Aluminium and its Position in the Periodic Table

FATHI HABASHI

Department of Mining and Metallurgy, Laval University, Quebec City, Canada GIK 7P4

Abstract

Aluminium seems to be misplaced in the Periodic Table. It is usually put below boron, a metalloid with which it has little in common. Aluminum should be transferred above scandium so that, together with the metals and the alkaline earth metals, these metals represent the typical metals with related properties and electronic configuration.

Introduction

The Periodic Table (Figure 1) was developed by chemists more than one hundred years ago as a correlation for the properties of the elements. With the discovery of the internal structure of the atom, it became recognized by physicists as a natural law. When the crystalline structure of solids was studied, the nature of the chemical bond was understood, and the theory of metals was put forward, it became an essential tool not only for chemists and physicists, but for metallurgists as well. Of the 87 naturally occurring elements, 63, i.e. about three fourth can be described as *metals*, 16 as *nonmetals*, and 9 as *metalloids* which have properties in between the two classes (Table 1)[1,2].

In the solid state metals are composed of crystals made of closely packed atoms whose outer electrons are so loosely held that they are free to move through the crystal lattice. This structure explains their mechanical, physical, and chemical properties. Nonmetals include the inert gases, hydrogen, oxygen, fluorine, and chlorine, liquid bromine, and the solid elements carbon, sulfur, phosphorus, and iodine. These elements do not have the properties of a metal. Except the inert gases which are monatomic, nonmetals readily share electrons. Their atoms are united together by covalent bond, i.e., atoms that share their outer electrons. They often form diatomic molecules such as H_2, Cl_2, N_2, or larger molecules such as P_4 and S_8, or giant molecules, i.e., a network of atoms

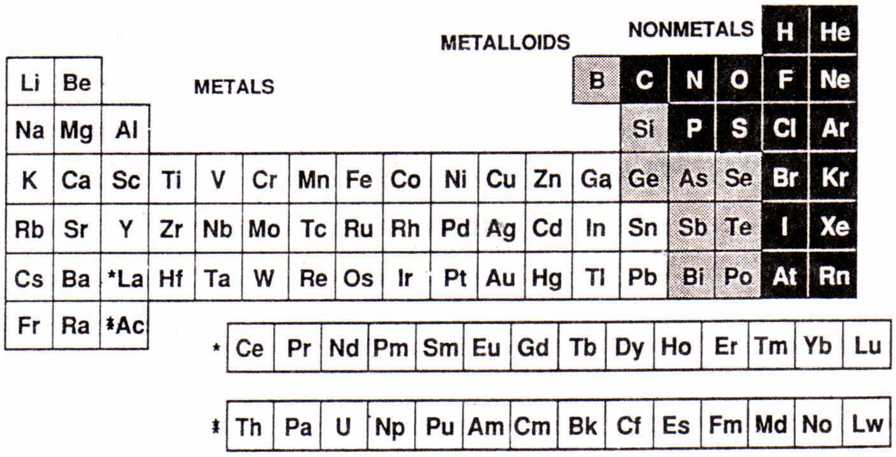

Figure 1 : The Periodic Table

of indefinitely large volume such as carbon in form of graphite or diamond. Metalloids have intermediate properties between metals and non-metals. Under certain conditions they behave as metals while under other conditions they behave as a nonmetals. Unlike metals their atoms are united together by covalent bond.

Since metals are those elements capable of losing electrons, therefore, they can be divided into typical, less typical, transition, and inner transition as a result of their electronic structure (Figure 2).

Typical And Less Typical Metals

Typical metals are the alkali metals, the alkaline earths, and aluminium. They have the following characteristics :

- They have an electronic structure similar to that of the inert gases with one, two or three electons 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.
- They form only colourless 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, aluminium 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, aluminium has higher density than magnesium, and magnesium higher than sodium.

The *less typical 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 result in an increased electrostatic attracion 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 zinc, and zinc is larger than copper. 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. This is obvious when the atomic volumes are compared (Figure 3)[3].
- The outermost electrons will not be easily lost, i.e., these metals are less reactive than their corresponding typical metals because of the stronger electrostatic attraction due to the smaller atomic radius.
- 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

