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WHERE SHOULD ALUMINUM GO IN THE PERIODIC TABLE ?

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CLASSIFICATION OF METALS

Since metals are those elements capable of losing electrons, therefore, they can be divided into typical, less typical, transition, and inner transition. This division is a result of their electronic structure (Figure 1).

Typical metals

These are the alkali metals, the alkaline earths, and aluminum. They have the following characteristics:

- They have an electronic structure similar to that of the inert gases with one, two, or three electrons in the outermost shell.
- They have single valency, i.e., they lose their outermost electrons in a single step.
- They are reactive, i.e., react readily with water and oxygen. The driving force for this reactivity is the inclination to achieve maximum stability by attaining the electronic structure of an inert gas. A reactive metal such as aluminum or magnesium may be used as a material of construction because of the protective oxide film that is formed rapidly on its surface.
- They form only colorless compounds.
- Within a certain vertical group the atomic radius increases with increasing atomic number because of the added electron shells. Within a certain vertical group the reactivity increases with increasing atomic number because of the ease with which the outermost electrons will be lost since they are further away from the nucleus. Thus cesium is more reactive than rubidium, and rubidium more than potassium, etc.
- With increasing charge on the nucleus, the electrostatic attraction for the electrons increases and the outermost electrons will not be easily lost hence the reactivity decreases. Thus magnesium is less reactive than sodium, calcium less than potassium, and so on.
- With increased electrostatic attraction for the electrons as a result of increasing charge on the nucleus, the size of the atom decreases. Thus, aluminum has a smaller radius than magnesium, and magnesium smaller than sodium.
- With decreased radius and increased atomic weight the atom becomes more compact, i.e. the density increases. Thus, aluminum has higher density than magnesium, and magnesium higher than sodium.
- They have appreciable solubility in mercury and form compounds with it except beryllium and aluminum.

