their atoms.

Various elements have different numbers of electrons, these electrons are configured around the nucleus in the atom, in an order called the electron configuration. Therefore, atoms of every element have a unique electron configuration whereby the electrons are ordered

The electrons in an atom are grouped around the nucleus into **shells**, roughly like the layers in an onion, according to the energy of the electrons. The farther a shell is from the nucleus, the larger it is, and the more electrons it can hold.

The first shell (n=1),the one nearest the nucleus, can hold only **2** electrons,

The second shell (n=2) can hold **8**,

The third shell (n=3) can hold **18**,

and The fourth shell (n=4) can hold **32** electrons.

The Number of electrons in a shell = $2n^2$

Within shells, electrons are further grouped into **subshells** of four different types, identified in order of increasing energy by the letters ***s*, *p*, *d*, and *f***.

Here is a summary of the types of orbitals and how many electrons each can contain:

Orbitals and Electron Capacity of the First Four Principle Energy Levels				
Principle energy level (n)	Type of sublevel	Number of orbitals per type	Number of orbitals per level(n^2)	Maximum number of electrons ($2n^2$)
1	s	1	1	2
2	s	1	4	8
	p	3		
3	s	1	9	18
	p	3		
	d	5		
4	s	1	16	32
	p	3		
	d	5		
	f	7		

Example:
How many electrons are present in an atom that has its first and second shells filled and has 4 electrons in its third shell? Specify its atomic number (Z).

Answer:
The first shell of an atom holds **2** electrons in its 1*s* orbital,
and the second shell holds **8** electrons (2 in a 2*S* orbital and 6 in three 2*p* orbitals).
Thus, the atom has a total of 2 + 8 + 4 = **14** electrons.

atomic number = number of protons = number of electrons in the atom.
Then the atomic No. (Z)=14 .

Exercise:
How many electrons are present in an atom in which the 1*S*, 2*S*, and 2*p* sub-shells are filled?

Exercise:
What is the atomic number of the atom having the first , second shells and the 3S subshell filled with electrons ?

Electron Configurations

The exact arrangement of electrons in atom shells and subshells is called the **electron configuration** and can be predicted by applying three rules:

RULE 1. (Aufbau Principle)

This principle shows that secondary energy levels are filled with electrons according to their energy level, from the lowest to the highest, **Electrons occupy the lowest-energy orbitals available, beginning with 1*S* and continuing in the order shown in the following Figure.** Within each shell, the orbital energies increase in the order *S, p, d, f.* they follow this order:

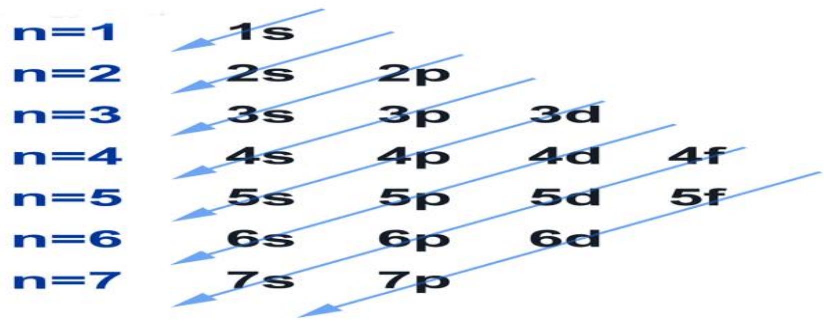


Figure 1: arrangement of the secondary energy levels(orbitals).

Orbital (1S) is filled first with electrons then 2S then 2P then 3S then 3P then 4S then 3d and so far:

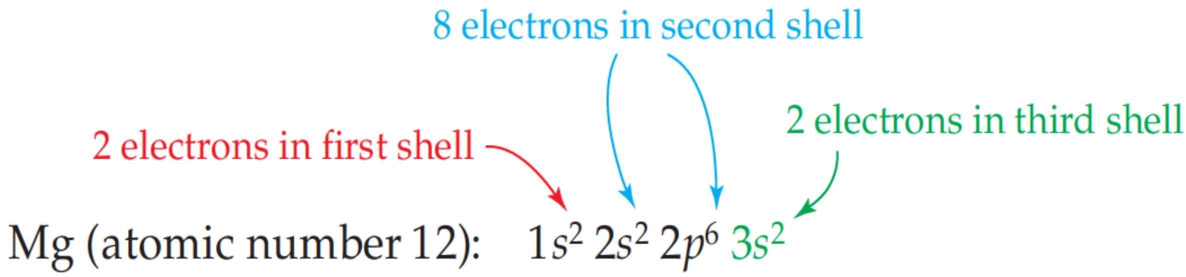
IS 2S 2P 3S 3P 4S 3d 4P 5S 4d 5P 6S 4f ....

