ALMUSTAQBAL UNIVERSITY COLLEGE

Biomedical Engineering Department

Stage : Second year students

Subject : General chemistry - Lecture 3

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(Electron configuration and periodic table)

ELECTRON CONFIGURATION

Various elements have different numbers of electrons, these electrons are configured around the nucleus in the atom, this order is called the electron configuration. Therefore, atoms of every element have unique electron configuration whereby the electrons are ordered in the atoms in such a way that the total energy is at the minimum, and the following rules are considered when electron are ordered in levels:

Aufbau Principle:

This principle shows that secondary energy levels are filled with electrons according to their energy level ,from the lowest to the highest, they follow this order:

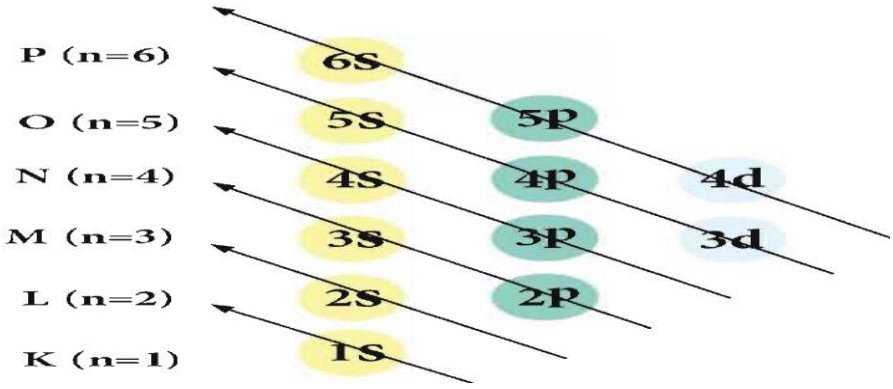

Figure 1: arrangement of the secondary energy levels(orbitals).

While writing the electron configuration for any atom, the atomic number must be known, whereby the number of electrons of the electrically balanced natural atom must be equal to its atomic number, commonly written at left down corner side of the symbol.

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 .....

It is noted that the higher the number of the primary shell, the higher the energy of the electrons and less distance between shells, therefore, there is an overlap between secondary shells which belong to various primary shells. Keep in mind, the number on the left of the symbol of the secondary energy level indicates the primary quantum number (n), while the number on the upper right of the symbol (S) represents the number of electrons in this level this goes to all symbols.

Hund's Rule:

This rule shows that no two electrons are doubly occupied in the sub shell (secondary level) unless its orbitals are singly filled. Two electrons cannot occupy one orbital until assigning one electron to each orbital in the secondary energy level.

Example 1:

Write the electron configuration for the elements ($_4$Be, $_3$Li , $_2$He, $_1$H).

Solution:

Elements Electron configuration

$_1$H	$1S^1$	$_2$He	$1S^2$
$_3$Li	$1S^2\ 2S^1$	$_4$Be	$1S^2\ 2S^2$

Exercise 1:

Write the electron configuration for the elements ($_9$F , $_{14}$Si , $_{18}$Ar).

Example 2:

Write the electron configuration for each of the elements ($_5$B , $_8$O, $_{10}$Ne , $_{l2}$Mg , $_{13}$A1 , $_{15}$P)

$_5$B $1S^2\ 2S^2\ 2P^1$

$_8$O $1S^2\ 2S^2\ 2P^4$

$_{10}$Ne $1S^2\ 2S^2\ 2P^6$

$_{l2}$Mg $1S^2\ 2S^2\ 2P^6\ 3S^2$

$_{13}$A1 $1S^2\ 2S^2\ 2P^6\ 3S^3$

$_{15}$P $1S^2\ 2S^2\ 2P^6\ 3S^5$

PERIODIC TABLE

The periodic table is a table that logically organize 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. Changes in the properties occur because the number of protons and electrons increases from left to right across a period or row. The increase in number of electrons is important because the outer electrons determine the element's chemical properties. 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. For instance, the noble gasses have their outermost orbit filled and therefore atoms from this family do not bond with other atoms.

