

A NEW LOOK AT THE PERIODIC TABLE

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Abstract: Scientific knowledge expands as the number of laws it establishes decreases. Here, the Periodic Law—the cornerstone of the Periodic Table—offers a significant advancement for the field of chemistry as it allows for the elements to be naturally classified. More than a century ago, scientists created the Periodic Table to organise the characteristics of the elements. Physicists acknowledged it as a rule of nature after learning about the atomic nucleus. It became an indispensable tool for metallurgists, chemists, and physicists as the study of solid crystalline structure,

chemical bond nature, and metal theory was advanced. The 87 elements that exist in nature are divided into three categories: metals (63 elements or almost 3/4 of the total), nonmetals (16 elements), and metalloids (9 elements). Instead of assigning numerical values to the Periodic Table groupings, chemists should give them descriptive names.

Keywords: metals, non-metals, metalloids, typical metals, less-typical metals, transition metals, inner transition metals, lanthanides, actinides.

INTRODUCTION

Metals are ubiquitous; most often, they are found as alloys, which are mixtures of many metals. The metallurgical industry cannot function without them. With the exception of carbon, the ordinary human seldom makes use of nonmetals. Air, a gas composed of mostly inert elements, is real yet invisible to the naked eye. The chemical industry relies on nonmetals. The electronics sector and cutting-edge technologies both rely on metalloids (Figure 1). Crystals of densely packed atoms with their outer electrons so loosely bound that they may travel freely across the crystal lattice make up metals in their solid form. This arrangement clarifies their chemical, physical, and mechanical characteristics.

Hydrogen, oxygen, nitrogen, fluorine, chlorine, and the inert gases 1 and 2 are examples of nonmetals. Liquid bromine and solid carbon, sulphur, phosphorus, and iodine are also nonmetals. These substances lack the characteristics that are associated with metals. Electrons are easily shared by nonmetals with the exception of inert gases. Atoms that share their outer electrons form a covalent bond, which joins them together. Diatomic molecules like H₂, Cl₂, and N₂ are common, as are bigger molecules like P₄ and S₈, and huge molecules like graphite and diamond, which are carbon networks with an infinitely enormous volume, are also common.

1Gases in the periodic table, beginning with helium and ending with radon, were once known as inert gases. Due to the scarcity of xenon-fluorine compounds, they were thought to be inert until the 1960s.



Figure 1. The periodic table

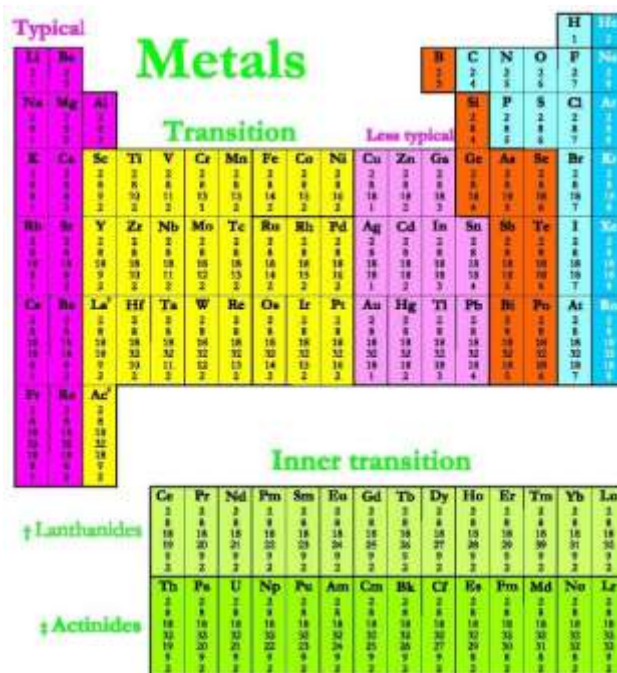


Figure 2. Electronic configuration of the metals.

Table 1. General characteristics of metals, metalloids, and nonmetals.

