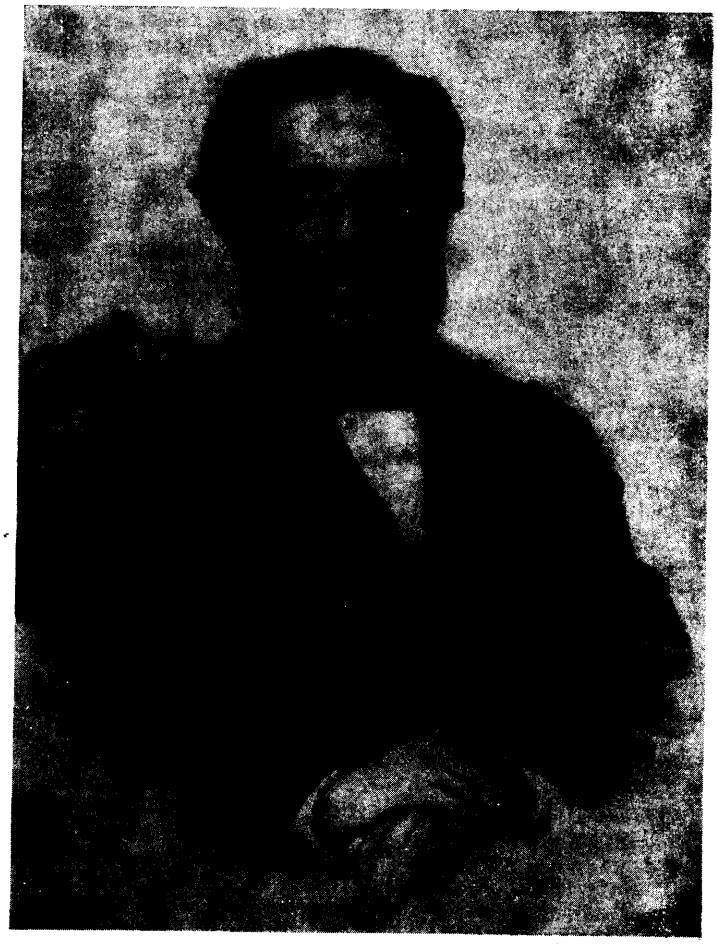
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D.I.Mendeleev in 1869.

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D.I.Mendeleev and His Discovery

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The centenary of D. I. Mendeleev's discovery of the Periodic Law of the chemical elements is being celebrated during 1969 by scientists in many parts of the world. The significance of this discovery in relation to the development of chemistry, physics and borderline sciences as well as for the advancement of agriculture and industry cannot be overestimated. It is impossible to visualise modern science and technology without the Periodic Law which has been a guide line both for theoretical and practical advances.

Mendeleev's discovery of the Periodic Law was not an accident. He, like many scientists whose names are associated with some scientific discovery, had predecessors. The discovery of the Periodic Law is a particularly good example of the sequential type of scientific thinking which is characteristic of all sciences and of all periods in time.

In order to understand the position which chemistry had reached by the time the Periodic Law was discovered one must recall the great chemical revolution associated with Lavoisier, the celebrated French scientist. In Lavoisier's time chemistry consisted of a mass of facts which needed to be arranged systematically and comprehended. Lavoisier created a new system in chemistry by drawing on the experience of his predecessors and his contemporaries and on the results of his own experiments; this system completely refuted the then accepted phlogiston theory. In his "Elementary Course of Chemistry" written in 1789, Lavoisier gives a list of "simple substances" which formed the foundation of the chemical nomenclature which he introduced. Among these simple substances Lavoisier distinguished what he termed "elements", defining them as simple substances which could not be broken down by chemical analysis. He included among them oxygen, nitrogen and hydrogen which he had discovered. Lavoisier was clearly aware of the fact that the concept of "a simple substance—an element" which he had introduced was in a sense relative, since, as he said in his text-book of chemistry "a given substance represents the contemporary limit of chemical analysis" since "at the present stage of our knowledge, analysis cannot further fragment it".