Each column is called a **group** where the elements have the same number of electrons in the outer **orbital**. Those outer electrons are also called **valence electrons**. They are the electrons involved in chemical bonds with other elements.

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..

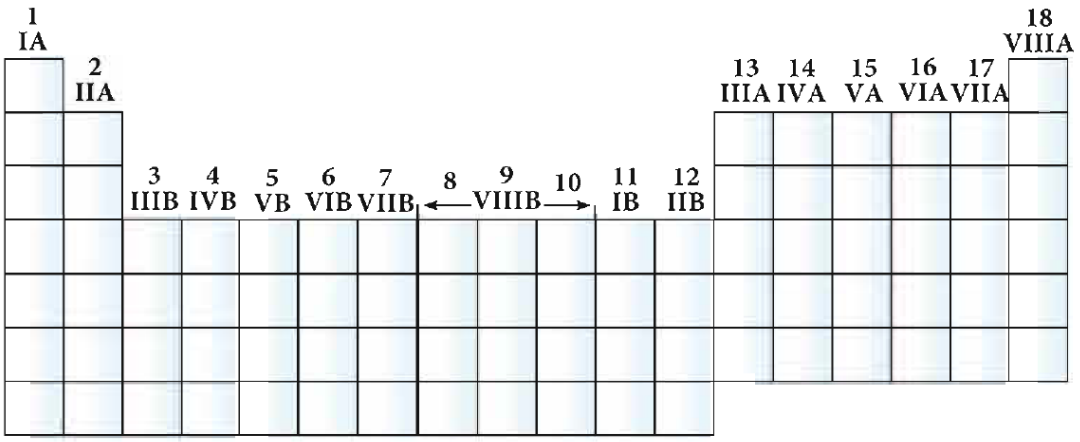

Figure 2 **:** Periodic table scheme

CLASSIFICATION OF ELEMENTS IN THE PERIODIC TABLE ACCORDING TO ELECTRON CONFIGURATION

Electrons play an important part in determining physical and chemical properties of an element, especially those electrons in the outer energy levels, known as valence electrons,

Classification of elements

In the periodic table, depends on these valence electrons. 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.

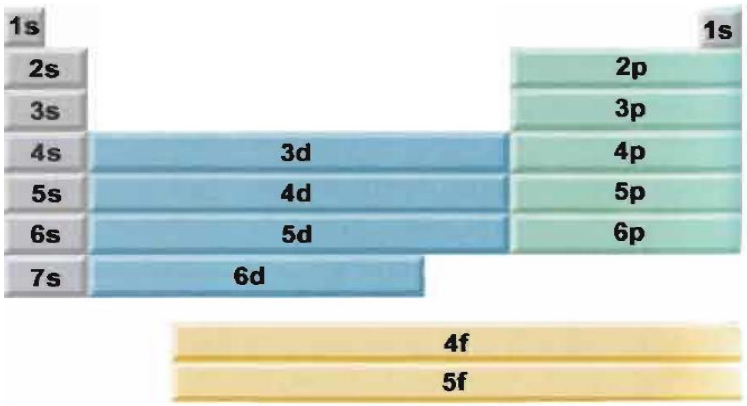

Figure 3: periodic table blocks

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) ,it is added to the noble elements at the far right.

Group IA includes elements whose last secondary energy level (S) have one electron, for Group IIA, it includes elements whose last secondary energy level (S)have 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 or group zero), it is called the noble gases group.

Elements partly filled with electrons at the secondary shells (S and P), and noble elements are called (represtative elements), other names are used for other groups like alkaline metals (IA), as for group IIA it is called alkaline earth metals,(VIIA) are called halogens.

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, 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 sixth and seventh periods.