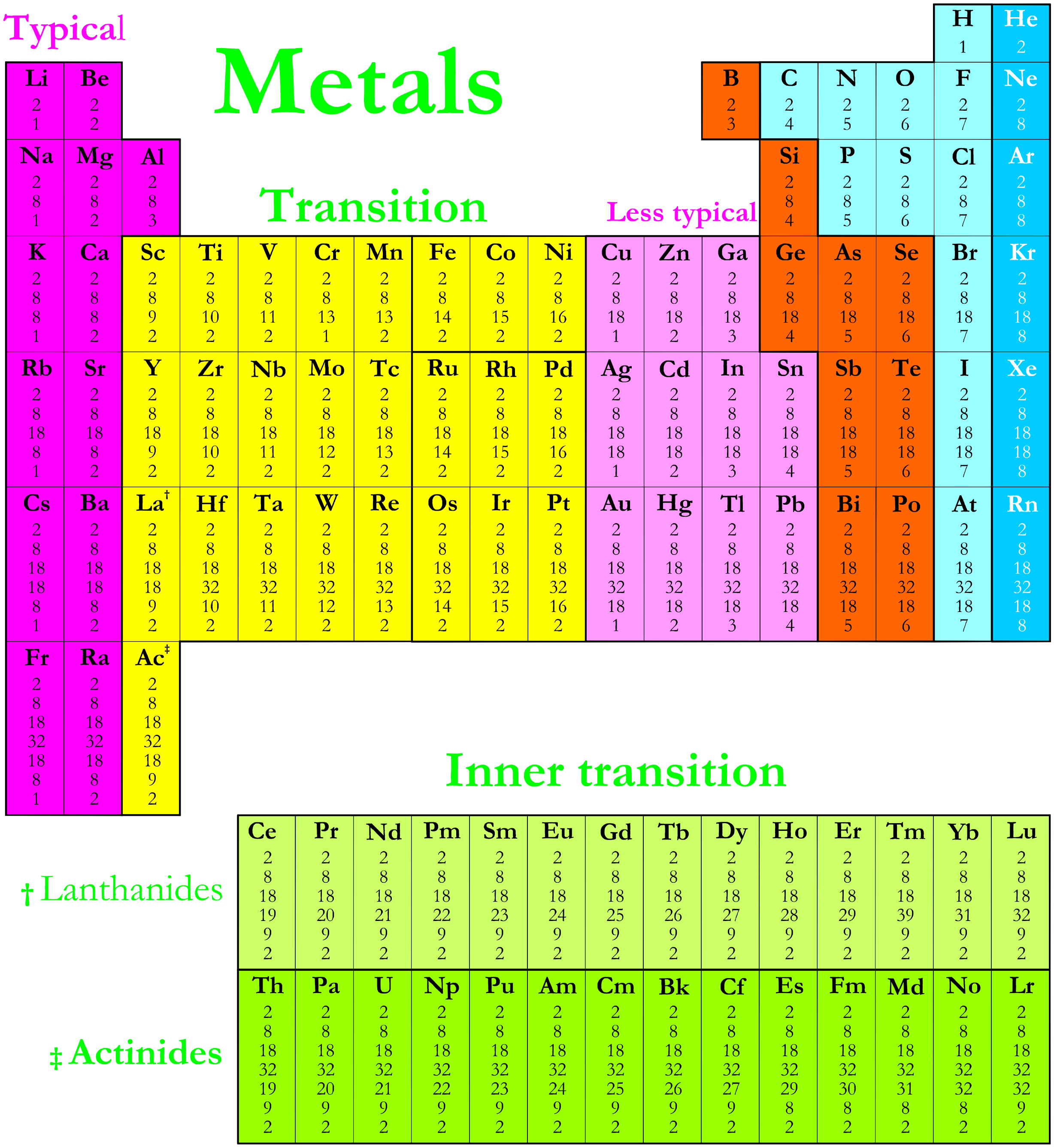

Figure 1- Electronic configuration of the elements

Less typical metals

These metals are: copper, silver, gold, zinc, cadmium, mercury, gallium, indium, thallium, tin, and lead. They differ from the typical metals in that they do not have an electronic structure similar to the inert gases; the outermost shell may contain up to four electrons and the next inner shell contains 18 instead of 8 electrons as in the inert gas structure. As a result of their electronic configuration they are characterized by the following:

- The atomic radius is less than the corresponding typical metals in the same horizontal group because the presence of 18 electrons in one shell results in an increased electrostatic attraction with the nucleus. Thus, the atomic radius of copper is less than potassium, silver less than rubidium, and gold less than cesium. However, the atomic radius increases with increased number of electrons in the outermost shell (which is contrary to the typical metals), i.e. the atomic radius of gallium is larger than that of zinc, and zinc is larger than copper. This is demonstrated in Figure 2: The atomic volume of the typical metals decreases with increased atomic number while the reverse is true for the less typical metals. the reason for this is the shielding effect of the 18-electron shell, the increased repulsion of the additional electron in the outmost shell and that shell, and also the increased repulsion between the electrons themselves in that shell.