In fact all the subsequent advances in analytical chemistry, which had been enriched by new techniques (to mention a few; gas analysis, electrolysis, spectroscopic analysis), have characteristically led to the separation of a number of Lavoisier's simple substances into elements which could not be broken down any further. Many new elements were already discovered in the first quarter of the nineteenth century. Among the alkali metals discovered then were lithium, sodium and potassium, among the halogens—fluorine, chlorine and iodine; a series of metals of the platinum group—palladium, rhodium, osmium, iridium and ruthenium (platinum itself having been discovered in the eighteenth century), a number of alkali and alkaline earth elements, and many other elements were discovered at the beginning of the nineteenth century. Their discovery is associated with the names of such renowned scientists as Berzelius, Gay-Lussac, Davy and many others—chemists and physicists from many countries: Britain, France, Germany, Holland, Sweden, Russia and others.

The large numbers of elements newly discovered now had to be designated. The symbols for the elements which have been retained to the present day were proposed by Berzelius in 1813.

Dalton's atomic theory (1803), published by him in 1808 in a two-volume book "A New System of Chemical Philosophy", played a key role in defining the concept of an element. One of the hypotheses of Dalton's theory states that "every chemical element consists of simple atoms proper to it, which are different from the atoms of every other element". With the help of symbols which he introduced, Dalton could represent the molecules of various compounds (he called molecules complex atoms) by formulae which were sometimes quantitative. Thus, for example, he represented the molecule of carbon dioxide by a formula which consisted of three spheres—two white ones (representing atoms of oxygen) on the outside and a black one (representing an atom of carbon) between them. These were the first quantitative chemical formulae.

The discovery of new elements was parallelled by investigations of the chemical and physical properties of the elements and their compounds. In order to determine the structure of various substances it was essential to know their elemental composition quantitatively. Bergman in Sweden was the first to begin quantitative study of the composition of chemical substances, he started as early as 1770. Lavoisier placed exceptional emphasis on the quantitative analysis of substances and he, together with the French school of chemists and chemists in other countries, were responsible for developing many new techniques for exact quantitative analysis. The quantitative study of chemical substances is based on the law of the constancy of their composition. The French chemist Proust, who was much concerned with the preparation of chemically pure substances and their analysis, formulated this hypothesis as follows: "Nature, even through the agency of man, never produces compounds without scales in her hands—by weight and by measure". The techniques of quantitative analysis naturally continued to improve with time, reaching ever greater degrees of precision.

Much attention was paid to the study of the chemical properties of the elements. Investigations of the composition of chemical substances led to the concept that all substances were formed by the substitution of some atoms or groups of atoms by other ones. This view was first expressed by Dumas and Regnault in 1840 and developed further by Gerhardt and Kekulé. It followed from this concept that the atom of each element has a certain "capacity", i.e. can substitute or add on only a quite definite number of other atoms. This conclusion was reached by the British chemist Frankland in 1848-1852. The quantitative measure of such capacity is given by the valency of the atom—a term introduced by Wichelhaus in 1868.

Long before this, at the end of the eighteenth and beginning of the nineteenth century it was shown that chemical elements combined with one another in definite ratios by weight, which in itself asserted the constant nature of chemical substances. Analysis of oxygen compounds resulted in determining the amounts of various elements which combined with the same amount of oxygen. Wollaston (1814) called these values chemical equivalents. Ber-

zelius and Richter were largely responsible for determining precisely the weight ratios and chemical equivalents of elements. Faraday, working on the quantitative studies of electrolytic phenomena, opened up new possibilities for measuring chemical equivalents. Faraday measured the amounts of various elements by passing the same current through a solution of electrolyte and referring these to the amount of one of the elements, calling them electrochemical equivalents. The latter proved to be identical with the chemical equivalents. When the concept of valency was introduced it became evident that these values represented the ratio of the atomic weight to the valency of an element.

