



# A NEW LOOK AT THE PERIODIC TABLE

Fathi Habashi

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As science advances, its laws become fewer but of greater scope. In this respect the Periodic Law, which is the basis of the Periodic Table, represents a major step in the progress of chemistry — it affords the natural classification of the elements. The Periodic Table 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 bonds 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 are described as *metals*, 16 as *nonmetals*, and 9 as *metalloids*. Chemists should abandon numbering the groups in the Periodic Table and to give descriptive names instead.

\* Corresponding Author

E-Mail: [Fathi.Habashi@arul.ulaval.ca](mailto:Fathi.Habashi@arul.ulaval.ca)

[a] Department of Mining, Metallurgical, and Materials Engineering, Laval University, Quebec City, Canada

## INTRODUCTION

Metals are the most common articles in everyday life; they are usually used in form of alloys, which are a combination of two or more metals. They are the basis of the metallurgical industry. Nonmetals, except carbon, are hardly used by an average person. Air, a mixture of nonmetals is known to exist but is not seen by people. Nonmetals are the basis of the chemical industry. Metalloids are the basis of advanced technology and the electronics industry (Figure 1). 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 throughout the crystal lattice. This structure explains their mechanical, physical, and chemical properties.

Nonmetals include the inert gases<sup>1</sup>, hydrogen, oxygen, nitrogen, fluorine, and chlorine, liquid bromine, and the solid elements carbon, sulfur, phosphorus, and iodine. These elements do not have the properties of a metal. Nonmetals except the inert gases 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<sub>2</sub>, Cl<sub>2</sub>, N<sub>2</sub> or larger molecules such as P<sub>4</sub> and S<sub>8</sub>, or giant molecules, i.e. a network of atoms of indefinitely large volume such as carbon in form of graphite or diamond.

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Inert gases is a historical name for the group of gases starting with helium and ending with radon. They were until the 1960's considered inert when few compounds of xenon with fluorine were prepared.

Figure 1 shows the periodic table with the following labels:

- Metals:** Groups 1-10 (Li, Be, Na, Mg, Al, K, Ca, Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn, Ga, Ge, As, Se, Br, Kr, Rb, Sr, Y, Zr, Nb, Mo, Tc, Ru, Rh, Pd, Ag, Cd, In, Sn, Sb, Te, I, Xe, Cs, Ba, La', Hf, Ta, W, Re, Os, Ir, Pt, Au, Hg, Tl, Pb, Bi, Po, At, Rn, Fr, Ra, Ac')
- Metalloids:** Groups 11-16 (B, C, N, O, F, Ne, Si, P, S, Cl, Ar)
- Nonmetals:** Groups 17-18 (H, He)
- Metallurgical Industry:** Groups 1-10
- Hi-tech Industry:** Groups 11-16
- Chemical Industry:** Groups 17-18

† Ce, Pr, Nd, Pm, Sm, Eu, Gd, Tb, Dy, Ho, Er, Tm, Yb, Lu  
‡ Th, Pa, U, Np, Pu, Am, Cm, Bk, Cf, Es, Fm, Md, No, Lr

Figure 1. The periodic table

Figure 2 shows the electronic configuration of the metals, categorized into:

- Typical Metals:** Groups 1-10 (Li, Be, Na, Mg, Al, K, Ca, Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn, Ga, Ge, As, Se, Br, Kr, Rb, Sr, Y, Zr, Nb, Mo, Tc, Ru, Rh, Pd, Ag, Cd, In, Sn, Sb, Te, I, Xe, Cs, Ba, La', Hf, Ta, W, Re, Os, Ir, Pt, Au, Hg, Tl, Pb, Bi, Po, At, Rn, Fr, Ra, Ac')
- Transition Metals:** Groups 11-16 (Cu, Zn, Ga, Ge, As, Se, Br, Kr, Rb, Sr, Y, Zr, Nb, Mo, Tc, Ru, Rh, Pd, Ag, Cd, In, Sn, Sb, Te, I, Xe, Cs, Ba, La', Hf, Ta, W, Re, Os, Ir, Pt, Au, Hg, Tl, Pb, Bi, Po, At, Rn, Fr, Ra, Ac')
- Less typical:** Groups 17-18 (H, He)
- Inner transition:** Lanthanides (Ce, Pr, Nd, Pm, Sm, Eu, Gd, Tb, Dy, Ho, Er, Tm, Yb, Lu) and Actinides (Th, Pa, U, Np, Pu, Am, Cm, Bk, Cf, Es, Fm, Md, No, Lr)

Figure 2. Electronic configuration of the metals.

