

“Semi-Long-Periodical” Tabulated Version of Chemical Elements Periodic System

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Abstract: A possible modification of the tabulated version of the Chemical Elements Periodical System, proposed to overcome deficiencies of the two most used versions—the “short-periodical” proposed by Mendeleev and the “long-periodical” introduced by Werner and recommended by IUPAC at the present time—is discussed in this paper.

Keywords: chemical elements periodical system; short-periodical version; long-periodical version; semi-long-periodical version

1. Introduction

One of the key problems of modern chemistry is the systematization of chemical elements. This seems quite natural because all known chemical substances are composed of these elements (or rather, from their atoms). The basis of the chemical elements’ systematics is well known to be the Periodic Law of D.I. Mendeleev. Its most rigorous modern formulation is as follows: “Most of the chemical element properties are in close to periodic dependence on the charge of their corresponding nuclides (atomic nuclei)”. The peculiarity of this law, distinguishing it from many other physics laws (the Periodic Law belongs precisely to this field of science—in any case, it is included among its 500 laws), is that a strict mathematical formula has not yet been found in which these very properties (at least the most important of them) would be quantitatively related to the element’s serial number in the Periodic System (that numerically equal to the atomic nucleus charge). Therefore, over the approximately 150 years that have passed since the discovery of the Periodic Law, many different versions of its interpretation have been proposed, such as various so-called “Periodic Tables”. At different times, other forms were also proposed; however, they did not receive any significant distribution among professional chemists or other researchers. The earliest was the tabulated form proposed by Mendeleev himself in his classical works [1,2]. On this basis, after his death, the current so-called “short-periodical version” (Table 1), which subsequently entered all chemistry textbooks and was dominant for a long time, was created. In this version, all currently known chemical elements (118 items) are divided into eight groups and seven so-called periods (while the last of them, the seventh period, has not yet been completed). Shortly before the death of D.I. Mendeleev, another great scientist, the Swiss chemist A. Werner (later Nobel Laureate in Chemistry) in 1905 proposed a slightly different interpretation of the Periodic Law, called the “long-period version”, within which there were already 18 groups of chemical elements in the same seven periods (Table 2). Without dwelling here in detail on the specifics of these and other tabulated forms, with chemical elements and the determination of their locations, we note that in the framework of the modern theory of the atom structure, all elements are divided into four categories depending on the structure of the outer electron shells of their atoms, the electrons of which are capable of participating in the formation of the chemical bond. That is, the chemicals are categorized based on the *s*-elements (ns^k , $k = 1$ or 2), *p*-elements (ns^2np^k , $k = 1-6$), *d*-elements ($ns^2(n-1)d^k$, $k = 1-10$), and *f*-elements ($ns^2(n-1)f^k$, $k = 1-14$), where *s*, *p*, *d*,



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and f are symbols of atomic orbitals (AO) on which electrons are arranged, n is the principal quantum of the corresponding AO, and k is the number of electrons in the given AO.

Table 1. The current short-periodical version of the Chemical Elements Periodic System. The numerical indices before the symbols of chemical elements are the charges (Z) of the nuclides corresponding to these elements. *s*-, *p*-, *d*-, and *f*-elements are highlighted in red, yellow, blue, and green, respectively.

	I	II	III	IV	V	VI	VII	VIII																
1	¹ H						(¹ H)	² He																
2	³ Li	⁴ Be	⁵ B	⁶ C	⁷ N	⁸ O	⁹ F	¹⁰ Ne																
3	¹¹ Na	¹² 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 Gd	⁴⁹ In	⁵⁰ Sn	⁵¹ Sb	⁵² Te	⁵³ J	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 Db	¹⁰⁶ Sg	¹⁰⁷ Bh	¹⁰⁸ Hs	¹⁰⁹ Mt	¹¹⁰ Ds														
	¹¹¹ Rg	¹¹² Cn	¹¹³ Nh	¹¹⁴ Fl	¹¹⁵ Mc	¹¹⁶ Lv	¹¹⁷ Ts	¹¹⁸ Og																
⁵⁸Ce												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
⁹⁰Th												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

Table 2. The current long-periodical version of the Chemical Elements Periodic System. The numerical indices before the symbols of chemical elements are the charges (Z) of the nuclides corresponding to these elements. *s*-, *p*-, *d*-, and *f*-elements are highlighted in red, yellow, blue, and green, respectively.