RULE 2.

Each orbital can hold only two electrons, which must be of opposite spin.

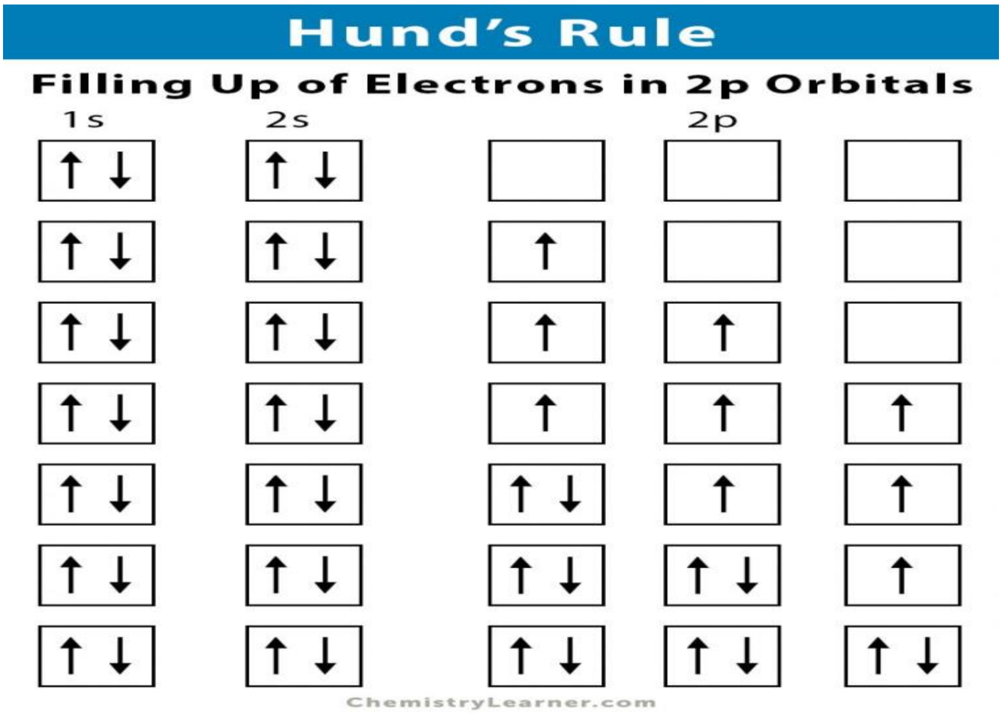
Example:

The electronic configuration of $_{12}Mg$ is as follows:



RULE 3. (Hunds rule)

Two or more orbitals with the same energy(the three *p* orbitals or the five *d* orbitals in a given shell, for example) are each half filled by one electron before any one orbital is completely filled by addition of the second electron.



Example 1:

Write the electron configuration for the elements (${}_{1}H$, ${}_{2}He$, ${}_{3}Li$, ${}_{4}Be$).

Solution:

Elements	Electron configuration	Elements	Electron configuration
${}_{1}H$	$1S^1$	${}_{3}Li$	$1S^2\ 2S^1$
${}_{2}He$	$1S^2$	${}_{4}Be$	$1S^2\ 2S^2$

Example 2:

Write the electron configuration for each of the elements :

(${}_{5}B$, ${}_{8}O$, ${}_{10}Ne$, ${}_{13}Al$ and ${}_{20}Ca$)

${}_{5}B$	$1S^2\ 2S^2\ 2P^1$	${}_{8}O$	$1S^2\ 2S^2\ 2P^4$
${}_{10}Ne$	$1S^2\ 2S^2\ 2P^6$	${}_{13}Al$	$1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^1$

${}_{20}Ca$ $1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^6\ 4S^2$

Exercise 1:

Write the electron configuration for the elements (${}_{9}F$, ${}_{14}Si$, ${}_{15}P$, ${}_{18}Ar$ and ${}_{19}K$)

PERIODIC TABLE

The periodic table is a table that logically organizes all the known elements. Each **element** has a specific location according to its atomic structure. Each row and column has specific characteristics. it is considered the most important tool for those who study chemistry, it is useful in predicting and understanding some properties of elements.