Thanks to the accuracy of determining the weight ratios of the elements Berzelius could determine equally precisely the atomic weights of many of the elements known in his day. The Belgian chemist Stas should be mentioned here; he determined very accurately the atomic weights of oxygen, silver, nitrogen, sulphur, chlorine, bromine, iodine, potassium, sodium, lithium, lead and carbon using his measurements of chemical equivalents. The Avogadro hypothesis played an important part in the exact determination of gaseous substances. Taking this hypothesis as a basis, Cannizzaro proceeded in the eighteen-fifties to establish more precisely the atomic weights of a series of elements, including hydrogen, oxygen, chlorine and others. Dulong and Petit found in 1819 that the product of the heat capacity of a solid element and its atomic weight had a constant value (the Dulong and Petit Law). Using this Law, Regnault first found the correct atomic weights of bismuth, silver, potassium and sodium, and Cannizzaro determined the atomic weight of copper and other elements.

As the result of the Herculean labours of many chemists and physicists from various countries and nations by 1869 there were 64 elements instead of the 35 simple substances of Lavoisier (1789), of which no more than 20 could be accurately described as elements; for most of these 64 elements the atomic weights were known sufficiently accurately, as were their chemical and some physical properties. All this accummulated material now needed to be systematically arranged.

The first attempt in this direction was made by Doebereiner who as early as 1829 noted that some elements which were similar in their chemical properties could be grouped into threes (triads). Such triads included: chlorine, bromine and iodine; lithium, sodium and potassium; and calcium, strontium and barium. At the same time in each such triad there was some consistency in relation to the atomic weights of the elements forming the given triad, viz., the atomic weight of the middle element in the triad was approximately equal to the arithmetic mean of the atomic weights of the two end elements. In 1851 Dumas gave some attention to Doebereiner's triads.

In 1850 Pettenkofer published a table of the equivalent weights of eighteen elements, arranged in five groups consisting of elements with similar chemical properties. Three of these groups included the elements which figured in Doebereiner's triads, but they now consisted of four elements, e.g. oxygen, sulphur, selenium and tellurium formed a group. He also noted that the difference of the equivalent weights in each group was approximately a multiple of a whole number. In the lithium, magnesium and oxygen groups this number equalled 8. Analogous groups of elements also appeared in Dumas (1851). In 1857 Lenssen's table appeared, in which he quite formally grouped all the elements then known into triads, with the result that some triads contained as chemically dissimilar elements as, for example, oxygen, nitrogen and carbon. In 1862 de Chancourtois arranged all the elements in an ascending order of atomic weights along a helix encircling a cylinder whose circumference he divided into 16 equal parts. Some elements with similar properties were found to lie one over the other along a vertical line, but this was not always the case. De Chancourtois worked on the hypothesis that the properties of the elements change periodically, each period consisting of 16 elements.

In his book "Modern Theories of Chemistry and their Significance for Chemical Statics" which was published in 1864, Lothar Meyer using new, more accurate atomic weights, made up three tables in which elements with similar valency (or atomicity as he termed it) were grouped together; in other words the chemical properties of the elements in each group were similar and there was an approximately constant difference in their atomic weights. The first table contained quadrivalent elements from carbon to lead, trivalent ones from nitrogen to bismuth, divalent ones from oxygen to tellurium, monovalent ones from fluorine to iodine, monovalent metals from lithium to caesium (Meyer included in this group also thallium, which he erroneously considered as a monovalent metal) and divalent metals from beryllium to barium. In the second table Meyer put quadrivalent titanium, zirconium and tantalum and hexavalent molybdenum, vanadium and tungsten. The third table contained quadrivalent manganese, iron, ruthenium, platinum, nickel, rhodium, iridium, cobalt, palladium and osmium, divalent zinc, cadmium and mercury, and without indicating their valency, copper, silver and gold. Meyer wrote: "Undoubtedly, there is some regular consistency in the numerical values of the atomic weights ..." without attempting to discover what this consistency might be. In the same year, 1864, Odling presented a table of the elements in his text-book of chemistry also using the new atomic numbers without indicating the principle underlying the arrangement of the elements in the table and only remarking that "the regular arithmetic ratios represented by the atomic weights are too numerous and too clear not to depend on some unknown law". Finally, one must also mention the table made up by Newlands in 1866, in which the elements were arranged in order of ascending atomic weights, eight to a row. Newlands called his table "the Law of Octaves" and noted a periodic recurrence of the elements' properties. At the same time there are obvious inconsistencies in Newland's table, such as for example the appearance of such dissimilar elements as hydrogen, fluorine and cobalt in one group.