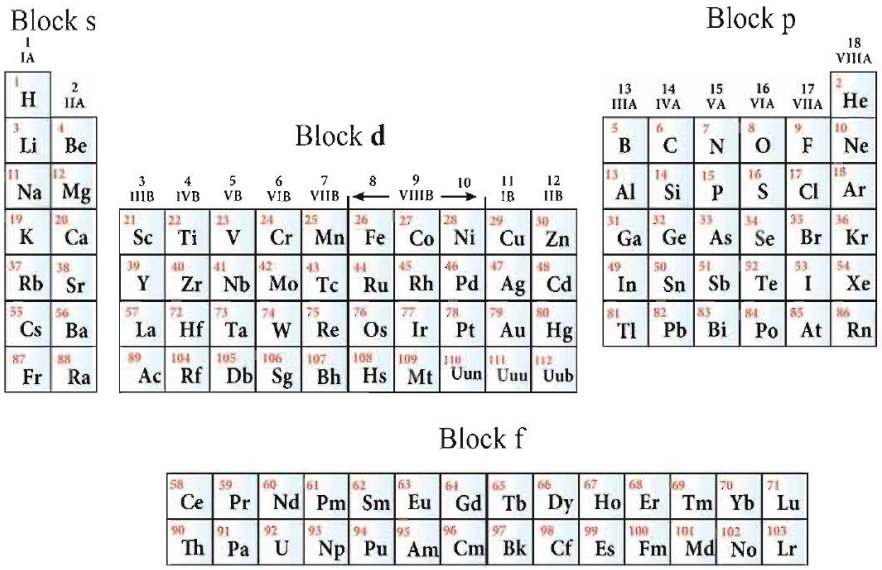

Figure 4: periodic table(atomic number is written on the top left corner in this table).

FINDING PERIOD AND GROUP NUMBER OF ANY ELEMENT

IN GROUP A

To find number 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) , 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 contain two electrons only.

Example 3:

What are the period and group for the following elements?

$_{19}$K , $_{10}$Ne , $_{17}$C1 , $_{8}$O

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 in 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) of 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 (zero) group or (VIIIA) group 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.

Exercise 2:

What are the period and group for the following elements?

$_{13}$ AI , $_{6}$C , $_{3}$Li

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 **($_{15}P$, $_{6}C$, $_{14}Si$)** 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. atomic radius*:*
The radius of the atom determines its volume. Theoretically, atom radius can be calculated by the last occupied level of electrons.

Period:
"It is noticed that the radius of the elements in one period decreases as we move from left to right, as their atomic numbers increase. The attraction energy between the electrons within one main level and the positive charge of the nucleus increases with increasing in the number of electrons in it.

Group:
The radius of the elements in the group, on the other hand, increases as we move from top to bottom in the periodic table and as the outer electrons keep distance from the nucleus, as in the figure below

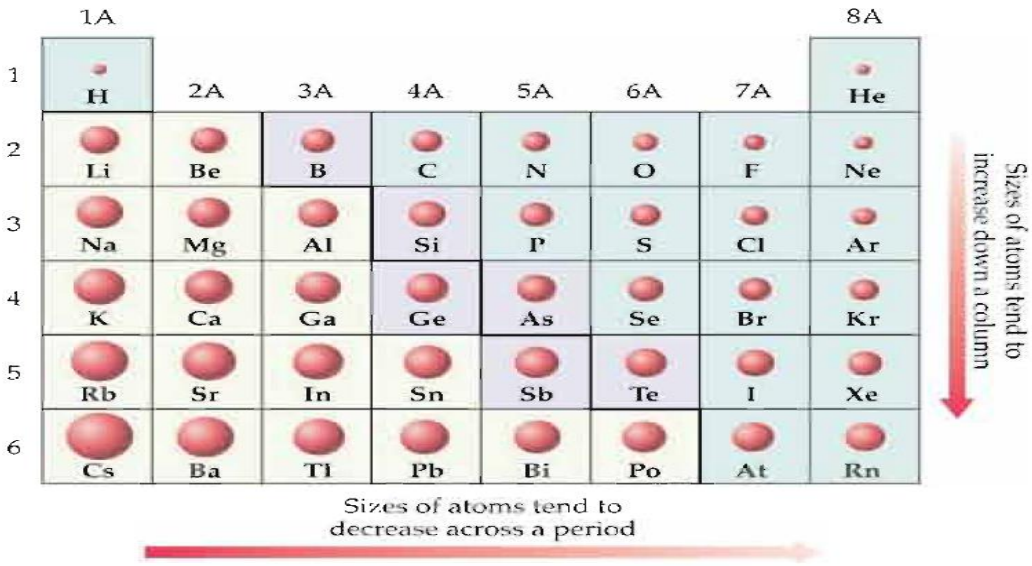

Figure 5: atomic Radius of some elements

Example 6:
Arrange the elements ($_{9}F$, $_{6}C$, $_{8}O$, $_{3}Li$) according to the increase in the atomic radius.