Metals	Metalloids	Nonmetals
Crystalline solids (except mercury) with metallic lustre Do not readily share electrons, their vapours are monoatomic	May be crystalline or amorphous sometimes have metallic lustre Readily share electrons even in the elemental form	Form volatile or non-volatile molecules having no metallic lustre 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 Have high density and useful mechanical properties	Low electrical and thermal conductivity Moderate density, no useful mechanical properties	Do not conduct electricity or heat. Electrical resistance decreases with increased temperature Low density of no useful mechanical properties
Electropositive, form cations, e.g., Cu^{2+} , Na^+ , etc. Form basic oxides, e.g., CaO Deposit on the cathode during electrolysis Either form no compounds with hydrogen or form unstable compounds usually nonvolatile (metal hydrides)	Sometimes electropositive, sometimes electronegative Form acidic oxides Deposit on the cathode Form stable compounds with hydrogen, e.g., AsH_3 , H_2Se	Electronegative, form anions, e.g., S^{2-} , Cl^- , etc. Form acidic oxides, e.g., SO_2 Deposit on the anode, e.g., O_2 , Cl_2 Form stable compounds with hydrogen, usually volatile, e.g., NH_3 , PH_3 , H_2S , etc.

Metalloids are a kind of material that has characteristics of both metals and nonmetals, with the covalent bonding characteristic of nonmetals. All three types of materials have their characteristics outlined in Table 1.

CLASSIFICATION OF METALS

Metals may be categorised into four types: typical, less typical, transition, and inner transition. This is because metals are elements that can lose electrons. Their electrical nature causes this divide (Figure 2).

Typical metals

Alkali metals, alkaline earths, and aluminium are all part of this category. The following are some of their traits:

- Their electrical structure is comparable to that of inert gases, with an outer shell consisting of one, two, or three electrons.

- Their outermost electrons are lost in a single step, giving them single valency.

Their reactivity means that they react easily with oxygen and water. The need to acquire the electronic structure of an inert gas in order to achieve maximal stability is the motivating reason behind this reactivity. The quick formation of a protective oxide covering on the surface of reactive metals like magnesium or aluminium makes them suitable building materials.

All the chemicals they produce are colourless.

- Due to the addition of electron shells, the atomic radius grows with increasing atomic number within a certain vertical group.

- The reactivity increases as the atomic number rises within a certain vertical group due to the fact that the outermost electrons are more easily lost when they are farthest from the nucleus. Rubidium is more reactive than potassium, cesium is more reactive than rubidium, etc.

Reactivity drops when the nucleus charge rises because more electrons are attracted to it electrostatically, making it harder to lose the outermost electrons. As a result, salts like magnesium and potassium are more reactive than calcium and magnesium, respectively.

Atoms shrink in size due to an increase in the nucleus's charge, which increases the electrostatic attraction between the electrons. Consequently, the radius of aluminium is smaller than that of magnesium, and magnesium is smaller than that of sodium.

- The density of an atom grows as its radius shrinks and its atomic weight rises, since an atom becomes denser as it gets heavier. The density of aluminium is lower than that of magnesium, while magnesium is lower than that of sodium.

- With the exception of beryllium and aluminium, they may form compounds with mercury due to their significant solubility in the metal.

Less typical metals

The following metals are included in this group: lead, zinc, cadmium, copper, indium, gallium, thallium, and zinc. They deviate from the electronic structure of inert gases, which is what sets them apart from ordinary metals. Instead of eight electrons in the inert gas structure, there are as many as 18 in the next innermost shell, and the outermost shell may contain as many as four. These traits are a direct consequence of their electrical make-up:

- With 18 electrons in a single shell, the atomic radius is less than that of ordinary metals in the same horizontal group, leading to an enhanced

connection between the nucleus and electrostatic forces. So, zinc has a smaller atomic radius than potassium, silver has a smaller atomic radius than rubidium, and gold has a smaller atomic radius than cesium. The atomic radii of gallium, zinc, and copper are all greater than that of copper because, contrary to what is expected for metals, the atomic radii of these elements grow as the number of electrons in their outermost shells rises. Figure 3 illustrates this: As the atomic number increases, the atomic volume of common metals decreases, while for rare metals, the opposite is true. The 18-electron shell's shielding effect, the greater repulsion between the outermost electron and that shell, and the higher repulsion of the electrons inside that shell all contribute to this phenomenon.