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
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	¹² 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 Gd	⁴⁹ In	⁵⁰ Sn	⁵¹ Sb	⁵² Te	⁵³ J	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 Db	¹⁰⁶ Sg	¹⁰⁷ Bh	¹⁰⁸ Hs	¹⁰⁹ Mt	¹¹⁰ Ds	¹¹¹ Rg	¹¹² Cn	¹¹³ Nh	¹¹⁴ Fl	¹¹⁵ Mc	¹¹⁶ Lv	¹¹⁷ Ts	¹¹⁸ Og
			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		
			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		

Both of these versions, however, have at least three disadvantages. Firstly, both the first and the second f -elements (so-called “lanthanoids” and “actinoids”) are actually not in the table itself but are taken out of it in the form of some “applications”, and there are no convincing arguments to justify such a decision from the standpoint of the theory of the atom structure. Secondly, with any of these two possible (and accepted) versions for the position of the first of all chemical elements—hydrogen (in the first or in the seventh (within the short-periodical version) or seventeenth (within the long-periodical version))—there are six (!) and 16 (!!) empty cells, respectively, and it is not clear how to fill them.

Along with these two drawbacks, each of these options has a third drawback. In the short-periodical version, as it can be noticed, *d*-elements within each of those periods where they are present (namely, starting from the fourth) are actually located on “two floors” (Table 1). This shortcoming in the long-periodical version is eliminated, but the table becomes less compact compared to the table in the short-periodical version. However, this is still half the trouble. Much worse is that the key table parameter, i.e., the group number, has completely lost its physical meaning; in the short-period version, it is in most cases associated with the numerical value of the maximum oxidation degree of the corresponding element (for example, aluminum and yttrium are in group III, and their maximal degree of oxidation is +3; vanadium and bismuth are in the group V, and their maximum oxidation state is +5; and manganese and iodine are in group VII and their maximum degree of oxidation is +7), even though there are exceptions to this rule (for example, Cu, Ag, and Au are in group I, but their maximum of oxidation degrees are +4, +3, and +5, respectively). In the long-period variant, such accordance takes place only for *s*-elements and *d*-elements of groups III–VII. In the long-periodical version, as can be easily noticed, such an accordance takes place only for *s*-elements and *d*-elements of groups III–VII. Of course, by developing A. Werner’s approach logically, it is possible to neutralize the above-mentioned first drawback to place *f*-elements in the table in the same way; however, then the number of groups will become 32, which means that the table will eventually become much wider. It should be noted that the current number of known chemical elements (118) is by no means the end of their “world”. Another category of them is on the way, terra incognita for scientists, which includes 18 5g-elements. The first of them, in accordance with the modern theory of the atom structure, should have a nuclear charge $Z = 121$, and the latter a nuclear charge $Z = 138$. In addition, after their discovery, the number of groups in the long-periodical version will increase to 50 (!). This is not very user-friendly. The short-periodical version is more convenient in this regard, because when 5g-elements appear, they, like *f*-elements, can be placed in the form of similar “applications” that will only slightly expand it. However, the short-periodical version of the Periodic System of D.I. Mendeleev was officially canceled by the International Union of Pure and Applied Chemistry (IUPAC) in 1989; therefore, the long-periodical version is accepted in the modern literature of most countries. One of the few exceptions (and perhaps the only one) is Russia, where both versions are used, and in the educational process and among students, the short-periodical version is the most popular.

After all of the above, the problem of creating such an interpretation of the Periodic Law, where the above shortcomings would be avoided, becomes quite real. As practice has shown, the two-dimensional tabulated form, and not any other two-dimensional (for example, spiral, stair, or concentric) or three-dimensional (pyramidal, spherical, cylindrical, etc.) is the most convenient to use among all the possible geometric shapes associated with the interpretation of this law. Therefore, the most optimal way to solve this problem is to modify the generally accepted tabulated versions of the Periodic System indicated above. A possible solution to this problem will be presented in this article.

2. Methodology

In the short-periodical version, the location of any chemical element is determined by two parameters, namely with the nuclear charge Z , coinciding with the element atomic number, and its maximal possible degree of oxidation. This approach is quite well justified, and as can be seen from Table 1, the system of chemical elements as a whole turns out to be harmonious and logical. However, the maximal oxidation degree of a chemical element is only one of its properties, which, like many others, is only a manifestation or consequence of the same CAUSE—the electronic structure of the chemical element atoms. Thus, as one of the characteristics determining the position of a particular chemical element in the Periodic System, we choose in all cases the property that is the CONSEQUENCE of this very CAUSE. Such an approach, obviously, in itself does not seem quite logical.