It can thus be seen that by 1869 there were already many attempts at systematising all the then known chemical elements. Mendeleev was aware of all these attempts and divided them into artificial, i.e. formal ones, and natural ones. He assigned attempts such as Newland's octaves to the former, and those tables which were based on chemical and other properties of the elements—to the latter. This principle, which was the only correct one, also underlay Mendeleev's own system of the elements, which he rightly termed natural.

Unlike his predecessors, who could not see the general law which some of them only vaguely guessed at beyond the particular relationships noted for individual groups of elements, Mendeleev aimed right from the start at finding a general law which would be characteristic of the system of the elements as a whole. In his article "The Problem of a System of the Elements", Mendeleev wrote: "I took, with a few exceptions, the same groups of similar elements as did my predecessors, but my aim was to study the consistent features in the relationships between the groups. This led me to ..... a general principle which is applicable to

all the elements and embraces many of the previously expressed analogies, but also allows consequences which would have been formerly impossible". Speaking to the Chemical Society in England in 1889, Mendeleev said that his predecessors de Chancourtois and Newlands had "stood foremost in the way towards the discovery of the periodic law, and that they merely wanted the boldness necessary to place the whole question at such a height that its reflection on the facts could be clearly seen". It is not, of course, a question of their having lacked decisiveness and his having had the necessary daring and persistence. The discovery of the Periodic Law was the culmination of the whole of Mendeleev's earlier scientific endeavour. As his colleague, Ozarovskaya, relates, when she asked Mendeleev how his periodic system occurred to him, he said: "I have been pondering on it for perhaps twenty years".

Referring to one of his first scientific papers "Isomorphism in Connection with Other Relationships of the Crystalline Form to Composition", which he presented as a dissertation on graduating from the Principal Pedagogical Institute in St. Petersburg in 1855, Mendeleev wrote: "The writing of this dissertation involved me predominantly in the study of chemical relationships. It determined much of what was to come".

Mendeleev's particular scientific outlook and the breadth of his views and interests played an exceptional role in the fact that he was the first to see that the laws of the periodic system of the elements were of outstanding significance for the development of science. L. A. Chugaev, in characterising Mendeleev as a scientist of exceptionally wide horizons, wrote in his biographical article: "Mendeleev .... had the ability to bring a philosophical outlook to chemistry, physics and other natural sciences with which he came into contact, and a scientific experimental approach to problems in philosophy, economics and sociology". Further on Chugaev wrote: "in his endeavours to take into account the physical conditions and factors of chemical processes, Mendeleev was imperceptibly led into probing questions poised on the borders between chemistry and physics; he became interested in solutions and isomorphism, capillary phenomena, changes in state for substances and, in particular, the change in the volume of gases as a function of pressure and temperature. Work with rarefied gases opened up for him wide vistas of possible applications to meteorology; he developed a desire to probe the conditions prevailing in the upper strata of the atmosphere; he carried out remarkable work on the flow resistance of liquids, closely related to the problems of aeronautics, etc.". And further on: "With such a broad sweep of thought and many-sidedness of activity, all Mendeleev's work was both deeply pondered and carefully carried out. This became possible only thanks to his extraordinary capacity for work, which enabled him to work throughout whole nights, with hardly a few hours' rest".