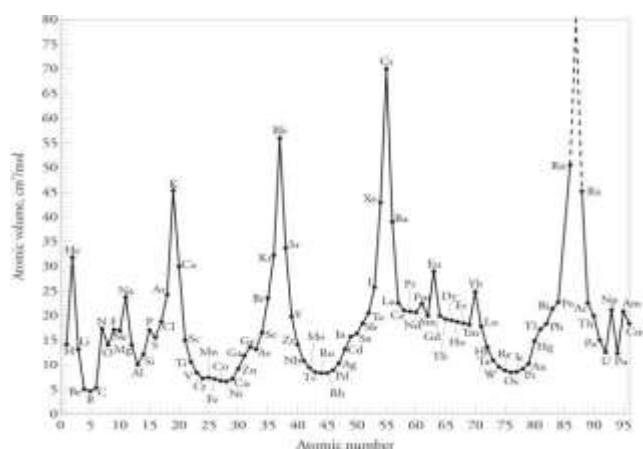


Figure 3. Atomic volume of elements.

- These metals are less reactive than their counterparts in normal metals for two reasons: first, the outermost electrons will not be readily lost.

Since the achievement of an inert gas electrical structure is not going to happen, there is no need to let electrons escape.

Because its atomic radius is less than that of other metals, its electrostatic attraction is greater.

- Denser than their normal metal counterparts, these metals have a greater atomic weight and a smaller atomic radius.

- Different valency states are shown by some of these metals; for instance, CuI and CuII represent copper, AuI and AuIII represent gold, HgI and HgII represent mercury, SnII and SnIV represent tin, and PbII and PbIV represent lead. This is due to the fact that the 18-electron shell is susceptible to loss of one or two electrons.

The volume, in cubic centimetres, that a solid gramme of an element occupies is the atomic volume. Since there is a constant number of atoms in one gramme of atomic weight, it may be used as a qualitative reference to the relative volumes of the individual atoms.

- A small number of these metals exist as coloured ions in solution, such as CuII and AuIII, or as coloured compounds, such as blue copper sulphate pentahydrate, yellow cadmium sulphide, etc. (Table 2). This is because there's a chance that electrons may jump from the lowest level of the 18 electron shell to the highest level.

- Their electrical structure is very similar to mercury's, which is why they dissolve in mercury at a rapid rate. They also don't react with mercury to create a compound.

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

	M ⁺	M ²⁺	M ³⁺	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

Metals ranging from scandium to nickel are included in the Periodic Table's vertical groupings. Their electrical arrangement differs from that of inert gases, and they are defined by an even distribution of electrons across their shells, with an increasing density in the innermost shell. On the other hand, the quantity of electrons in the outermost electron shells seems to be non-normal. This is because of energy levels, which are found by measuring spectra. The characteristics of transition metals fall in between those of more common and less common metals, as the name suggests.

Because they do not attain the inert gas structure when they lose their outermost electrons, these metals are less reactive than usual metals; yet, they are more reactive than less typical metals. Some characteristics that both of them have are:

They have the same amount of outermost electrons, which makes them quite similar to one other and shows the normal group interactions.

μIons with larger charges may be formed if they lose more electrons from the lowermost shell. Because of this, their valence might vary. As an example, titanium may be found in +2, +3, and +4 oxidation levels, whereas vanadium can be found in +2, +3, +4, and +5 oxidation states.

Adding one electron to an inner shell marginally reduces the size of the atom due to increased electrostatic attraction, which causes the atomic radius of the subsequent metals in a specific horizontal period to drop slightly as the atomic number increases.

With the exception of the groups Sc, Y, La, and Ac, which exclusively produce colourless compounds, the majority of these substances produce coloured ions in solution as a result of electronic transition (Table 3).

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

	M ²⁺	M ³⁺	MO ²⁺	MO ₂ ⁺	MO ₂ ²⁺
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	—	—	—
Pb	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	—	—

Their covalent compound formation includes several examples, such as the carbonyls of iron and nickel, the chlorides of titanium, and the oxyacids of chromium, molybdenum, and tungsten.