As is known, D.I. Mendeleev revealed the fact of a regular change in the chemical elements with a change in atomic masses and formulated the Periodic Law, but he was not able to explain a reason for periodicity. Additionally, this was not because of a lack of talent or imagination. The point is that to understand the given phenomenon's nature, it is necessary to have a rigorous theory of the atom structure, but it began to develop only in the second decade and acquired a fairly complete form in the 1920s when Mendeleev was no longer alive. Additionally, therefore, building his system, and not having data on the cause of the phenomenon, he was forced to focus on that consequence (in the given case, some property of the chemical element), in light of the fact that the periodic repetition of their properties itself would look the most distinct. Such a property in his time turned out to be the maximal valency, although even when it was used as a second parameter (along with atomic mass) when constructing a system of chemical elements, it was not without certain costs. (At present, the concept of "valence" due to the ambiguity of interpretation of this term was replaced by a more adequate concept of "oxidation degree", which, as indicated in a number of our works, in particular [3,4], "is the charge in units of electron charge that would occur on the atom of a given element in a given chemical compound under the assumption that within the framework of each of the conditionally-existing, in this compound, two-center two-electron chemical bonds formed by the exchange mechanism, there would be a complete transfer of electrons from the atom with less electronegativity to the atom with more electronegativity".) However, we have one very significant advantage over D.I. Mendeleev: we know the CAUSE of periodicity and its violations, which are directly connected with the specifics of the filling of atomic orbitals (AO) with electrons and, accordingly, the electronic configuration of atoms, taking into account both the number of electrons in the outer layer and the range of AOs they are located in. It is such a characteristic that should be taken as the second parameter in constructing a system of chemical elements.

One should probably start creating a new tabulated version of the Chemical Elements Periodic System by establishing the place in this system of an element with $Z = 1$, i.e., a hydrogen atom. As is known, within the framework of the short-periodical variant, its position is ambiguous: some researchers consider it expedient to place it in the first group (together with alkali metals or ns^1 -elements), while others prefer to place it in the seventh group (together with halogens or ns^2np^5 -elements). For each of these two approaches, there are arguments "for" and "against", because for some properties, hydrogen is similar to alkali metals, and for others it is similar to halogens, and an unambiguous choice between them within the framework of the short-period version (as well as within the framework of the long-period version) does not seem possible. Therefore, we will try to consider the current situation from a different angle.

Let us place hydrogen (H) at the very beginning of our table in its upper left corner, and after it, without any gaps, the following elements in order of increasing nuclear charge Z : He, Li, Be, B, C, N, etc. Having carefully looked at their properties, we will not find elements similar to each other, neither in chemical properties nor, more importantly, in the structure of the outer electron shells, up to the element with $Z = 8$ (oxygen O) inclusive. However, the element with $Z = 9$ (fluorine F), following oxygen in terms of the structure of the electron shell, CAN already be considered as an ANALOGUE of H, because despite the fact that their electronic configurations of the outer shells differ significantly from each other ($1s^1$ for H and $2s^22p^5$ for F), both of these elements lack one electron each before filling the outer electron shell. The next element after fluorine, Ne, with $Z = 10$, is similar in its electronic structure ($2s^22p^6$) to He ($1s^2$) and, like it, has completely filled AO (although here there is still a difference in electronic configurations, albeit smaller than between hydrogen and fluorine). With a further increase in Z , the analogy in the electron shells' structure becomes complete: Li and Na— $2s^1$ and $3s^1$, respectively; Be and Mg— $2s^2$ and $3s^2$, respectively; B and Al— $2s^22p^1$ and $3s^23p^1$, etc. to S, which turns out to be an analog of O ($2s^22p^4$ and $3s^23p^4$, respectively). With this approach, it is possible to eliminate the common defect inherent in both the short-periodical and long-periodical versions of the

Periodic System; that is, one can eliminate those same empty cells between the first two chemical elements already mentioned above.