While studying capillary phenomena Mendeleev discovered critical temperature which he termed "the temperature of absolute boiling". Mendeleev's studies of the compressibility of liquids, of solutions, of ammonia-metal compounds and thionic acids, as much of his other research, invariably represented new advances in science. The range of Mendeleev's scientific views and interests can be judged from his works on fuels, petroleum technology, explosives, metallurgical problems, metrology, agriculture and many other topics.

D. I. Mendeleev was born in 1834, the son of the Director of the Tobolsk Gymnasium. He entered the Gymnasium in 1841, finishing school in 1849, and in 1850 he was accepted at the Principal Pedagogical Institute in St. Petersburg. He graduated from the Institute in 1855 with a Gold Medal and was appointed to teach science in Simferopol, from which he moved to Odessa when the school in Simferopol was closed because of the Crimean War. Towards the end of 1856 Mendeleev returned to St. Petersburg. He gained his Master's degree there and defended a dissertation for the title of Privatdozent; in 1857–1858 he began teaching organic chemistry at the University of St. Petersburg. In January 1859 Mendeleev was sent abroad, to Heidelberg, for two years, returning in 1861. Whilst abroad, Mendeleev took part in the famous International Chemical Congress in Karlsruhe in 1860. When Mendeleev returned to Russia he resumed lecturing at the University. In 1861 he published his book "Organic Chemistry" which was the first Russian text-book of organic chemistry. In 1864 Mendeleev was elected Professor at the Department of Chemistry of the Petersburg Technological Institute. In 1865 he presented his doctoral thesis "The Combination of Alcohol with Water" and at the end of that year his appointment as Professor of St. Petersburg University in the Department of Technological Chemistry was ratified. In 1867 he was appointed to the Chair of General Chemistry.

In 1876 Mendeleev was elected Corresponding Member of the Petersburg Academy of Sciences. He was an honorary member of five Russian Universities. Mendeleev was also awarded honorary degrees by Cambridge, Oxford and several more of the oldest Universities of Europe. He was elected Fellow of the Royal Society of London; member of the Rome, Paris, Berlin and other Academies of Science and honorary member of many scientific Societies. In 1896 Mendeleev became head of the Principal Bureau of Weights and Measures in St. Petersburg. Mendeleev died in 1907, in St. Petersburg.

Mendeleev was convinced of the existence of a periodic law by the periodicity of the chemical properties of the elements which he had clearly demonstrated in his Natural System of the Elements in 1870, by the formulae of the higher oxides and hydrides of the elements, the regularities in the realtionships of the atomic weights of the elements in each group of chemical analogues, the periodicity in the change of specific gravities and atomic volumes, in the properties of the various compounds of the elements and in a number of their other special features.

Mendeleev was deeply convinced that the Periodic Law was objectively correct and of the utmost scientific and practical importance and he spared no effort in making it widely available to science as soon as possible. He wrote: "Each law of Nature becomes of scientific significance only if it allows, as it were, some practical consequences, i.e. logical deductions of the sort that would explain the hitherto unexplained and indicate phenomena unknown until then, and especially if the law leads to predictions which can be verified experimentally".

Mendeleev used the Periodic Law to predict the existence and properties of many of the then unknown elements. The first one was gallium, discovered by Lecocq de Boisbaudran in France in 1875, scandium, discovered by Nielsen in Sweden in 1880 and germanium, discovered by Winckler in Germany in 1886. The discovery of these elements was a triumph for the Periodic Law, and many of those who at first doubted that Mendeleev's Law was in fact general and correct, became his supporters.

The same conviction of the rightness of the Periodic Law enabled Mendeleev to alter the atomic weights of a number of elements. Thus, for example, he indicated that according to his Periodic System uranium should be hexavalent and not trivalent as had been accepted up to then, and hence the atomic weight of uranium should be 240 and not 120.