PERIODS:

In the modern periodic table, each horizontal row of the table is called a **period**. Along a period, a gradual change in chemical properties occurs from one element to another. The periodic table consists of seven periods.

GROUPS :

The modern periodic table of the elements contains 18 groups or vertical columns. elements in a group are different but have almost similar chemical properties because they have the same number of outer electrons.

Each column is called a **group** where the elements have the same number of electrons in the outer **orbital**.

Every element in the first column (group one) has one electron in its outer shell. Every element in the second column (group two) has two electrons in the outer shell ...etc..

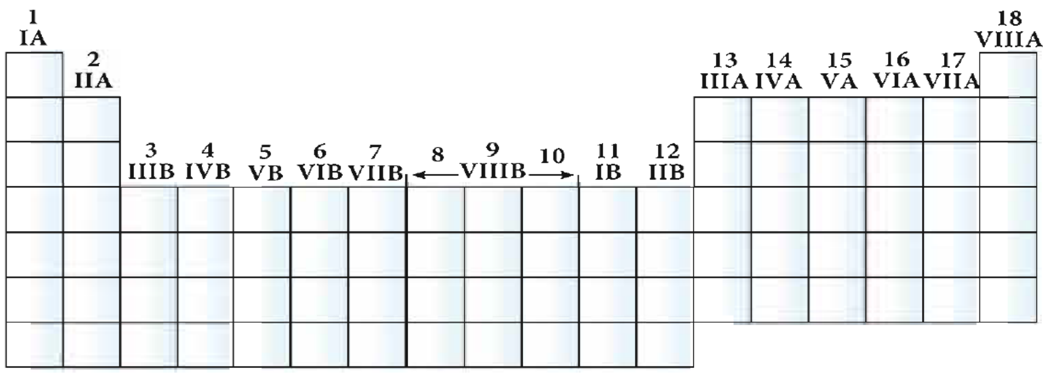


Figure 2 : Periodic table scheme

CLASSIFICATION OF ELEMENTS IN THE PERIODIC TABLE ACCORDING TO ELECTRON CONFIGURATION

In the periodic table Elements can be divided into four blocks, according to the types of the secondary level with which the electron configuration of the elements ends with (s, p, d, f), as illustrated in fig. 3.

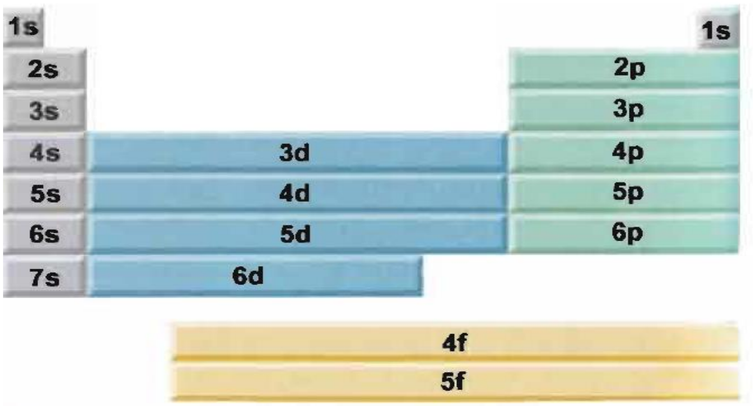


Figure 3: periodic table blocks

1. S- Block Elements*:*
They are elements on the far left of the periodic table including groups IA and IIA, whose electron configuration ends with (S), except for helium (He), which is added to the noble elements at the far right.
Group IA includes elements whose last secondary energy level (S) has one electron, while Group IIA, includes elements whose last secondary energy level (S) has two electrons

2. P-Block Elements*:*
These elements are located on the right side of the periodic table, (see figure 3) whose electron configuration ends with (P) and include six groups, the first five of which are (IIIA , IVA , VA , VIA ,VIIA) and the last group on the far right of the periodic table (group VIIIA), it is called the noble gases group.