However, the question remains open as to which groups all the elements listed above should then be placed into. To answer it, let us now pay attention to the following fact. The third element in the Periodic System with $Z = 3$ (Li) has one electron in the outer layer, the element with $Z = 4$ (Be) has two, the element with $Z = 5$ (B) has three, etc., up to the element with $Z = 8$ (O), which has six electrons. Therefore, it is quite natural to place them into groups, giving them numbers corresponding to the NUMBER OF ELECTRONS ON THE OUTER LAYER. Then, Li should be positioned in group I, Be in group II, B in group III, C in group IV, N in group V, and O in group VI. So far, as we can see, there is a complete correspondence between the group numbers and the number of electrons in the outer layer within the framework of our approach. Going back, we note that an atom with $Z = 2$ (He) has a completely filled two-electron shell $1s^2$ and has NOT A SINGLE electron on its outer electron shell. In accordance with the above logic of numbering groups of chemical elements, it should be positioned in group 0, but not in group VIII, as it is in the modern short-periodical (and, all the more so, not in group XVIII, as it is in the long-periodical) version. It should be noted that D.I. Mendeleev, in HIS original version of the Periodic Table, placed helium, as well as all other elements similar to it (the so-called inert gases) with the configuration ns^2np^6 , in the group 0 and was categorically against placing them in group VIII [1,2]. However, the first element of the Periodic System, hydrogen (H), as mentioned above, lacks one electron before filling the outer electron shell, and now, in full accordance with the aforesaid logic, it should be positioned in the minus first group (–I). This conclusion seems paradoxical, but this decision is precisely what allows us to completely eliminate all empty cells in the first period. Naturally, halogens also fall into the same group (–I), the atoms of which also lack one electron until the completion of their outer shells.

Continuing the “settlement” of chemical elements in the above table, we can easily place in it all the currently known chemical elements within the framework of the principle “group number = number of electrons in the outer shell”. Some difficulties, however, arise with Co and Ni, as well as their analogs that have similar electronic configurations $(n - 1)d^7ns^2$ and $(n - 1)d^8ns^2$, namely Rh and Pd, for the simple reason that the numbers of electrons in the valence levels of them are 9 and 10, respectively. In the current version of the Chemical Elements Periodic System, as well as in its original interpretation by D.I. Mendeleev, they are placed together with Fe and its analog, Ru, in one large eighth group, which in itself looks unnatural [5]. On the other hand, the Co–Rh and Ni–Pd pairs have no analogs at all among all elements with $Z = 1–57$, and therefore, they should be placed in two independent groups, namely in IX and X, respectively. The same “population” principle can and should be used for f -elements. The first of them, Ce, with $Z = 58$, has a reduced electronic configuration of the outer AO $6s^25d^14f^1$ and four valence electrons, so it belongs in group IV, and not in III, where it is currently listed. Other $4f$ elements will be arranged similarly in the table: Pr with the $6s^24f^3$ configuration in V, Sm ($6s^24f^6$) in VIII, Tb ($6s^24f^9$) in XI, and so on, up to Er ($6s^24f^{12}$), which will occupy a cell in group XIV. The next element, Tm, having a $6s^24f^{13}$ configuration with the $4f$ -AO completely filled with electrons and one electron per $6s$ -AO, should, however, be placed in group I. Finally, the last two $4f$ elements, Yb and Lu, will occupy the corresponding cells in groups II and III. After all, placing the rest of the currently known chemical elements with $Z = 72–118$, namely categories $5d$, $6p$, $7s$, $5f$, $6d$, $7p$, and $8s$, as has been said above, will no longer be difficult. Moreover, our reasoning can be extended to those chemical elements that have not yet been obtained artificially or found in nature, namely, those forming the category $5g$ with $Z = 122–139$. It is worth noting that in this new format, the f - and g -elements are placed on par with the rest of the elements in their respective groups; moreover, even the concepts of “lanthanoids” and “actinoids”, referring to $4f$ - and $5f$ -elements, lose their original meaning. Additionally, outwardly, the new table looks more elegant compared to the short-periodical and long-periodical versions.

Concluding this section, it is necessary to explain the origin of the electronic configurations of the outer shells, which strictly speaking, is a consequence of one important regularity, namely the Madelung–Klechkovsky rule, including two provisions: (1) the filling of AO with electrons occurs in ascending order of the sum their main (n) and orbital (l) quantum numbers, and (2) if the indicated sum ($n + l$) is equal for two or more AOs, they are filled in the order of the growth of the main quantum number n . Using these rules, it is possible to derive the sequence of atomic orbitals in order of increasing energy indicated below:

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p < 8s < 5g < 6f < 7d < 8p < \dots$$

Thus, knowing the atom nucleus charge of the corresponding element Z and this sequence, it is possible with sufficient reliability to determine both the total atom electronic configuration in its ground state (i.e., the state with the lowest energy) and the reduced one, the electrons of which, in principle, may participate in the formation of chemical bonds.