**Table 1.** General characteristics of 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 having no metallic lustre
Do not readily share electrons, their vapours are monoatomic	Readily share electrons even in the elemental form	Readily share electrons; form diatomic, large or giant molecules; inert gases are monoatomic
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., $\text{Cu}^{2+}$ , $\text{Na}^+$ , etc.	Sometimes electropositive, sometimes electronegative	Electronegative, form anions, e.g., $\text{S}^{2-}$ , $\text{Cl}^-$ , etc.
Form basic oxides, e.g., $\text{CaO}$	Form acidic oxides	Form acidic oxides, e.g., $\text{SO}_2$
Deposit on the cathode during electrolysis	Deposit on the cathode	Deposit on the anode, e.g., $\text{O}_2$ , $\text{Cl}_2$
Either form no compounds with hydrogen or form unstable compounds usually nonvolatile (metal hydrides)	Form stable compounds with hydrogen, e.g., $\text{AsH}_3$ , $\text{H}_2\text{Se}$	Form stable compounds with hydrogen, usually volatile, e.g., $\text{NH}_3$ , $\text{PH}_3$ , $\text{H}_2\text{S}$ , etc.

Metalloids have covalent bond like nonmetals, but have intermediate properties between metals and nonmetals. Table 1 summarizes the properties of metals, nonmetals, and metalloids.

## 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 2).

### 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.

### 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 3: The atomic volume<sup>2</sup> 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.

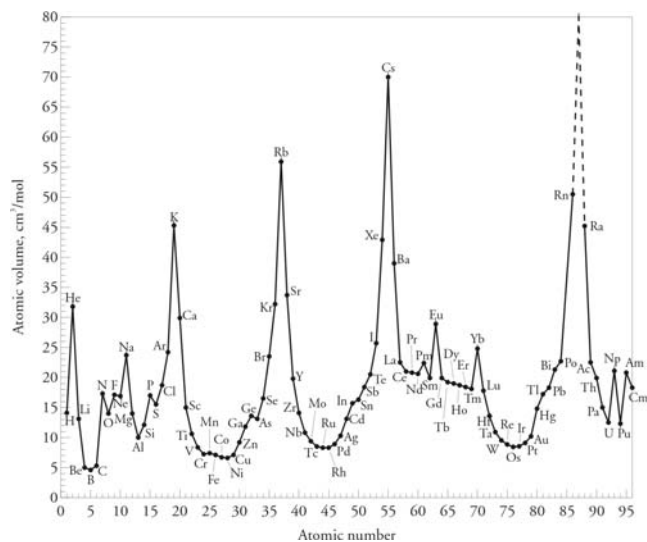


Figure 3. Atomic volume of elements.

- 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  $\text{Cu}^{\text{I}}$  and  $\text{Cu}^{\text{II}}$ , gold as  $\text{Au}^{\text{I}}$  and  $\text{Au}^{\text{III}}$ , mercury as  $\text{Hg}^{\text{I}}$  and  $\text{Hg}^{\text{II}}$ , tin as  $\text{Sn}^{\text{II}}$  and  $\text{Sn}^{\text{IV}}$ , and lead as  $\text{Pb}^{\text{II}}$  and  $\text{Pb}^{\text{IV}}$ . This is because of the possibility of removing one or two electrons from the 18-electron shell.

- Few of these metals form colored ions in solution, e.g.,  $\text{Cu}^{\text{II}}$  and  $\text{Au}^{\text{III}}$ , or colored compounds, e.g., copper sulfate pentahydrate (blue), cadmium sulfide (yellow), etc. (Table 2). 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.

Table 2. Colour of the less typical metal ions in solution.

	$\text{M}^+$	$\text{M}^{2+}$	$\text{M}^{3+}$	$\text{M}^{4+}$
Cu	colourless	blue	—	—
Zn	—	colourless	—	—
Ga	—	—	colourless	—
Ag	colourless	—	—	—
Cd	—	colourless	—	—
In	—	—	colourless	—
Sn	—	colourless	—	colourless
Au	yellow	—	red	—
Hg	colourless	colourless	—	—
Tl	colourless	—	colourless	—
Pb	—	colourless	—	colourless

### 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 (Table 3).

<sup>2</sup> Atomic volume is the volume in cubic centimeters occupied by one gram atomic weight of the element in the solid state. It 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.

**Table 3.** Colour of transition metal ions in solution. Incomplete list because many compounds are insoluble or when soluble, hydrolyse and precipitate

	M <sup>2+</sup>	M <sup>3+</sup>	MO <sup>2+</sup>	MO <sub>2</sub> <sup>+</sup>	MO <sub>2</sub> <sup>2+</sup>
Sc	—	colourless	—	—	—
Ti	—	violet	colourless	—	—
V	violet	green	blue	yellow	—
Cr	blue	green	—	—	—
Mn	pink	violet	—	—	—
Fe	green	yellow	—	—	—
Co	red	blue	—	—	—
Ni	green	—	—	—	—
Y	—	colourless	—	—	—
Zr	—	—	colourless	—	—
Nb	—	—	—	colourless	—
Mo	—	red	green	blue	—
Ru	red	—	—	—	—
Rh	red	red	—	—	—
Pd	yellow	—	—	—	—
La	—	colourless	—	—	—
Hf	—	—	colourless	—	—
Ta	—	—	—	colourless	—
W	—	—	—	green	yellow
Re	—	red	—	—	—
Os	brown	green	red brown	—	—
Ir	—	yellow	brown	—	—
Pt	green	green-black	red, yellow	—	—

○ 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 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 transition metal.* This is the group iron, cobalt, and nickel. They show similarity in the horizontal direction.