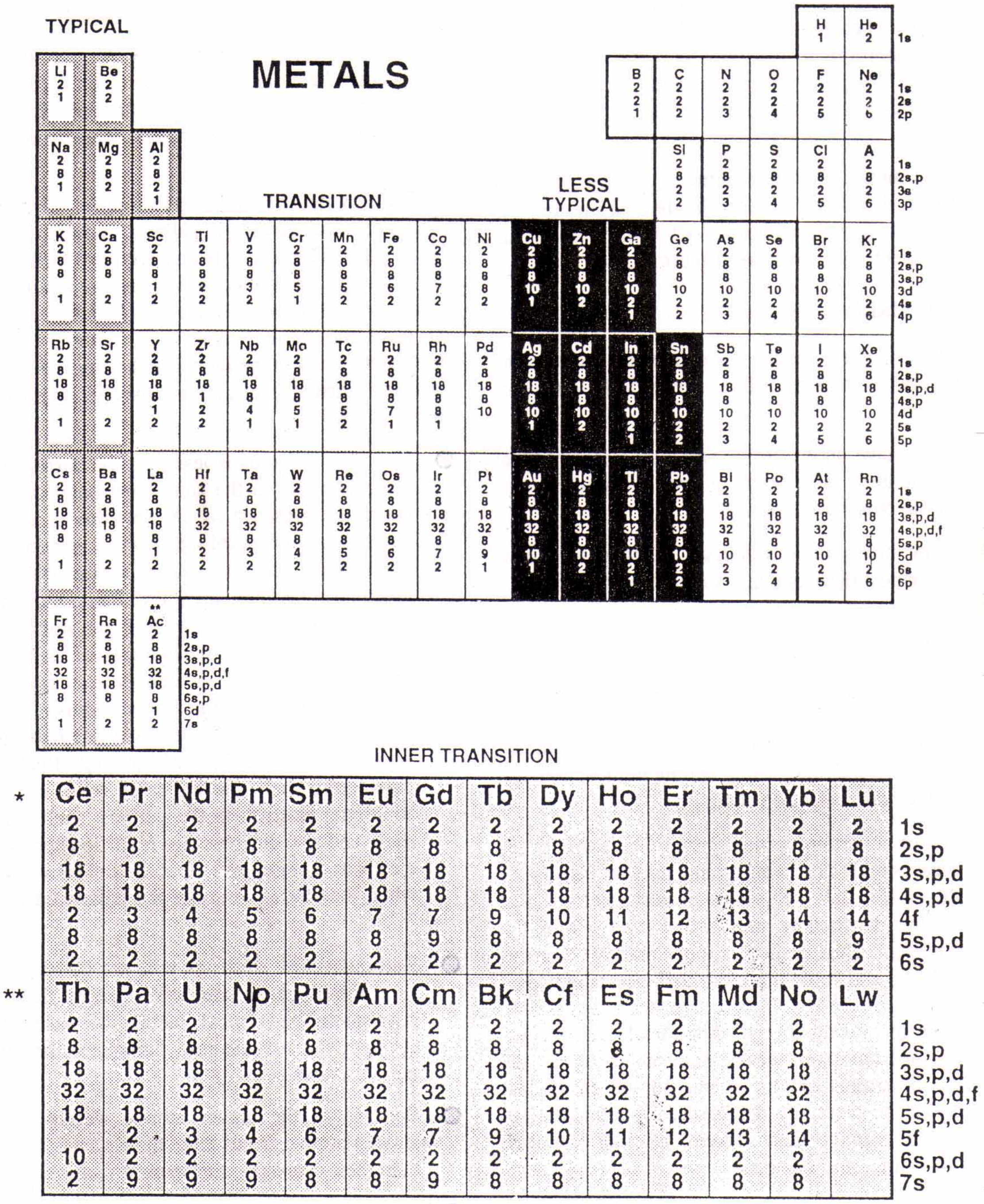

Figure 2 : Electronic configuration of the elements

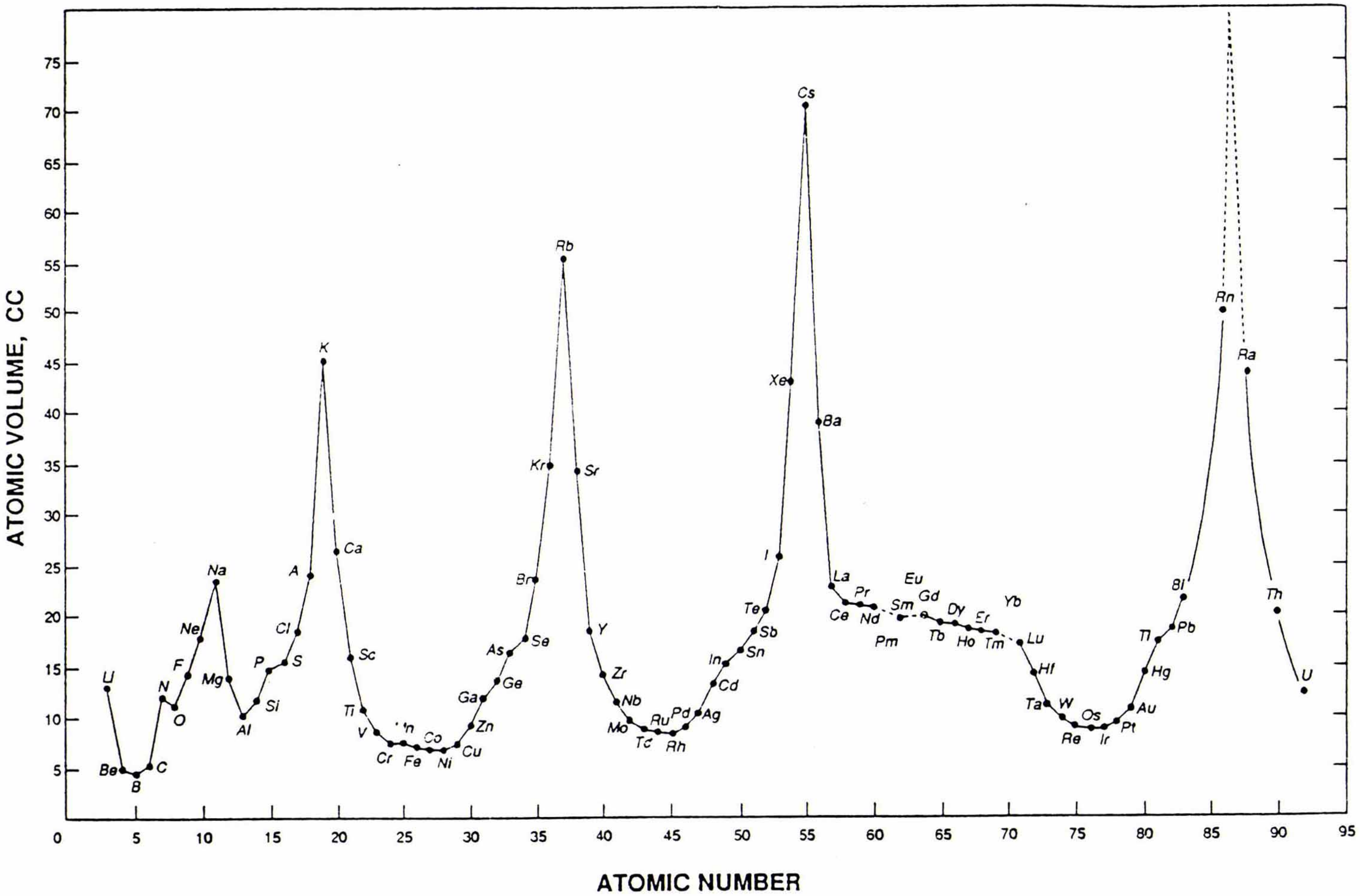

Figure 3. Atomic volumes of the elements. Atomic volume is the volume in cubic centimeters occupied by one gram atomic weight of the element in the solid state:Atomic volume = atomic weight/density Atomic volumes can be used as qualitative guides to the relative volumes of the individual atoms since all gram atomic weights contain the same number of atoms.

Sn^{IV} and lead as Pb^{II} and Pb^{IV}. This is because of the possibility of removing one or two electrons form the 18-electron shell.

- Few of these metals form coloured ions in solution, e.g., Cu^{II} and Au^{III} or coloured compounds, e.g., copper sulfate pentahydrate (blue), cadmium sulfide (yellow), etc. This is due to the possibility of movement of electrons from the 18-electron shell to a higher level.

Transition and Inner Transition Metals

The *transition metals* 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 shells. There are, however, some apparent irregularities in the number of electrons in the outermost electron shells. This is due to energy levels of electrons which are determined from spectroscopic measurements. As their name implies the transition metals are a bridge 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 coloured ions in solution which is due to electronic transitions.
- They form 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 form borides, carbides, nitrides, and hydrides which have mostly metallic character.

The *inner transition 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: These are the metals following lanthanum, namely cerium to lutetium.
- The actinides: These are the metals following actinium, namely thorium to lawrencium.

The lanthanides are not only similar to the transition metals but some of the characteristics of transition metals are even more pronounced because of their unique electronic structure. Thus, beside showing multiple valency, forming coloured ions in solution and being less reactive than the typical metals, they are so similar in chemical properties that their separations is difficult and is usually done by making use of differences in physical properties. Of the actinides, thorium, uranium, and plutonium are of practical importance. Uranium and plutonium have multiple valency (+3, +4, +5, and +6) and form coloured compounds.