3.* d -Block Elements*:
These are metal elements whose electron configuration ends with (S and d) , they are called transition elements or d- block elements, they are located at the center of the periodic table. Figure (3)

4. f -Block Elements*:*

These elements are located at the bottom of the periodic table whose electron configuration ends with (**f**), and called the inner transition elements, including 14 groups belonging to the sixth and seventh periods.

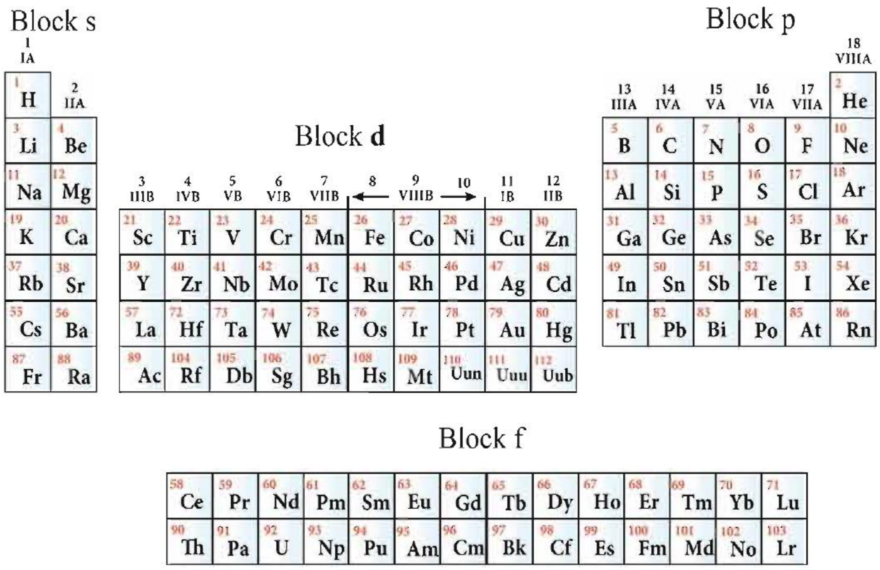


Figure 4: periodic table(atomic number is written on the top left corner in this table).

FINDING PERIOD AND GROUP NUMBER OF ANY ELEMENT

IN GROUP A

To find the period number and group number for group A, the following steps should be followed:

1 - Write the electron configuration of the element.

2- The number of the period is the highest number of the (n), at which the electron configuration of the element ends.

3- The number of the group can be found as follows:

a- If the electron configuration ends with (S), thus the number of electrons in this level is the number of the group.

b- If the electron configuration ends with the (p), thus the number of electrons at this level as well as the secondary level(s) in the primary level which fills before it represents the number of the group. If the total number of electrons is 8 , then it means that this element is in the noble gases group, except for helium, the last energy level of it ends with(S) and contains two electrons only.

Example 3:
What are the period and group for the following elements?
$_{8}O$, $_{10}Ne$, $_{17}C1$, $_{19}K$,

Solution:

$_{8}O = IS^2\ 2S^2\ 2p^4$
The last main level is level (2). Thus, the period is the second period. The last secondary level (p) contains (4) electrons. 2 electrons from (S) are added and the total number is : 2 + 4 = 6 (group six) Oxygen is in the **second period in group 6** A in the periodic table.

$_{10}Ne = 1S^2\ 2S^2\ 2P^6$
The last main level is level (2) so it is **2nd period**.
The last secondary level (p) contains (6) electrons in addition to (2) electrons from the underlying level (2S). The total number is (8). Thus, its group is the **eighth**.
Accordingly, Neon belongs to the **second period in the (VIIIA) group** of the periodic table.

$_{17}C1 =1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^5$
The last main level is level (3). Thus, its period is the **third period**.
Its last secondary level (p) contains (5) electrons in addition to (2) electrons from the underlying level (3S). The total number is (7). Chlorine belongs to **group seven** of the periodic table.
Thus, chlorine is in the **third period of group (7 A)** of the periodic table.