- All three metals are ferromagnetic.

- 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.

- They form carbonyls with CO.

- All three metals have nearly the same melting point (about 1500°C).

- All three metals occur in nature together in the native state in the minerals awarait, Fe(Ni,Co)<sub>3</sub>, and josephinite, Fe(Ni,Co)<sub>2</sub>.

- *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* These are the metals between lanthanum and hafnium, namely cerium to lutetium (Figure 4). Although they have two electrons in the outermost shell, and one would expect that they would form divalent compounds, yet their most common valency state is three. This is one of the exceptions in the Periodic Table.

- Beside showing multiple valency they also form colored ions in solution (Table 4).



**Figure 4.** The actual position of lanthanides in the Periodic Table

**Table 4.** Colour of inner transition metal ions in solution

	$M^{2+}$	$M^{3+}$	$M^{4+}$	$MO_2^{2+}$
Ce	—	colourless	orange red	—
Pr	—	yellow green	—	—
Nd	—	red violet	—	—
Sm	red brown	pale yellow	—	—
Eu	—	pale pink	—	—
Gd	—	colourless	—	—
Lu	—	colourless	—	—
Yb	green	colourless	—	—
Tm	—	pale green	—	—
Er	—	rose	—	—
Ho	—	brownish yellow	—	—
Dy	—	pale yellow green	—	—
Tb	—	pale pink	—	—
Th	—	—	colourless	—
U	—	—	green	yellow
Pu	—	green	pink	—

- Being less reactive than the typical metals, they are so similar in chemical properties that their separation is usually done by making use of differences in physical properties.

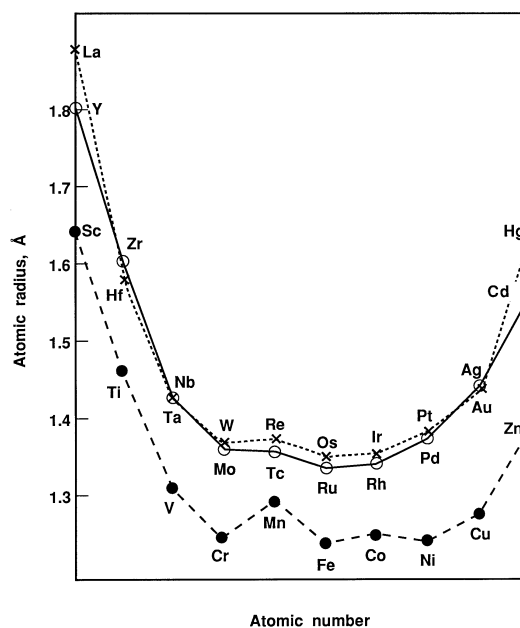
- They form ionic carbides that decompose with water liberating hydrogen. However, these carbides have some metal like character, e.g., they have metallic luster and electrical conductivity.

- As in typical and transition metals, the atomic radius in the lanthanides gradually decreases from lanthanum to lutetium with the result that although the atomic radius of lanthanum is larger than that of yttrium, which is above it in the Table, the atomic radius of lutetium is smaller than that of yttrium, i.e., La, Y, Lu. This is known as the "lanthanide contraction" and is responsible for the fact that the atomic radii of the metals lanthanum to mercury are nearly identical to those above them in the Table (Figure 5).

#### The actinides

These are the metals following actinium, namely thorium to lawrencium. Of this only thorium, uranium, and plutonium are of practical importance. In this group the 6d and 5f energy levels of electrons may be too close to each other which results

in cases where f-levels may be occupied in one atom but not in another. Uranium has multiple valency (+3, +4, +5, and +6) and form colored compounds; plutonium exists in +3 and +4 oxidation states and also forms colored compounds (Table 4).



**Figure 5.** Lanthanide contraction. The lanthanides (metals between La and Hf) are not included in this chart.

- The actinides differ from the lanthanides by the fact that they can be readily separated one from the other by differences in the chemical properties while this is not the case with the lanthanides.

- They are similar to the lanthanides in forming ionic carbides with some metal-like behavior that decompose with water yielding hydrogen.