Solution:

$_{3}Li = 1S^2\ 2S^1$
$_{6}C = 1S^2\ 2S^2\ 2P^2$
$_{8}O = 1S^2\ 2S^2\ 2P^4$
$_{9}F = 1S^2\ 2S^2\ 2P^5$

Notice that all the elements above end with the second main level. This Means that they are all in second period of the periodic table. Thus, the arrangement of these elements according to the increase in their radius is as follow: (atomic radius decreases in the period (→) on increase of atomic number)

$_{3}Li > _{6}C > _{8}O > _{9}F$

Exercise 4

Arrange the elements ($_{20}Ca$, $_{12}M g$, $_{4}Be$) according to the increase in their atomic radius.

2. Ionization Energy:

Ionization Energy is defined as:" The amount of energy required to remove one electron from the outer energy level of a gaseous atom." As in the ionization of Sodium atom:

Na + ionization energy —> Na^{+} + e

In the group:

Ionization energies are arranged in the groups from top to bottom. As the atomic number becomes greater, the ionization energy of an element decreases. The reason behind this is that the outer shells' electrons stay away from the nucleus which in turn , increases the tendency of the atom to lose one of the electrons.

In the period:

The ionization energies in the period increases as the atomic number of an element increases because of the increase in the positive charge of the nucleus and the occurrence of the electrons in the same main level of energy. The attraction energy to attract the electron by the positive charges of the nucleus becomes greater.

However, there is an exception to this. If an atom has a secondary saturated shell such as (ns^2) or half saturated shell such as np^3,its ionization energy is greater than the ionization energy of the following atom .

An example for this is $_{7}N$ which has greater ionization energy than the ionization energy of $_{8}O$ in spite of the fact that oxygen atom has a bigger atomic number than the atomic number of nitrogen and that they both belong to the same period.

$_{8}O = 1S^2\ 2S^2\ 2P^4$

$_{7}N = 1S^2\ 2S^2\ 2P^3$

Noble gases have the greatest ionization energy so they do not lose electrons easily.

3. Electronegativity:

In many chemical compounds, the negative charge of the bonded electrons is centered 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 and thus, is given number (4) as a measure for its electronegativity. This number fluoride electronegativity is used as a measurement for all other elements.

In the period:

In general electronegativity increases as the atomic number increases in the period with some exceptions.

In the group:

Electronegativities are arranged in the groups from top to bottom. As the atomic number becomes greater, the electronegativity of an element decreases.

4 . Metallic and Nonmetallic Properties.

The metallic and nonmetal properties change according to the changes in the atomic number of the atoms in a same group and in a same period. As the atomic number of the atoms in the same period increases, the metallic properties decrease on one hand, and the nonmetallic properties increase on the other hand.

For example, lithium and beryllium in the second period show metallic properites .Boron and silicon come after with properties of metalloid. The rest of the elements in the period such as nitrogen, oxygen and fluoride come at the end with nonmetallic properties.

In one group, the metallic properties increase and the nonmetallic properties decrease as the atomic number increases.

All the elements in Group IA and Group IIA are metals.
The elements in Group VIA and Group VIIA are nonmetals.
The elements in the rest of the groups are not of the same type. For example, nitrogen in Group VA shows nonmetallic properties whereas antimony and arsenic show metalloid properties. Bismuth is the last element in group V and it shows metallic properties.

As far as the periods are concerned, the two elements in the first period (hydrogen & helium) are nonmetals. In the following four periods, there is a gradual change from metallic to nonmetallic properties. All the elements in the sixth period are metals except the last two elements which are nonmetals. The seventh period includes only metals, taking into account the fact that transition elements, lanthanide and actinides which are internal transitional elements show metal properties.