For instance, the amines of nickel and cobalt may be formed into coordination compounds with ammonia.

Typically, they produce borides, carbides, nitrides, and hydrides, most of which are metallic in nature.

When it comes to mercury, they are the least soluble.

Three classes of transition metals exist:

Metals that undergo a vertical transition. Scandium and manganese form vertical groups. Zr-Hf, Nb-Ta, and Mo-W are a few examples of such compounds that exhibit vertical resemblance. All of the compounds formed by the Sc, Y, La, and Ac groups are colourless and have a valence of 3.

- Metal undergoing horizontal transition. This is the group that includes nickel, cobalt, and iron. With respect to the horizontal axis, they are comparable.

"Ferromagnetic" describes all three metals.

Carbides of these metals exhibit characteristics that lie somewhere in the middle, between those of ordinary ions and the metal-like nature of transition metals. So, they're shiny and

conduct electricity, but water and weak acids dissolve them.

The chemicals they produce are divalent and trivalent.

Carbonyls are formed when they react with carbon monoxide.

The melting point of all three metals is quite close to 1500 degrees Celsius.

- Awarait and josphinite are minerals that naturally contain all three elements in their native states: Fe(Ni,Co)₃ and Fe(Ni,Co)₂, respectively.

- Transition metals having a horizontal to vertical orientation. All six of these metals belong to the platinum group, and their similarities run both horizontally and vertically.

- Corrosion does not affect them.

In their natural habitat, they coexist.

Inner transition metals

Every inner shell of these metals has an increasing amount of electrons, yet the number of electrons in the two outermost shells is equal. Two categories make them up:

They are called lanthanides. The metals cerium through lutetium, which fall between lanthanum and hafnium, are shown in Figure 4. Despite the fact that they are expected to form divalent compounds due to their two electrons in the outermost shell, their most frequent valency state is actually three. Among the periodic table's exceptions, this one stands out.

Table 4 shows that these compounds exhibit multiple valency and the formation of coloured ions in solution.

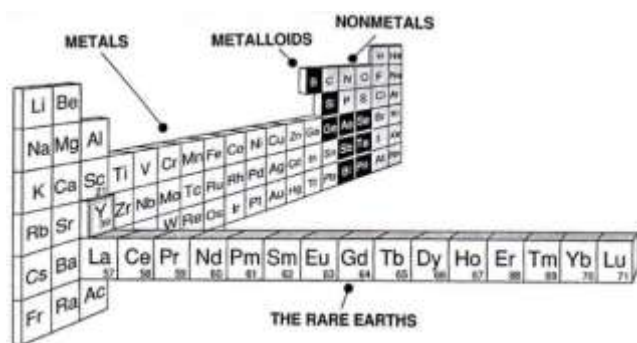


Figure 4. The actual position of lanthanides in the Periodic Table

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	—

- Because of their low reactivity and similarity in chemical characteristics, physical property differences are often used for their separation from more reactive metals.
- When exposed to water, they break down into ionic carbides, which release hydrogen. Having said that, these carbides do exhibit some metal-like properties, such as electrical conductivity and a metallic sheen.
- Just like in ordinary and transition metals, the atomic radius of lutetium (La, Y, Lu) is smaller than that of yttrium (La, Y, Lu), even though lutetium's atomic radius is larger than yttrium's (right above it in the table), as is the case with lanthanum (La) and other lanthanides. A phenomenon called the "lanthanide contraction" is to blame for the striking similarity between the atomic radii of mercury and the metals above it in the Table (Figure 5).

The insulimides

Thorium through lawrencium are the metals that come after actinium. Thorium, uranium, and plutonium are the only ones that matter in a practical sense. Potentially too close proximity of the 6d and 5f electron energy levels in this set leads to

with atoms that may have occupied f-levels but unoccupied ones. The valency of uranium may be +3, +4, +5, or +6, and it can also create compounds with different colours. Similarly, plutonium can exist in +3 and +4 oxidation states, and it can also form compounds with different colours (Table 4).