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The atomic weight of indium was similarly corrected from the previously accepted 75 to 113. The atomic weights of uranium and indium are now known to be 238.07 and 114.82 respectively.

With time the atomic weights of elements were corrected and determined more precisely, new elements were gradually discovered as predicted by Mendeleev and marked with dashes in the empty positions of his Table. Some of these elements were not discovered until many years after Mendeleev's death in 1907. Thus, rhenium (Mendeleev's bimanganese) was discovered by V. and E. Noddack, Tacke and Berg in 1925, and francium (ekacaesium) by Perey in 1939.

Mendeleev, naturally, could not have predicted the group of inert gases because of their characteristic chemical inertness. Rayleigh and Ramsay discovered argon in 1894 in England, and in 1895 Ramsay discovered helium which had been discovered as early as 1868 in the sun's atmosphere by its spectrum (Jansen and Lockyer). Ramsay was convinced that helium and argon were chemical analogues, unlike any of the known elements in their chemical properties, and therefore must constitute a separate group in Mendeleev's Table; in 1897 he therefore predicted, on the basis of the quantitative relationships of the atomic weights of the elements in each group that other elements must exist in the helium group with the atomic weights of 4(helium) + 16 = 20, 40(argon) + 44 = 84 and 84 + 44 = 128, which he and his assistant Travers subsequently discovered in the atmosphere. The new rare gases, as Ramsay named them, had the following atomic weights: 20.18 (neon), 83.7 (krypton) and 130.2 (xenon). The group of rare elements was included by Mendeleev in his Table as the zero group.

Nevertheless, neither the creation of new groups, nor the more accurate determination of the atomic weights could remove those exceptions, although not numerous, which were observed in the distribution of the elements in Mendeleev's Table. Namely, in arranging the elements in an ascending order of atomic weights, Mendeleev had to make the following exceptions: potassium, with an atomic weight of 39.10 was placed before argon with an atomic weight of 39.94, nickel (58.69) after cobalt (58.93) and iodine (126.93) after tellurium (127.61). Apart from these exceptions Mendeleev encountered difficulties in placing the rare earth elements (lanthanides). He wrote in the eighth edition of "The Fundamentals of Chemistry": "For some rare metals with the R_2O_3 type of oxides there is no place in Group III of the Periodic Table .... I have not yet found any personal interpretation of this, and it appears to me as one of the most difficult problems presented by the Periodic Law".

These and similar difficulties (especially the difficulties in arranging the actinides) were not, and could not be, solved during Mendeleev's lifetime, since the atomic weight was the only reliable characteristic of an element then available.

Several years after Mendeleev's death his Periodic Law found an ally in the theory of atomic structure which began to develop vigorously from 1911, when Rutherford discovered the atomic nucleus. The next important step along the path of discovery of atomic structure and establishment of a connection between the structure of the atom and the Periodic Law was made by Moseley in 1913. Moseley studied the X-ray spectra of various elements and as a result of these established a simple connection between the wave length of certain lines of the characteristic X-ray spectra of the elements and their atomic number, which Moseley correctly identified with the nuclear charge, expressed in units of elementary charge. The truth of this conclusion was substantiated a little later in the very first experiments on alpha-particle scattering performed by Chadwick in Rutherford's laboratory. These experiments gave values for the nuclear charge in the atoms of copper, silver and platinum (29.3, 43.3 and 77.4) which coincided, within the limits of experimental error, with the atomic numbers of these elements in Mendeleev's Table (29, 47 and 78 respectively). It followed from this, and from later experiments, that the nuclear charge was always equal to the atomic number of the element in the periodic system.