3. Results and Discussion

The tabulated version of the Periodic Table of Chemical Elements with the above principle of “settlement” has the form presented in Table 3; it includes the first 216 chemical elements. This set is divided into four periods, the first of which contains 16, the second 36, the third 64, and the fourth 100 elements. Within the framework of this tabulated version (which can be called “semi-long-periodical”), the element number in various periods is different, distinguishing it from both the short-periodical and long-periodical versions (where, as is known, starting from the second period, there is pairwise equality between the numbers of elements in adjacent periods, namely the second and third (8 elements each), fourth and fifth (18 elements each), and sixth and seventh (32 elements each)). The total number of elements (S_E) in each period in this semi-long-periodical version is determined by the equation $S_E = 4(N + 1)^2$, where N is the period number. In each new period, a new category of chemical elements appears, the number of which is one greater than the number N . In the first period, there are two of these categories (s - and p -elements); in the second, three categories (s -, p - and d -elements); in the third, four categories (s -, p -, d - and f -elements); and in the fourth, five categories (s -, p -, d -, f - and g -elements). At the same time, the fourth period is so far represented by only two chemical elements, namely tennessine Ts ($Z = 117$) and oganesson Og ($Z = 118$), so in fact, here, we are dealing with three periods instead of seven, in each of the two above-mentioned main tabulated versions.

Obviously, in this version of the Periodic System, all the uncertainties associated with the hydrogen position (as well as empty cells between it and the next largest Z charge, helium) have disappeared. The halogens’ positions in this Periodical System, inert gases, lanthanides, and actinides have also become clearer and more definite.

However, it is much more important that, within the framework of this Periodic System version, new regularities are outlined, remaining either undiscovered or not clear enough in both the short-period and long-period versions. Below are some of these regularities.

Table 3. Semi-long-periodical version of the Periodic Table of Chemical Elements. The numerical indices before the symbols of chemical elements are the charges (Z) of the nuclides corresponding to these elements. s -, p -, d -, f -, and g -elements are highlighted in red, yellow, blue, green, and violet, respectively.

	- I	0	I	II	III	IV	V	VI	VII	VIII	IX	X	XI	XII	XIII	XIV	XV	XVI	XVII	XVIII
1	¹ H	² He	³ Li	⁴ Be	⁵ B	⁶ C	⁷ N	⁸ O												
	⁹ F	¹⁰ Ne	¹¹ Na	¹² Mg	¹³ Al	¹⁴ Si	¹⁵ P	¹⁶ S												
2	¹⁷ Cl	¹⁸ Ar	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	³⁷ Rb	³⁸ Sr	³⁹ Y	⁴⁰ Zr	⁴¹ Nb	⁴² Mo	⁴³ Tc	⁴⁴ Ru	⁴⁵ Rh	⁴⁶ Pd								
			⁴⁷ Ag	⁴⁸ Gd	⁴⁹ In	⁵⁰ Sn	⁵¹ Sb	⁵² Te												
3	⁵³ J	⁵⁴ Xe	⁵⁵ Cs	⁵⁶ Ba	⁵⁷ La	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	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	⁸⁷ Fr	⁸⁸ Ra	⁸⁹ Ac	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	¹⁰⁴ Rf	¹⁰⁵ Db	¹⁰⁶ Sg	¹⁰⁷ Bh	¹⁰⁸ Hs	¹⁰⁹ Mt	¹¹⁰ Ds								
4	¹¹⁷ Tc	¹¹⁸ Og	119	120	121	122	123	124	125	126	127	128	129	130	131	132	133	134	135	136
			137	138	139	140	141	142	143	144	145	146	147	148	149	150				
			151	152	153	154	155	156	157	158	159	160								
			161	162	163	164	165	166												
	167	168	169	170	171	172	173	174	175	176	177	178	179	180	181	182	183	184	185	186
			187	188	189	190	191	192	193	194	195	196	197	198	199	200				
			201	202	203	204	205	206	207	208	209	210								
			211	212	213	214	215	216												

In inorganic chemistry, one curious fact, which can be called the “effect of pairs”, has long been known. It consists of the following. According to the degree of proximity of their physicochemical characteristics, the elements of each of the groups with the same electronic configuration are divided into pairs, and within each of these pairs there is a significantly greater similarity than between pairs. For example, in group I, such pairs are Li ($Z = 3$) and Na ($Z = 11$), K ($Z = 19$) and Rb ($Z = 37$), and Cs ($Z = 65$) and Fr ($Z = 88$); in group III, they are B ($Z = 5$) and Al ($Z = 13$), Ga ($Z = 31$) and In ($Z = 49$), and Tl ($Z = 81$); in group V they are N ($Z = 7$) and P ($Z = 15$), As ($Z = 33$) and Sb ($Z = 51$), and Bi ($Z = 83$); etc. Additionally, it has long been noted that neon differs from helium much less than from argon; similar differences are observed between krypton and argon and between krypton and xenon. This circumstance within the framework of any of the current tabulated versions of the Periodic System—short-periodical as well as long-periodical—cannot be explained with sufficient clarity. In the proposed “semi-long-periodical” version (Table 3), it seems quite natural, since the named pairs within each of these groups belong to different periods (for example, He and Ne to the first period, Ar and Kr to the second, and Xe and Rn to the third). Additionally, in this version, the so-called phenomenon of secondary periodicity (nonmonotonicity of changes in the elements’ properties having the same type of electronic configurations) becomes quite natural. An example of this is the dynamics of changes in the compound stability with the highest oxidation state of halogens, decreasing from Cl to Br, increasing from Br to J, and decreasing again from J to At. This follows directly from the division of halogens into pairs (Cl, Br) and (J, At), belonging to the second and third periods, respectively.