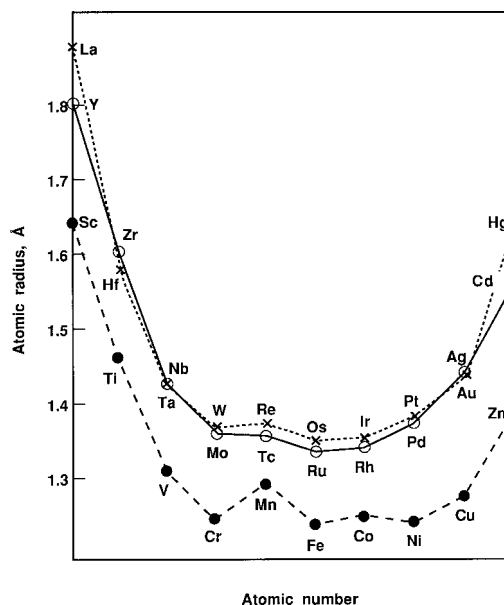


Figure 5. Lanthanide contraction. The lanthanides (metals between La and Hf) are not included in this chart.

The actinides are distinct from the lanthanides because their chemical characteristics allow for easy separation, whereas this is not the case for the lanthanides.

- They break down in water to produce hydrogen, much like the lanthanides, and they generate ionic carbides that behave much like metals.

- Although plutonium and samarium are comparable, thorium and uranium are very different from cerium and neodymium. Unlike thorium, which is essentially cerium in its tetravalent form, uranium is stablest in its hexavalent form. That is why they resemble transition metals more than lanthanides in this regard. While +3 and +6 valent states of plutonium do exist, the most stable of these is tetravalent.

Comparable to the lanthanides, ammonium and curium exhibit resemblance and separation difficulties.

DISCUSSION

There have been many iterations and revisions to the Periodic Table since its creation in 1869 by Mendeleev (Table 5).

IUPAC's Periodic Table

Groups 1–18 of the Periodic Table were proposed by the International Union of Pure and Applied Chemistry (IUPAC) in 1985; for example, Group 1 would include alkali metals while Group 18 would include inert gases. One, two As a middle ground, these numbers were acceptable to both the North

Figure 6 shows the periodic tables for the United States and Europe. However, this sparked a heated discussion among chemists, the results of which were documented in a string of editorials published in Chemical & Engineering News.³⁻⁵ The majority of these messages were against the new system of numbers.

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

If physicists and chemists give up on the traditional method of grouping them by number, they may be able to overcome this issue. A more practical approach would be to assign groups of elements names that represent their electrical structure, bonding type, and physical and chemical characteristics. In addition to helping pupils better comprehend the Periodic Table, these names should be simpler for them to recall. In this regard, the current categorization into 10 classes (Figure 7) ought to be useful. Aluminium is rearranged in this categorization to sit next to magnesium, creating a group of typical metals that also includes alkali metals and alkaline earth metals. Metalloids, monatomic nonmetals, covalent link nonmetals, inner transition metals (lanthanides and actinides), horizontal and vertical transition metals, less common metals, and transition metals with a vertical and horizontal component are the other categories.

Figure 7. A proposed Periodic Table without group numbers

Placing each group member in its proper place is one benefit of this categorization. Take this case in point:

Boron is a metalloid, aluminium is a common metal, and gallium, indium, and thallium are less common metals; these three groups make up Group 13 of the IUPAC nomenclature, which begins with boron and ends with thallium.

- There are three distinct groups that make up the group that begins with carbon and ends with lead (Group 14 in the IUPAC notation): Tin and lead are less common metals, silicon and germanium are metalloids, and carbon is not a metal at all.

Since these two groups are so dissimilar, it might be best to keep them separate and call them Groups 13 and 14 (or III A and IVA). The melting points of the two groups couldn't be more different. The nitrogen and oxygen groups are also composed of a variety of nonmetals and metalloids, however the commonalities outweigh the variances. Differentiating between the two would be beneficial. Instead of arranging the metals in three vertical groups 8, 9, and 10, it would be more beneficial to classify iron, cobalt, and nickel as one group and platinum as another.