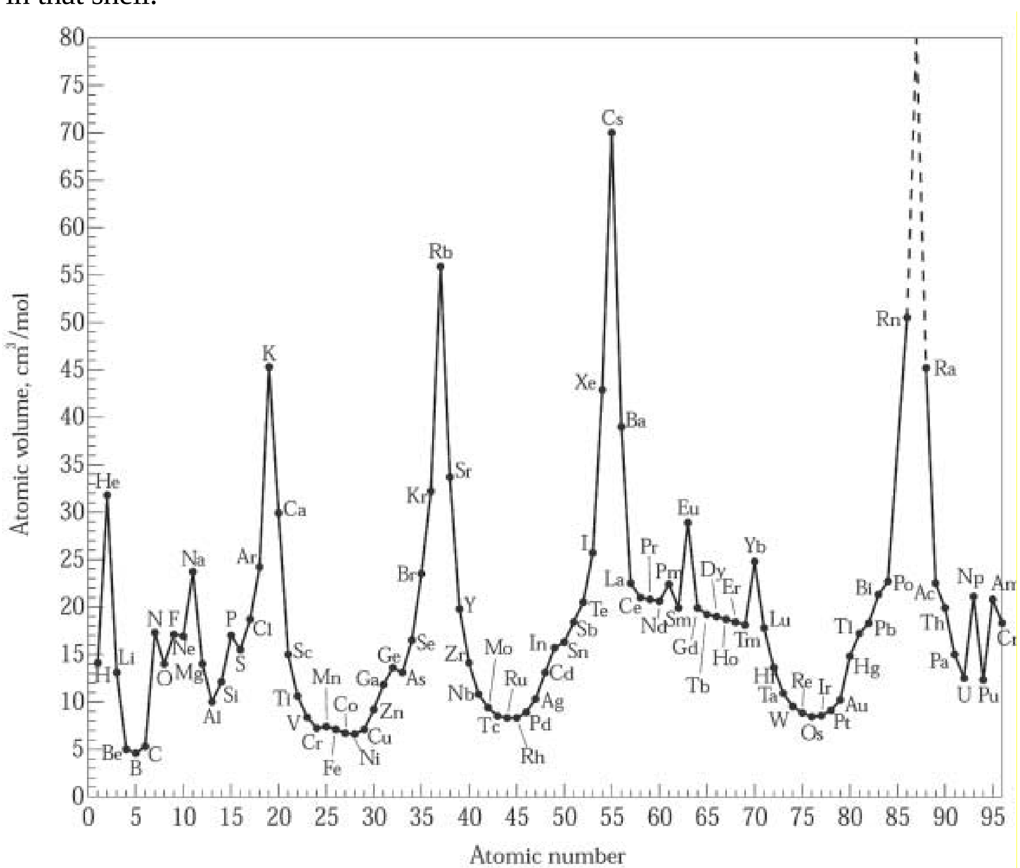

Figure 2- Atomic volume of the elements

- The outermost electrons will not be easily lost, i.e. these metals are less reactive than their corresponding typical metals for two reasons:
    - There is no driving force to lose electrons since an inert gas electronic structure will not be achieved.
    - There is a stronger electrostatic attraction due to the smaller atomic radius as compared to that of the typical metals.
- Because of the higher atomic weight and the smaller atomic radius these metals are more dense than their corresponding typical metals.
- Some of these metals show two different valency states, e.g., copper as Cu^{I} and Cu^{II}, gold as Au^{I} and Au^{III}, mercury as Hg^{I} and Hg^{II}, tin as Sn^{II} and Sn^{IV}, and lead as Pb^{II}

and Pb^{IV}. This is because of the possibility of removing one or two electrons from the 18-electron shell.

- Few of these metals from colored ions in solution, e.g., Cu^{II} and Au^{III}, or colored compounds, e.g., copper sulfate pentahydrate (blue), cadmium sulfide (yellow), etc. This is due to the possibility of movement of electrons from the 18 electrons shell to a higher level.
- They have the highest solubility in mercury since their electronic structure is similar as that of mercury. Also, they do not form compound with mercury

Transition metals

These are the metals in the vertical groups in the Periodic Table from scandium to nickel. They not only have electronic configuration different from the inert gases but they are characterized by having the same number of electrons in their outermost shell and a progressively greater number of electrons in the next inner shell. There are, however, some apparent irregularities in the number of electrons in the outermost electron shells. This is due to energy levels, which are determined from spectroscopic measurements. As their name implies the transition metals have properties between the typical and less typical metals.