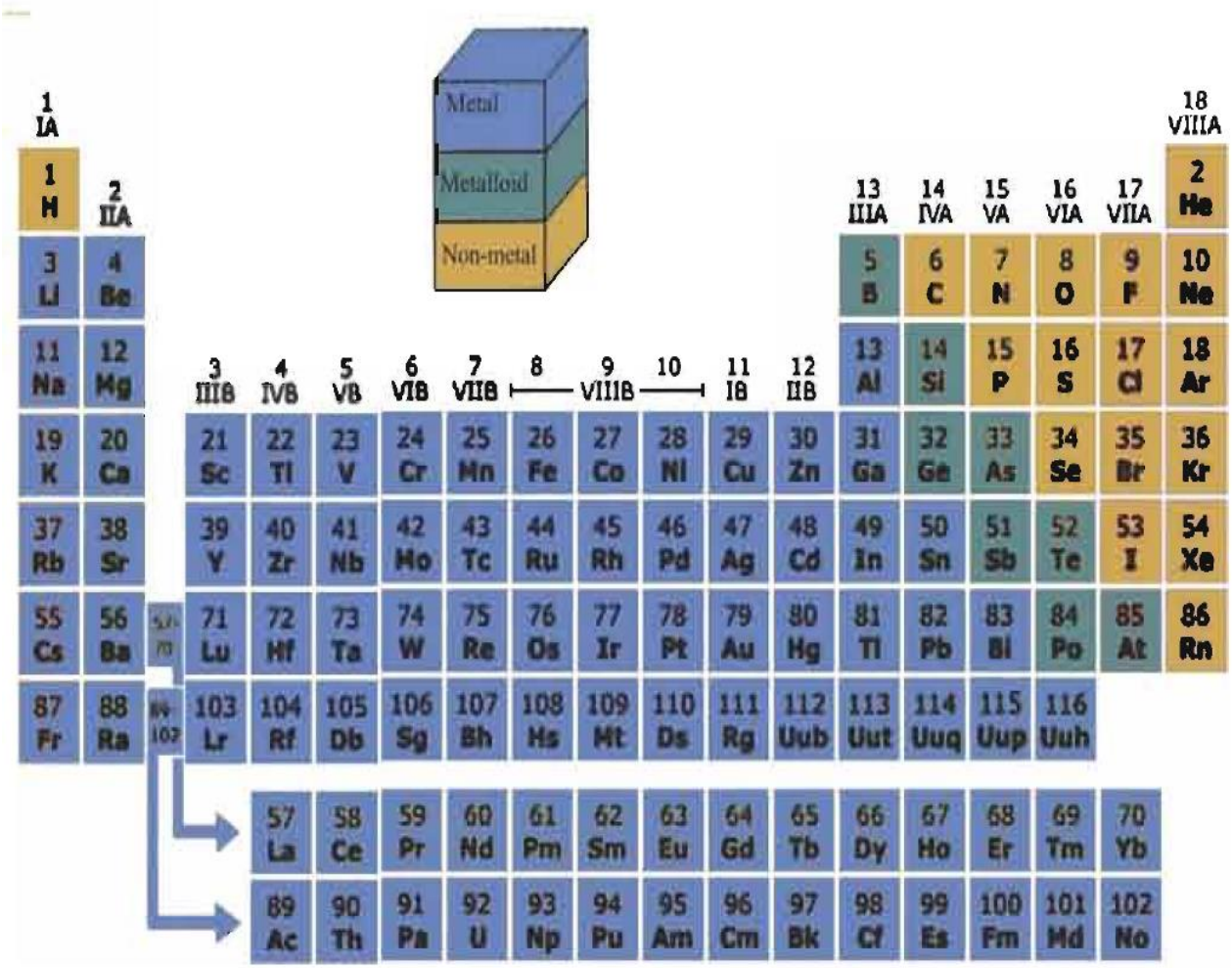

Figure 6: Metals and Non-metals

Periodic Table of the Elements

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	11 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 *La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 +Ac	104 Rf	105 Ha	106 Sg	107 Ns	108 Hs	109 Mt	110 110	111 111	112 112	113 113					

* Lanthanide Series	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
+ Actinide Series	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

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**