The position of aluminum

Moving aluminum further away from gallium with which Some may oppose if it is found in bauxite. But keep in mind that the two metals are quite different from one another:

- Gallium does not oxidise as quickly as aluminium, which means it does not develop a non-porous protective layer as quickly.

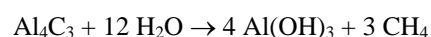
Electrodeposition of gallium from aqueous solutions is possible, but aluminum is not.

- While gadoxal (OH)₃ dissolves in a solution of ammonium hydroxide, al(OH)₃ does not.

- Hydrogen sulphide (H₂S) may remove gallium from water, while aluminium isn't affected.

- While gallium does not create carbides, aluminium does.

- Like the carbides in the previous three groups, aluminium carbide (Al₄C₃) is an ionic, colourless compound that breaks down in water due to the presence of the C²⁻ anion. Unlike the other members of the group, however, it releases methane instead of acetylene:



- While aluminium creates a white solid polymer hydride, (AlH₃)_x, gallium makes a gaseous hydride, Ga₂H₆.

• Aluminium oxide has a long history of being associated with the scandium group and rare earth elements.

• By connecting scandium with aluminium, Mendeleev accurately anticipated its characteristics prior to its discovery.

• Scandium has many chemical properties with aluminium; its ions, Al^{3+} , Sc^{3+} , Y^{3+} , La^{3+} , and Ac^{3+} , are in a series with Mg^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+} , and Ra^{2+} .

• In the Earth's crust, gallium is distributed and has a relative abundance of 1.5×10^{-3} %; aluminium, after oxygen and silicon, is the third most plentiful element with a relative abundance of 8.13%.

An further concern that might emerge from this transfer is the fact that aluminum's electron orbital structure differs from Sc, Y, and La. Specifically, aluminium has two s and one p electrons, whereas the other three have two s and one d electron. Here is a response to this objection. In the Periodic Table, elements are categorised as s-, p-, d-, or f-block elements based on the population of electrons in their subshells. However, this categorization fails to reveal or explain other phenomena, including:

• Elements of the boron group are all on the p-block, but they have distinct physical and chemical characteristics; for example, although most of the elements in the group are metals, boron is a metalloid. Additionally, aluminium hydroxide and boron hydroxide (boric acid) have structurally distinct differences. Unlike the latter, the former does not include any sharing of oxygen atoms among the $M(OH)_n$ groups.

• Elements of the Boron group have two sigma and one proton in their outer shell. With the exception of gallium and thallium to a lesser degree, all three electrons are lost in a single step. Conversely, all three electrons in the outer shell of an element in the scandium group—two positive and one negative—are lost simultaneously. Therefore, the total amount of electrons should be considered, not the difference in electron orbitals, as a significant requirement for element categorization.

• The d-block, which includes transition metals like zinc and copper, is unique among its members in that, unlike the rest of the block, the atomic radius grows as the atomic number rises, rather than the other way around. Therefore, they do not belong in the category of transition metals since their chemical characteristics are distinct.

• The electron orbitals in the subshells are very distorted due to quantum mechanical factors. Take this case in point:

The d-electrons of chromium, manganese, molybdenum, and technetium are equal, but the s-electrons of these elements are different, rather than the other way around.

Instead of the other way around, gadolinium and europium have distinct d-electrons but comparable f-electrons.

The entire amount of electrons in a shell should be considered instead of rigidly adhering to the s-p-d and f-classifications, as this makes them useless.

Also, physicists, who aren't always concerned with chemical characteristics, are the ones who came up with this categorization system.

CONCLUSIONS

Instead of assigning numerical values to the Periodic Table groupings, chemists should use the following more descriptive names:

Monoatomic elements that are not metals

nonmetals that are covalent

• Microorganisms

• Common metals

• Fewer common metals

• Metals that undergo a transition and have a vertical grain

• Iron, nickel, and cobalt, which are transition metals with horizontal similarities

• Metals of the platinum group that are transitional and similarly structured both vertically and horizontally

• Transition metals found within:

The lanthanides

• Actinides can

Placing each group member in its proper place is one benefit of this categorization.

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