They are less reactive than the typical metals because they will not achieve the inert gas structure when they lose their outermost electrons, but they are nevertheless more reactive than the less typical metals. They share the following properties:

- They resemble each other quite closely besides showing the usual group relationships because they have the same number of the outermost electrons.
- They may lose additional electrons from the next lower shell to form ions with higher charges. As a result they show a variable valence. For example, vanadium exists in +2, +3, +4, and +5 oxidation states, and titanium in +2, +3, and +4.
- The atomic radius of the successive metals in a certain horizontal period decreases slightly as the atomic number rises because when an electron is added to an inner shell it decreases slightly the size of the atom as a result of increased electrostatic attraction.
- Most of them form colored ions in solution due to electronic transition with the exception of the group Sc, Y, La and Ac that form only colorless compounds.
- They form many covalent compounds, e.g., the carbonyls of iron and nickel, the chlorides of titanium, and the oxyacids of chromium, molybdenum and tungsten.
- They form coordination compounds with ammonia, e.g., the ammines of cobalt and nickel.
- They mostly form borides, carbides, nitrides, and hydrides, which have mostly metallic character.
- They have the lowest solubility in mercury.

The transition metals can be divided into three groups:

- *Vertical similarilty transition metals.* These are the vertical groups scandium to manganese. They show similarity in the vertical direction, e.g., Zr-Hf, Nb-Ta, and

Mo-W. The group Sc, Y, La, and Ac form colorless compounds and have the same valency (+3).

- *Horizontal similarity transition metals*. This is the group titanium to nickel. They show similarity in the horizontal direction.

  - Their carbides have intermediate properties between the metal-like character of the transition metals and the ionic character of the typical metals. Thus they have metallic luster and electrically conductive, but they are attacked by water and dilute acids.
  - They form di- and trivalent compounds.
  - These metals have melting points in the range 1220 to 1800^{o}C.
  - Iron, cobalt, and nickel occur in nature together in the native state in the minerals awariait, $Fe(Ni,Co)_3$, and josephinite, $Fe(Ni,Co)_2$.

- *Horizontal - vertical transition metals*. This is the platinum metals group where the similarity between the six metals is in the horizontal and vertical direction.
  - They resist corrosion.
  - They occur together in nature in the native state.

Inner transition metals

These metals have the same number of electrons in the two outermost shells but a progressively greater number of electrons in the next inner shell. They form two groups: the lanthanides and the actinides

DISCUSSION

In 1985 the *International Union of Pure and Applied Chemistry* (IUPAC) recommended that the groups in the Periodic Table should be numbered from 1 to 18, i.e., the alkali metals are Group 1 and the inert gases are Group 18 [1,2]. This numbering was a compromise between the North American and the European Periodic Tables (Figure 3). But this raised a hot debate among chemists that was published in *Chemical & Engineering News* in a series of letters to the editor [3-5]. Almost all these letters opposed the new numbering system. It is believed that this conflict may be resolved if chemists and physicists abandon the old idea of numbering the groups. Instead, it would be more reasonable to give names to each group of the elements that reflect their electronic structure, the nature of their bonding, and their properties (both physical and chemical). These names should be easier for students to remember and give them a deeper understanding of the Periodic Table.

The present classification in ten groups (Figure 4) should be helpful in this respect. In this classification aluminum is moved to a position immediately next to magnesium so that the alkali metals, the alkaline earth metals, and aluminum form a group of typical metals. The other groups are: the less typical metals, the vertical transition metals, the horizontal transition metals, the vertical-horizontal transition metals, the inner transition metals (the lanthanides and the actinides), the metalloids, the monatomic nonmetals, and the covalent bond nonmetals.