As is known, before the advent of the “actinoid” concept of G. Seaborg, the most common version of the Periodic System was a slightly different version of the short-periodical version of the Periodic System, in which the elements Th ($Z = 90$), Pa ($Z = 91$), and U ($Z = 92$) were positioned, respectively, in groups IV, V and VI, followed by the so-called “trans-uranium elements”, which, as the lanthanoids, were taken out of the main table in the form of a separate “appendix”. Moreover, in chemistry textbooks up to the 1970s, Th was considered an analog of the elements Ti, Zr, and Hf; Pa as an analog of V, Nb, and Ta; and U as an analog of Cr, Mo, and W. After the above Seaborg’s hypothesis received experimental confirmation, the above chemical elements with $Z = 90$ – 92 were assigned to actinoids, and the positions they occupied in groups IV–VI remained unoccupied for some time. Currently, they contain 6d elements of rutherfordium Rf ($Z = 104$), dubnium Db ($Z = 105$), and seaborgium Sg ($Z = 106$) (named after the creator of the “actinoid concept”). Returning to this concept, within its framework, it remains (and still remains) incomprehensible why the first elements of this series—from Pa to Cm—exhibit oxidation degrees that significantly exceed +3, although formally, actinoids, such as lanthanoids, in both the main modern tabulated versions of the Periodic System—in the short-periodical as well as in the long-periodical versions—adjoin group III. The “semi-long-periodical” version of the Periodic System gives the answer to this, because in it, Th, Pa, and U appear in groups IV, V, and VI, respectively, so their maximal oxidation degrees (+4, +5, and +6) look natural and expected. Moreover, Np is positioned here in the seventh group and Pu in the eighth one, so the presence of both elements in the oxidation degree +7, and in Pu, even +8, seems quite natural.

The chemical elements thulium (Tm, $Z = 69$) and mendelevium (Md, $Z = 101$) in some cases exhibit an oxidation degree of +1; moreover, in the case of Md, in acidic solutions, it becomes very stable (from these solutions, mendelevium even precipitates along with sparingly soluble compounds of *s*-elements of the same group) [6,7]. At the same time, as far as we know, no other *f*-element has yet been found in chemical compounds with such a degree of oxidation. Within the framework of the short-periodical and long-periodical versions, this fact is also in no way connected with the arrangement of these elements; in the “semi-long-periodical” version, it finds its explanation, because both of these elements, as well as alkali metals, must be in group I. It is noteworthy that without exception, all the

chemical elements placed in this group within the framework of the “semi-long-periodical” version have the $ns1$ fragment in their electronic configuration.

Ytterbium Yb ($Z = 70$) and its analog—nobelium No ($Z = 102$)—in many compounds exhibit an oxidation degree of +2, resembling in this respect the elements of group II-s-elements Sr ($Z = 38$, Ba ($Z = 56$) and Ra ($Z = 88$) and d-elements Zn ($Z = 30$), Cd ($Z = 48$) and Hg ($Z = 80$). Metallic ytterbium, like metallic strontium and barium, dissolves quite well in liquid ammonia, and its YbSO_4 salt is just as poorly soluble in water as the compounds of similar composition SrSO_4 , BaSO_4 , and RaSO_4 . For nobelium in aqueous solutions, only No^{2+} is stable, while No^{3+} is easily reduced to No^{2+} . These facts also do not find an explanation in the above most common tabulated versions of the Periodic System, but in the “semi-long-periodical” version, they become quite understandable since in this tabulated version, Yb and No, together with the named s-elements, are in group II.

One of the decisive criteria determining the value of a particular theory is that new facts can be predicted that have not yet been observed. What can be predicted using our proposed “semi-long-periodical” version of the system of elements? Let us give just one example of this.