$_{19}K$ $1S^2\ 2S^2\ 2p^6\ 3S^2\ 3p^6\ 4S^1$
The last main level is level (4) and its period is the fourth. The last secondary level (S) contains one electron and its group is the first.
Based on this, potassium belongs to the **fourth period and group 1 A**.

Exercise 2:
What are the period and group for the following elements?
$_{3}Li$, $_{6}C$, $_{13}Al$

Example 4:
What is the common property between the locations of the following elements in the periodic table? $_{3}Li$, $_{11}Na$, $_{12}Mg$.

$_{3}Li = 1S^2\ 2S^1$
group (1 A) / 2nd period

$_{11}Na = 1S^2\ 2S^2\ 2P^6\ 3S^1$
group (1 A) / 3rd period

$_{12}Mg = 1S^2\ 2S^2\ 2P^6\ 3S^2$
group (2A) / ***3rd*** period

According to the above, the common property between Li and Na is that they both have the same group (Group 1 A). The common property between Na and Mg is that they have the same period 3rd period.

Example 5:
What is the common property between the locations of the elements ($_{4}Be$, $_{5}B$, $_{7}N$) in the periodic table ? .

Solution:
$_{4}Be = 1S^2\ 2S^2$ group (2A) / 2nd period
$_{5}B = 1S^2\ 2S^2\ 2P^1$ group (3A) / 2nd period
$_{7}N = 1S^2\ 2S^2\ 2P^3$ group (5 A) / 2nd period

All these elements are in the same period (2nd period).
They differ from each other with respect to groups. Each element belongs to a different group. Beryllium (Be) is in the second group. Boron (B) in the third group and Nitrogen (N) in the fifth group.

Exercise 3:
What is the common property between the locations of the elements ($_{6}\mathbf{C}$, $_{14}\mathbf{Si}$, $_{15}\mathbf{P}$) in the periodic table?

PERIODIC PROPERTIES

The physical and chemical characteristics of the elements in the groups and periods of the periodic table vary according to their atomic radius, ionization energy, electron affinity and electronegativity as arranged below.

1. Electronegativity:

Electronegativity is the ability of an atom to pull electrons towards itself. In many chemical compounds, the negative charge of the bonded electrons is centred near a certain atom. This greatly affects the chemical properties of the compound. **Electronegativity is defined as: "The tendency of an atom to attract bonded electrons towards itself in any chemical compound".**

Fluoride, of all other elements, has the greatest electronegativity.

The electronegativity values increase from left to right and bottom to top in the periodic table excluding the Noble gases. The most electronegative element is Fluorine.