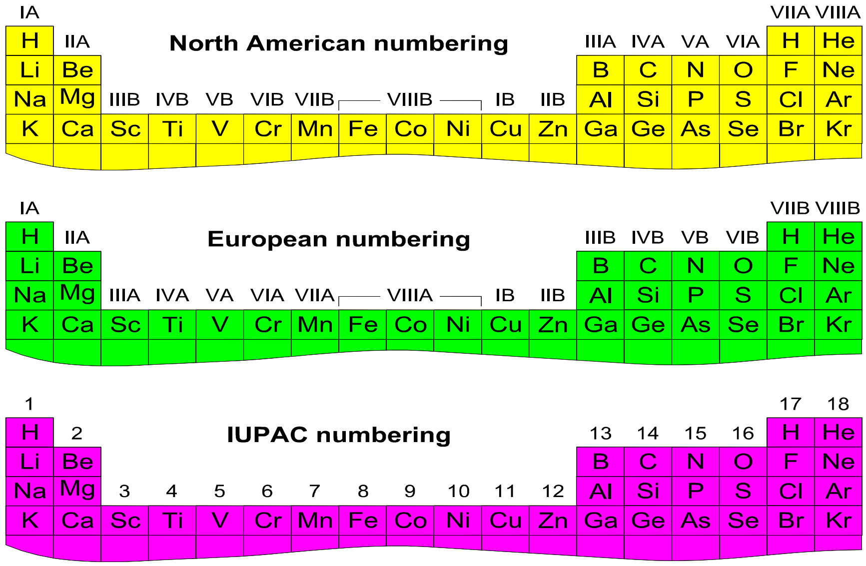

Figure 3- Different numbering of the groups in the Periodic Table

An advantage of this classification would be putting each member of a group in its logical position.

The position of aluminum

Moving aluminum further away from gallium with which it occurs in bauxite may raise objection. However, it should be recalled that there are marked differences between both metals:

- While aluminum oxidizes so rapidly that it soon forms a non-porous protective layer, gallium does not.
- Gallium can be electrodeposited from aqueous solution, while aluminun cannot.
- $Al(OH)_3$ does not dissolve in ammonium hydroxide solution, but $Ga(OH)_3$, does dissolve.
- Gallium is precipitated from aqueous solution by H_2S as a sulfide; aluminum does not.
- Aluminum forms carbides, gallium does not.
- Aluminum carbide, Al_4C_3 is similar to the carbides of the first three groups, being ionic colorless compound containing the C_2^{2-} anion that decomposes in water. However, it liberates methane instead of acetylene like the other members of the group:
  $Al_4C_3 + 12\ H_2O \rightarrow 4\ Al(OH)_3 + 3\ CH_4$
- Gallium forms a gaseous hybride, Ga_2H_6, while aluminum forms a white solid polymer hydride, $(AlH_3)_x$
- Historically, aluminum oxide was considered an earth like the rare earths and it fits with the scandium group.
- Mendeleev successfully predicted the properties of scandium before it was discovered by linking it with aluminum.
- The chemistry of scandium is very similar to that of aluminum and the ions Al^{3+}, Sc^{3+}, Y^{3+}, La^{3+}, and Ac^{3+} form a series similar to Mg^{2+}, Ca^{2+}, Sr^{2-}, Ba^{2+} and Ra^{2+}.
- Gallium is a typical dispersed element whose relative abundance in the Earth's crust is 1.5 x 10^{-3}%, aluminum is the third most abundant element with a relative abundance of 8.13% (after oxygen and silicon).