It has long been noted that $5f$ -AO, in terms of the number of electrons used to form chemical bonds, is very similar to $4d$ -AO. Therefore, uranium, in its possible oxidation degrees, repeats tungsten, neptunium repeats rhenium, americium repeats iridium, and curium repeats platinum. Therefore, it can be expected that such a similarity should be observed in the oxidation degrees of the osmium–plutonium pair. As is known, the maximal oxidation degree of osmium is +8, and it can be expected that the same oxidation degree should be observed in the case of plutonium as well. Indeed, relatively recently, compounds were discovered where Pu exhibits this degree of oxidation (see, for example, [8]). Moreover, it can even be expected that the “neighbor” of Pu ($Z = 94$) in terms of the nuclear charge—Am ($Z = 95$) will be able to show an oxidation degree +9 (for now, the maximal known oxidation degree of the given chemical element is +7). For Ir, which in the “semi-long-periodical” version is in the same group as Am, however, such an oxidation degree has not yet been observed experimentally (as, by the way, for any other of the known chemical elements). Nevertheless, there are fairly reliable theoretical calculations of the structure and thermodynamic characteristics of iridium(IX) nonafluoride IrF_9 using the CCSD, CCSD(T), and DFT B3LYP methods [9]; the $(\text{IrO}_4)^+$ cation and Rh(IX) and Mt(IX) cations similar to it in composition [10]; and iridium(IX) nitride-trioxide IrO_3N [11] and rhodium(IX) RhO_3N [12].

Of course, the following objection to the described “semi-long-periodical” tabular version of the Periodic System can be made: it is possible that it will have no less (or maybe even more) inconsistencies between the group number and the maximal oxidation degree exhibited by the chemical element than occurs in the short-periodical version. Therefore, it is unlikely that Pd and Pt, which will be in the X group (see Table 3), will be able to exhibit an oxidation state of +10 in their somewhat stable chemical compounds. It would be even less likely to expect Dy and Cf, placed in group XII, to have an oxidation state of +12. However, on the one hand, oxidation degree is just one of the “columns” in the “questionnaire” of a chemical element, and for this reason alone, it cannot serve as the only measure of the value of any version of the systematics of these same elements. Additionally, on the other hand, the Chemical Elements Periodic System approved by IUPAC in 1989 (i.e., the long-periodical version) is also very far from perfect in terms of the correspondence between the group number where the chemical element is located and the maximal degree of its oxidation.

As mentioned above, Table 3 shows chemical elements with $Z = 1$ –216 values. However, at the present time, this table is only slightly more than 50% filled with chemical elements. Additionally, the question arises as to when it will be completely filled, and whether such a number of chemical elements can even exist. According to the data of many studies (see, in particular, [13–21]), with increasing Z , the ability of atoms of a chemical element to radioactively decay rapidly increases, and therefore, their stability also decreases

just as quickly in general (in particular, the half-life of all currently synthesized isotopes of 7p-elements with $Z > 112$ is fractions of a second). Additionally, it is quite possible that the limiting Z value is already close. The problem associated with establishing this limit (and, consequently, the total number of chemical elements) is very interesting in itself, but its discussion, in our opinion, is already beyond this article's scope, and it rather belongs to the field of modern nuclear physics than modern general and inorganic chemistry. Nevertheless, in the conclusion of this article, it is worth mentioning a very important point associated with the Z growth and about which D.I. Mendeleev himself did not even guess (and modern inorganic chemists mostly know almost nothing about it). The author of this article is referring to the so-called "relativistic effect", according to which with increasing Z , the speed of electrons' movement increases and for heavy atoms becomes comparable with the light speed (c) (in particular, for a 1s electron in an Au atom ($Z = 79$), it is 0.58 c). This phenomenon leads to compression, first of all, of deep-lying AOs; however, higher-lying AOs are also subjected to compression to one degree or another. This phenomenon leads to compression, first of all, of deep-lying AOs; however, higher-lying AOs are also subjected to compression. Already for sixth-period elements (in the "semi-long-periodical" version, for the elements of the third period), this effect becomes so significant that it begins to affect the external AO, the electrons of which participate in the chemical bond formation, and, therefore, it must be taken into account in assessing the chemical properties of both the atoms of these elements and those chemical compounds formed by them. According to a number of nuclear physicists, a very strong compression of the 7s and 7p AOs should take place [14]. How significant this factor influence is can be judged from the results of [22], where the calculation of macroscopic characteristics for the heaviest of the known chemical elements, oganesson Og with $Z = 118$, was carried out with and without taking into account the relativistic effect. In the first variant, it turns out that oganesson must have a melting point of 220 K (which most likely implies a gaseous state at room temperature, as well as for radon), while in the second, it is a substance with a solid state of aggregation and a melting point of 325 K. The difference is very significant, more than 100 K. On the other hand, there is reason to believe that not only will simple substances formed by such elements as copernicium (Cn, $Z = 112$) and flerovium (Fl, $Z = 114$) have a very low reactivity but moreover, in their chemical properties, both of them should be close to noble gases (!) such as radon (Rn, $Z = 86$). What is really amazing is that a number of researchers, based on theoretical calculations, believe that copernicium in the form of a simple substance should be gaseous (!) already at room temperature [14]. However, there is also a directly opposite opinion: Cn will turn out to be a noble metal due to the relativistic splitting of 6d-AO [14]. In a recently published paper [23], it was predicted that copernicium, like its Group II neighbor mercury, should be a volatile liquid at room temperature and have a density of $14.7 \text{ g}\cdot\text{cm}^{-3}$ close to that of mercury ($13.55 \text{ g}\cdot\text{cm}^{-3}$); at the same time, it is noteworthy that its melting point (283 ± 11) K lies higher, and its boiling point (340 ± 10) K, on the contrary, is lower than the similar parameters for Hg (234.32 and 629.88 K, respectively). However, unlike metallic mercury, metallic copernicium has a rather large band gap of 6.4 eV, and various dispersion interactions are strongly pronounced in it, making it similar to the noble gases formed by elements of the zero group.