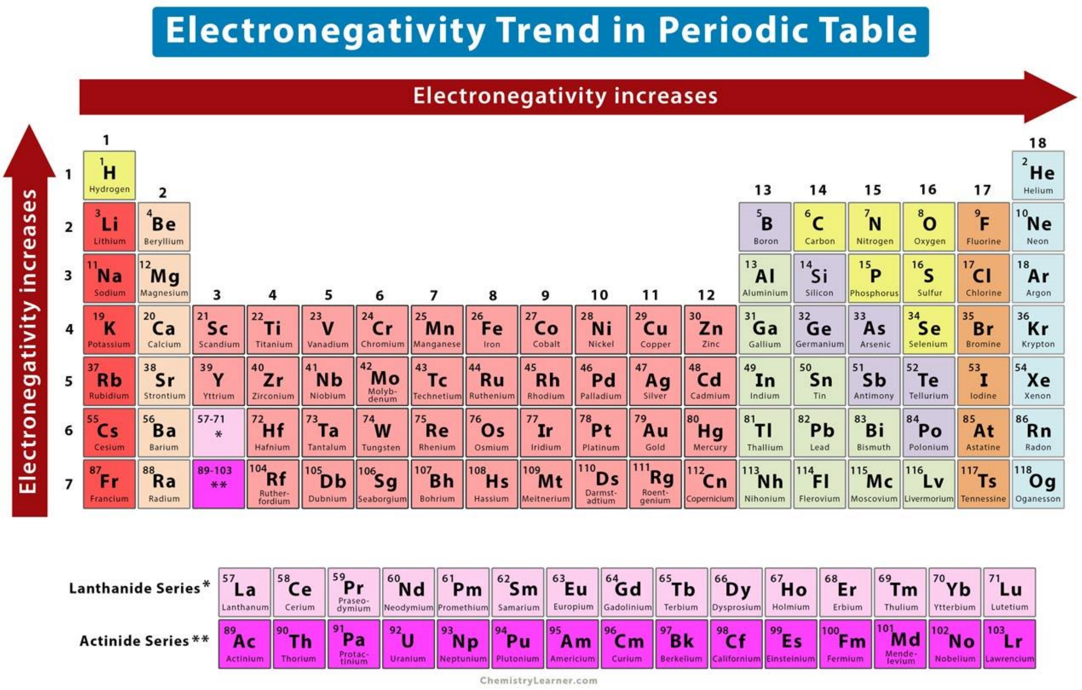


Figure 5: Electronegativity Trend in Periodic Table

From these electronegativity values, we can derive the patterns of two other periodic properties: Ionization Energy and Electron Affinity

2. **Ionization Energy**

Ionization energy is **the amount of energy required to remove an electron from the outer energy level of an atom**. The more electronegative the element, the higher the ionization energy.

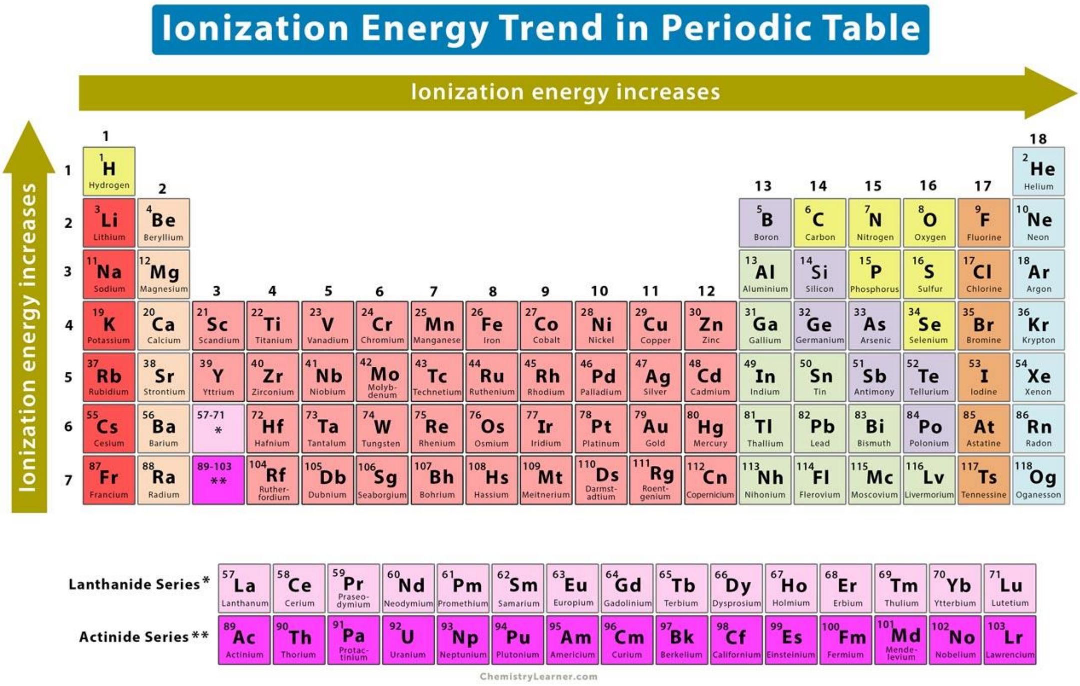


Figure 6: Ionization Energy Trend in Periodic Table

For example: In the ionization of Sodium atom:

Na + ionization energy —> Na^{+} + e

Ionization is not limited to the loss of a single electron from an atom. Two, three, or even more electrons can be lost sequentially from an atom, and the amount of energy associated with each step can be measured.

M + Energy → M^{+} + e- First ionization energy (E_i1)

M^{+} + Energy → M^{2+} + e- Second ionization energy(E_i2)

M^{2+} + Energy → M^{3+} + e- Third ionization energy (E_i3)

Successively larger amounts of energy are required for each ionization step because it is much harder to pull a negatively charged electron away from a positively charged ion than from a neutral atom. Removing the second electron from sodium takes nearly 10 times as much energy as removing the first one.