- While plutonium shows similarity to samarium, thorium and uranium differ much more than do cerium and neodymium. While thorium is like cerium in the tetravalent state, uranium is most stable in the hexavalent state. Thus they are, in this respect, more like transition metals than lanthanides. Plutonium is most stable in the tetravalent state, but also occurs in the +3 and +6 valent states.

- Americium and curium do show similarity and difficulty of separation comparable to the lanthanides.

## DISCUSSION

Since its inception in 1869 by Mendeleev, the Periodic Table has undergone numerous changes (Table 5) and a large number of forms have been proposed.

#### IUPAC's Periodic Table

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.<sup>1,2</sup> This numbering was a compromise between the North

American and the European Periodic Tables (Figure 6). But this raised a hot debate among chemists that was published in *Chemical & Engineering News* in a series of letters to the editor.<sup>3-5</sup> Almost all these letters opposed the new numbering system.

Figure 6 displays three versions of the periodic table illustrating different group numbering systems:

- North American numbering:** Groups are labeled IA through VIIIA, with a gap between VIIIB and IB.
- European numbering:** Groups are labeled IIA through VIIIB, with a gap between VIIIA and IIB.
- IUPAC numbering:** Groups are labeled with numbers 1 through 18.

**Figure 6.** The Periodic Table according to IUPAC in an attempt to resolve the difference between European and North American nomenclature regarding the sub-groups

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 7) 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 inner transition metals (the lanthanides and the actinides), the metalloids, the monatomic nonmetals, and the covalent nonmetals.

Figure 7 shows a proposed periodic table without group numbers, color-coded by element type:

- METALS:** Typical Metals (Li, Be), Transition Metals (Sc to Zn), Less Typical (Ga to Hg), and Inner Transition Metals (Lanthanides and Actinides).
- METALLOIDS:** Elements like B, Si, Ge, As, Sb, Te, and Po.
- NONMETALS:** Covalent (C, N, O, F, S, Se, Br, I, Xe) and Monatomic (H, He).

**Figure 7.** A proposed Periodic Table without group numbers

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

□ The group starting with boron and ending with thallium (Group 13 of the IUPAC notation) is in fact composed of three different groups: boron is a metalloid, aluminum is a typical metal, gallium, indium, and thallium are less typical metals.

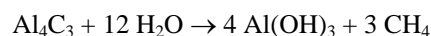
□ The group starting with carbon and ending with lead (Group 14 of the IUPAC notation) is again composed of three different groups: Carbon is a nonmetal; silicon and germanium are metalloids while tin and lead are less typical metals.

There is a little advantage in keeping these two groups as Groups 13 and 14 (or III A and IVA) - they have little in common. There is a great disparity in the melting points of the members of both groups. Similarly, the nitrogen and oxygen groups are a mixture of nonmetals and metalloids; the differences are more than the similarities. It would be advantageous to make this distinction. It would also be advantageous to identify the iron, cobalt, and nickel as a group, and the platinum metals as another group, instead of putting them in three vertical groups 8, 9 and 10.

### 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 aluminum cannot.
- $\text{Al}(\text{OH})_3$  does not dissolve in ammonium hydroxide solution, but  $\text{Ga}(\text{OH})_3$ , does dissolve.
- Gallium is precipitated from aqueous solution by  $\text{H}_2\text{S}$  as a sulfide; aluminum does not.
- Aluminum forms carbides, gallium does not.
- Aluminum carbide,  $\text{Al}_4\text{C}_3$  is similar to the carbides of the first three groups, being ionic colorless compound containing the  $\text{C}_2^{2-}$  anion that decomposes in water. However, it liberates methane instead of acetylene like the other members of the group:



- Gallium forms a gaseous hydride,  $\text{Ga}_2\text{H}_6$ , while aluminum forms a white solid polymer hydride,  $(\text{AlH}_3)_x$
- Historically, aluminum oxide was considered an earth like the rare earths and its 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  $\text{Al}^{3+}$ ,  $\text{Sc}^{3+}$ ,  $\text{Y}^{3+}$ ,  $\text{La}^{3+}$ , and  $\text{Ac}^{3+}$  form a series similar to  $\text{Mg}^{2+}$ ,  $\text{Ca}^{2+}$ ,  $\text{Sr}^{2+}$ ,  $\text{Ba}^{2+}$  and  $\text{Ra}^{2+}$ .

- Gallium is a typical dispersed element whose relative abundance in the Earth's crust is  $1.5 \times 10^{-3}$  %, aluminum is the third most abundant element with a relative abundance of 8.13 % (after oxygen and silicon).

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 two s and one p electrons, while the others have two s and one d 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  $\text{M}(\text{OH})_n$  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

Chemists should abandon 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 (iron, cobalt, nickel)
- Transition metals with vertical and horizontal similarity (platinum group metals)
- Inner transition metals:
  - The lanthanides
  - The actinides

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

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