In addition to the enhancement of the relativistic effect, with increasing Z , there is another important factor that directly affects the physicochemical properties of elements, namely, the convergence of the various AOs energies with an increase in their key parameters n and l . For example, for elements with $Z \geq 122$, where according to the current theoretical concepts, the outer level 5g should be filled with electrons, so "mixing" of the 5g- and the 6f levels following it in energy is very likely. In such a case, at least for some of these elements, it will be incorrect to consider them as 5g elements because, in their electronic configurations of the ground state, there will be both 5g and 6f electrons (and possibly also electrons of the 7d level, the next in energy according to the Madelung-Klechkovsky rule). Consequently, they can no longer be attributed to any one specific category in the same way as the "traditional" *s*-, *p*-, *d*-, and *f*-elements; rather, they will be some kind of "hybrid"

categories. However, it will not affect the positions of these elements in any of the tabulated versions of the Chemical Elements Periodic System discussed in this article.

4. Conclusions

“The Periodic Law is not threatened by destruction, but only a superstructure is promised” (in the original Russian version, “Периодическому Закону не угрожает разрушение, а лишь надстройка обещается”). The discoverer of the Periodic Law, D.I. Mendeleev, often repeated this phrase, which describes the significance of this law in the future of chemical science, both before and after the final recognition of the very fact of his existence (although what exactly is hidden under this “superstructure”, he, however, never specified). However, it was most likely about various graphical interpretations of the Periodic Law, because, apparently, he intuitively felt that finding a mathematical formula linking at least one of the properties of chemical elements with its atomic weight (or, more correctly, atomic mass) would be a very difficult (if not impossible) task. Additionally, more than 150 years already have passed since the discovery of this law, fully confirming the above forecast of this great Russian chemist, because during this time, a great many of the most diverse options for this “superstructure” were proposed, and there are good reasons to believe that research in this direction will continue for a long time. This is despite the fact that the creative thought of various “interpreters” of the system of chemical elements basically worked “idle”: only two tabulated versions, the short-periodical and long-periodical systems, were recognized among researchers, while the main array of graphic options remained unclaimed now and is unlikely to be in demand in the future. On the other hand, the totality of physicochemical data about properties of atoms of 7p- and 8s-elements obtained in recent years, and mentioned above, forces us to draw a conclusion that is probably not very pleasant for chemists in general or inorganic chemists in particular: for these and, probably, heavier elements, the Periodic Law, it seems, is no longer valid.

Studies devoted to solving the problem of the chemical elements’ systematics based on the use of the Periodic Law were, are, and probably will be relevant for a long time, not only for physicochemists who are compatriots of its discoverer, but also for physicochemists around the world. Additionally, they will be relevant because the optimal interpretation of the Periodic Law has yet to be created. What it will be—tabular or another geometric shape—the future will show. In the opinion of the author of these lines, a decision on the above problem should still be sought on the path of improving the two most important tabulated versions used in practice—the short-periodical as well as long-periodical,—and this circumstance prompted me to write the given article.

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