Discussion

Moving aluminium 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 aluminium oxidizes so rapidly that it soon forms a non-porous protective layer, gallium does not. Aluminium forms only one oxide, Al_2O_3, gallium forms two: Ga_2O_3 and Ga_2O. Gallium can be electrodeposited from aqueous solution, while aluminium cannot. Further, $Al(OH)_3$ does not dissolve in ammonium hydroxide solution, but $Ga(OH)_3$ does dissolve[4]. Also, gallium is a typical dispersed element whose relative abundance in the Earth's crust is $1.5 \times 10^{-3}\%$, aluminium is the third most abundant element with a relative abundance of 8.13% (after oxygen and silicon).

Table 1 - Metals, metalloids and nonmetals

Metals	Metalloids	Nonmetals
Crystalline solids (except mercury) with metallic lustre.	May be crystalline or amorphous Sometimes have metallic lustre.	Form volatile or non-volatile molecules, no metallic lustre.
Do not readily share electrons, their vapours are monatomic.	Readily share electrons even in the elemental form.	Readily share electrons; form diatomic, large or giant molecules; inert gases are monatomic.
Exhibit electrical and thermal conductivity. Electrical resistance usually increases with increased temperature.	Low electrical and thermal conductivity.	Do not conduct electricity or heat. Electrical resistance decreases with increased temperature.
Have high density and useful mechanical properties.	Moderate density, no useful mechanical properties.	Low density, of no useful mechanical properties.
Electropositive, form cations, e.g., Cu^{2+}, Na^{+}, etc.	Sometimes electropositive, sometimes electrtonegative.	Electronegative, form anions, e.g., S^{2-}, Cl^{-}, etc.
From basic oxides, e.g., CaO	Form acidic oxides.	Form acidic oxides, e.g., SO_2
Deposit on the cathode during electrolysis.	Deposit on the cathode.	Deposit on the anode.
Either form no compounds with hydrogen or form unstable compounds usually nonvolatile (metal hydrides)	Form stable compounds with hydrogen, e.g., AsH_3, H_2Se.	Form stable compounds with hydrogen, usually volatile, e.g., NH_3, PH_3, H_2S, etc.

Another objection which may arise from this transfer is the dissimilarity of the electron orbital structure of aluminium as compared to Sc, Y, and La: aluminium 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: [5,6]

- 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 aluminium hydroxide. In the former no oxygen atoms are shared between the $M(OH)_n$ coordination groups in contrast to the latter[7].
- 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 1d, 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. Thus aluminium fits as well in the scandium group to an advantage.
- The copper and zinc groups of elements are in the d-block which is commonly 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.

Further there are many irregularities in the electron orbitals in the subshells which arises from quantum mechanical considerations. For example, chromium and manganese have the same number of d-electrons and different numbers of s-electrons instead of the reverse. Similalry molybdenum and technetium, europium and gadolinium (in this case similar f- and different d- electrons), and others. 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

The group starting with boron and ending with thallium in the Periodic Table is in fact composed of three different groups: Boron is a metalloid with a high melting point, aluminium is a typical metal with a low melting point, gallium, indium, and thallium are less typical metals. It is therefore suggested to remove aluminium from this group and put it above scandium so that all the typical metals are on one side in the Table. When aluminium occupies this new position then the first three groups of the Periodic Table will be common in forming colourless compounds and in having single valency.

References

1. F. Habashi, *Metallurgical Chemistry*, ACS Audio Course, American Chemical Society, Washington, D.C. 1987.
2. H.H. Sisler, *Electronic Structure, Properties, and the Periodic Table*, Reinhold, New York, 1963.
3. H. Remy, *Lehrbuch der Anorganischen Chemie*, volume 1, pp. 24-26, Akademische Verlag, Leipzig 1965.
4. A.N. Zelikman, O.E. Krein, and G.V. Samsonov. *Metallurgy of Rare Metals*, translated from Russian, Israel Program of Scientific Translations, Jerusalem 1966, pp. 337-340.
5. J.C. Kotz and K.F. Purcell, *Chemistry and Chemical Reactivity*, Saunders, New York 1987, pp. 262-271.
6. K.M. Mackay and R.A. Mackay, Introduction to Modern Inorganic Chemistry, International Textbook Co., London, 1974. This textbook is based mainly on the s-, p-, d-, f-block elements classification.
7. A.F. Wells, Structural Inorganic Chemistry, Clarendon Press, Oxford 1962, pp. 551-555.

****