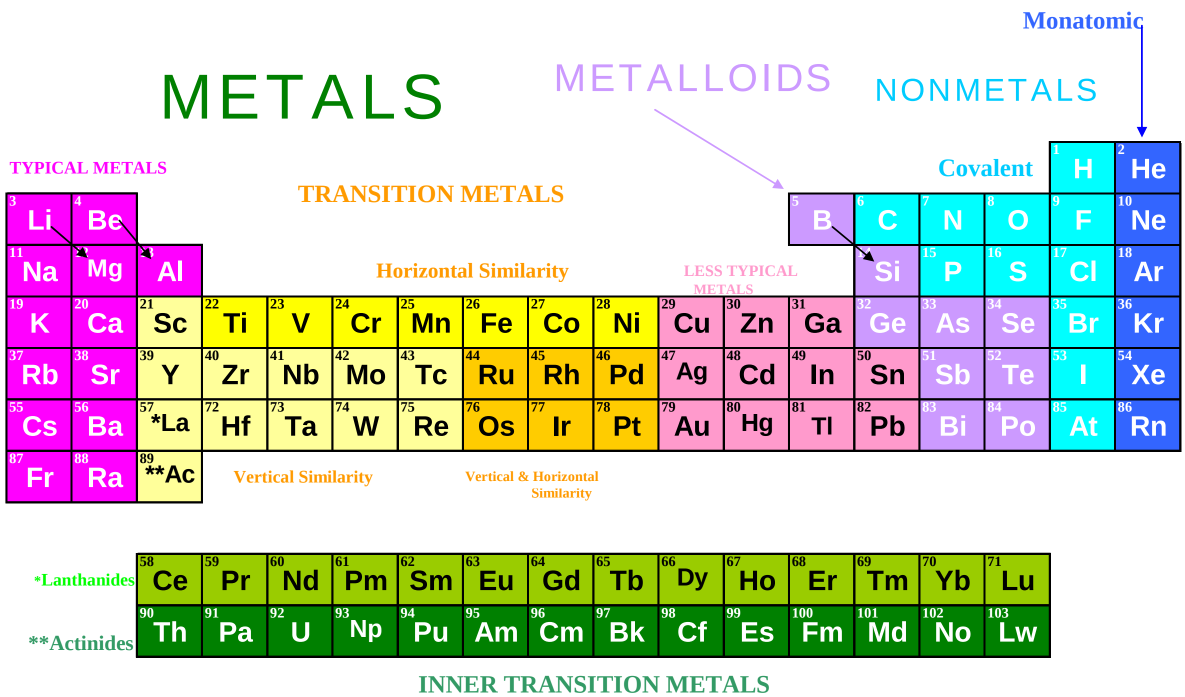

Figure 4 - A Periodic Table in ten groups

Another objection, which may arise from this transfer, is the dissimilarity of the electron orbital structure of aluminum as compared to Sc, Y and La: aluminum has 2s and 1p electrons, while the others have 2s and 1d electrons. This objection can be answered in the following. The classification of the elements in the Periodic Table according to the population of electrons in the subshells, i.e., as s-, p-, d, and f-block elements does not indicate or explain many facts such as:

- The boron group elements are in the p-block but show different chemical and physical properties, e.g., boron is a metalloid while the other elements are metals. Further, the structure of boron hydroxide (boric acid) is fundamentally different from aluminum hydroxide. In the former no oxygen atoms are shared between the $M(OH)_n$ coordination groups in contrast to the latter.
- Boron group elements have three electrons in their outer shell: 2s and 1p. However, all three electrons are lost in one step except thallium and to a minor extent gallium. On the other hand the scandium group elements have also three electrons in their outer shell: 2s and 1 d, and all three are lost in one step. Hence the distinction in electron orbital is not a critical criterion for classification of the elements, and it is the total number of electrons that should be taken into consideration.
- The copper and zinc groups of elements are in the d-block which is used to be known as transition metals but differ from the other members of the same block in that the atomic radius increases with atomic number whereas in the other members it is the opposite. As a result they have different chemical properties and should not be included in the transition metals group.
- There are many irregularities in the electron orbitals in the subshells, which arises from quantum mechanical considerations. For example:

- Chromium and manganese, and molybdenum and technetium have the same number of d-electrons and different numbers of s-electrons instead of the reverse.
- Europium and gadolinium have similar f- and different d-electrons instead of the reverse.

This renders the strict adherence to the s-p-d, and f-classification pointless, and it should be better to consider the total number of electrons in a shell. It should also be noted that this classification was introduced by physicists who are not necessarily interested in chemical properties.

CONCLUSIONS

Aluminum should be present above scandium in the Periodic Table so that, together with the alkali metals and the alkaline earth metals, they form the typical metals group with related properties and electronic configuration. In addition chemists abandon the old tradition of numbering the groups in the Periodic Table, and to give the following descriptive names instead:

- Monatomic nonmetals
- Covalent nonmetals
- Metalloids
- Typical metals
- Less typical metals
- Transition metals with vertical similarity
- Transition metals with horizontal similarity (Ti to Ni)
- Transition metals with vertical and horizontal similarity (platinum group metals)
- Inner transition metals:
    - ☐ The lanthanides
    - ☐ The actinides

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