Third ionization energy (E_i3) > second ionization energy (E_i2) > first ionization energy (E_i1)

3. Electron Affinity (*E*ea)

The Electron Affinity of an element (***E*ea**) is **the amount of energy gained or released with the addition of an electron to the atom to form an anion.**

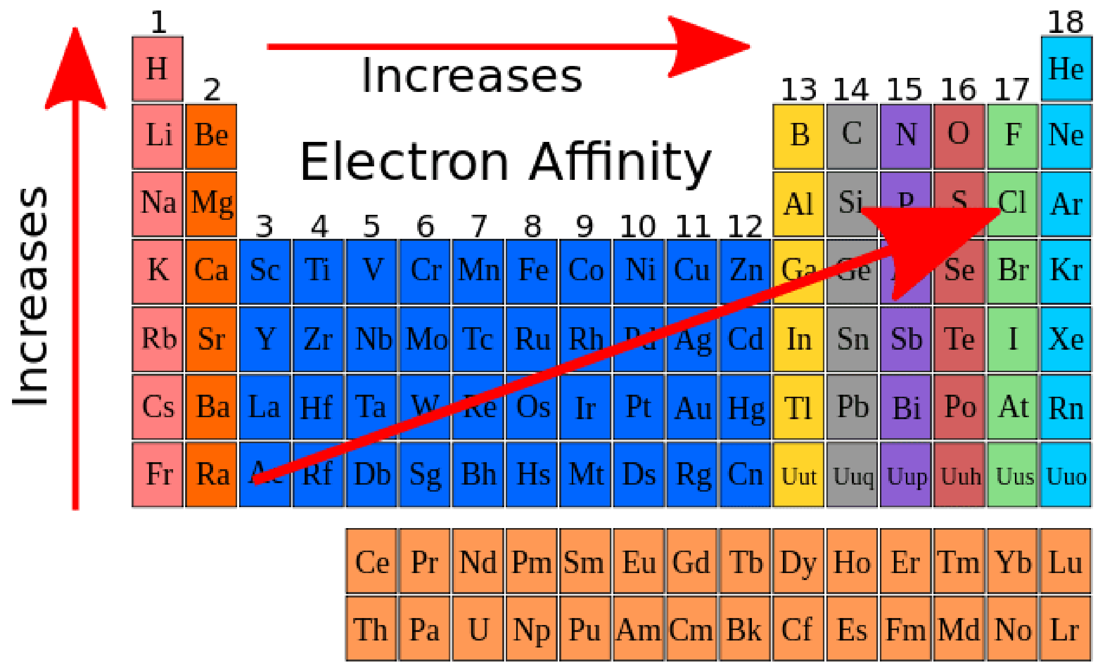
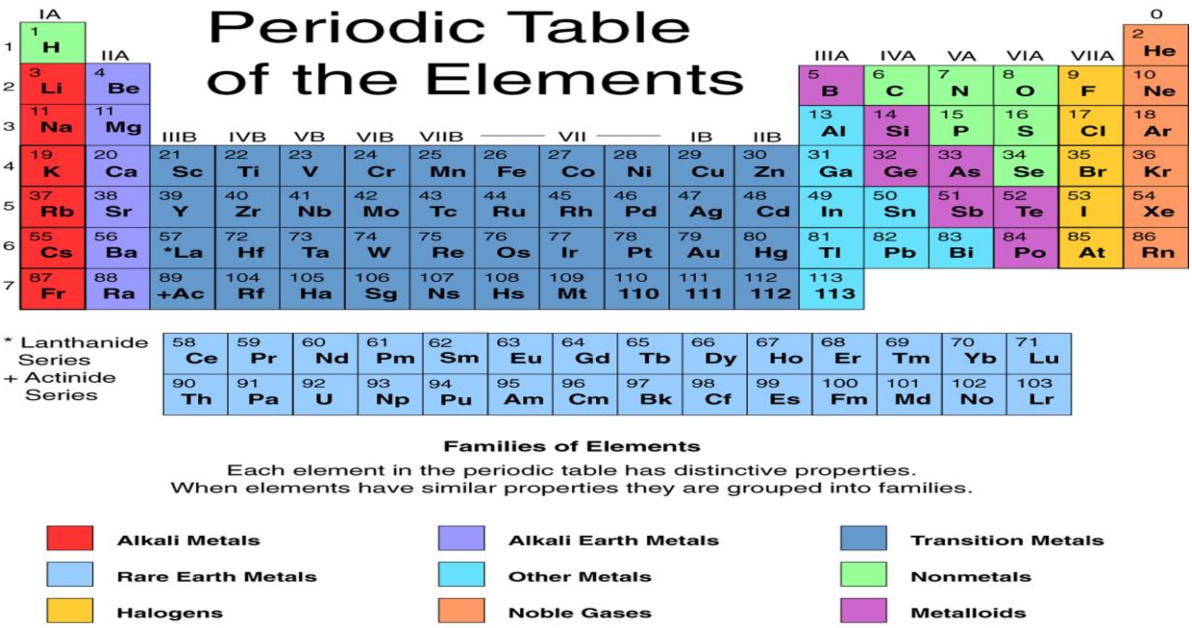
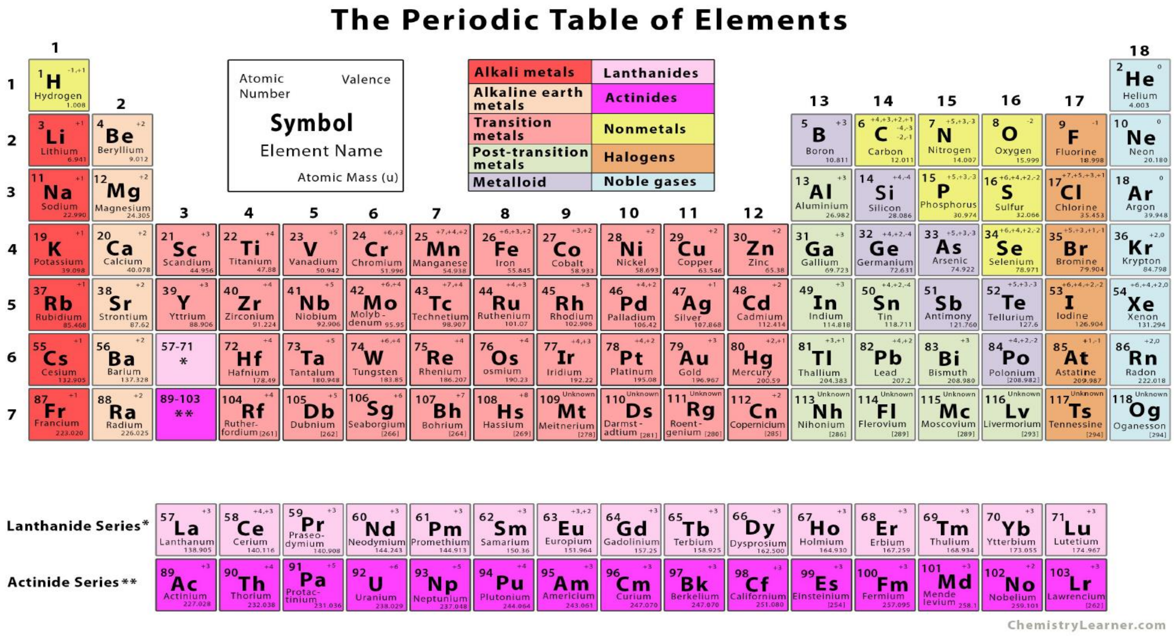


Figure 7: Electron Affinity Trend in Periodic Table

In conclusion:

The electronegativity , ionization energy and Electron Affinity increases in the same pattern in the periodic table. Left to right and bottom to top.


Periodic Table of the Elements
* Lanthanide Series
+ Actinide Series
Families of Elements
Each element in the periodic table has distinctive properties.
When elements have similar properties they are grouped into families.
Alkali Metals
Alkali Earth Metals
Transition Metals
Rare Earth Metals
Other Metals
Nonmetals
Halogens
Noble Gases
Metalloids


The Periodic Table of Elements
Atomic Number
Valence
Symbol
Element Name
Atomic Mass (u)
Alkali metals
Lanthanides
Alkaline earth metals
Actinides
Transition metals
Nonmetals
Post-transition metals
Halogens
Metalloid
Noble gases
Lanthanide Series*
Actinide Series**
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