It is also noteworthy that the experiments carried out by Barkla and others on X-ray diffraction by various elements revealed that for all the light elements investigated, the nuclear charge was equal to half the atomic weight. If a difference between half the atomic weight and the atomic number of the element amounting to 0.5 or less is taken to be negligible, then half the atomic weight is equal, with this degree of accuracy, to the atomic number from helium to argon, and only starting with argon does the atomic weight increasingly outstrip (twice) the atomic number of the element. The arrangement of the elements in a periodic table according to their atomic weight, as made by Mendeleev and his predecessors (before atomic numbers and their physical significance were known) acquired a theoretical basis only for the first three periods in the periodic system. The discovery of atomic numbers and the development of the atomic theory provided the Periodic Law, for the whole series, with a new and definitive formulation: the properties of the chemical elements show a periodic change related to their atomic number.

If the theory of atomic structure proved to be an ally of the Periodic Law, the periodic system from its side played an exceptionally important role in the development of the atomic theory. Moreover, the question of atomic structure itself arose only after the discovery of the Periodic Law. Nevertheless, the "bricks" from which the atom could be built up only became known some 30 years after the discovery of the periodic system of the elements. The concept of elementary electric charge was first formulated clearly by Helmholtz in 1881, but the electron itself—one of these "bricks", was only discovered by J. J. Thomson in 1897. The other "brick"—the atomic nucleus, as already mentioned, was discovered by Rutherford in 1911. Following the discovery of radioactivity (Becquerel, 1896) and the formation of helium as the result of radium decay (Ramsay and Soddy, 1903) it became evident that the elements could be transmuted one into the other, thus confirming the genetic interconnection of the elements, their mutual kinship. In 1913 Niels Bohr put forward the theory of the hydrogen atom. The hydrogen atom, with its electron levels, proved, in a sense, to be a prototype of all the elements in the periodic system. In 1924 Born and Heisenberg extended the theory of the hydrogen atom to the alkali metal atoms. In 1926 Heisenberg, using quantum mechanics as a basis, solved the problem of the helium atom, and then between 1926–1930 the theoretical problem of the atom with many electrons was solved.

The Pauli exclusion principle (1925), according to which an atom cannot have more than one electron characterised by the same set of four quantum numbers, the number n being able to have the values of 1, 2, 3, 4 etc., the quantum number l which can have values 0, 1, 2 ... $(l - 1)$, the quantum number m which can have values l, $(l - 1)$ ... $-l$ and the quantum number s which can be $+\frac{1}{2}$ or $-\frac{1}{2}$ (inserted by Ed. of Translation), the position of the electron in the atom is determined. This set of quantum numbers, together with Pauli's principle and the postulate that the

ground state of each atom is the state with the lowest energy, give the physical foundations of the periodic system of the elements. It follows from the values established in the quantum atomic theory and the relationship of the quantum numbers and the Pauli principle that, for a given value of the so-called principal quantum number n, which may be 1, 2, 3, 4, etc., the maximal number of electrons in an atom cannot exceed the value $2n^2$, i.e. 2, 8, 18, 32 etc. These electrons form a closed electron shell and are distributed among subgroups with the maximal number of electrons in each orbital being equal to $2(2l + 1)$ where the quantum number l has the values 0, 1, 2 ... $(n - 1)$. The electrons with the quantum numbers $l = 1$, 2, 3 etc. are called s-, p-, d-, f- etc. electrons, respectively.

It follows from the above formulae that the single-quantum shell of the atom ($n = 1$) contains two s-electrons. This shell cannot accommodate more than two electrons and is therefore a closed electron shell which is first filled in helium the atom and then remains unchanged throughout Mendeleev's Table. The two-quantum shell can accommodate six p-electrons in addition to the two s-electrons. This gives a second closed eight-electron shell, which is completed in neon and remains, like the single-quantum shell, closed throughout the Mendeleev Table. The next, three-quantum, shell contains a maximum of 18 electrons: two s-electrons, 6 p-electrons and 10 d-electrons. It was found that only the heavier elements, starting with copper (atomic number 29), possessed the maximal number of electrons, viz. 18. In the three lighter elements, argon, potassium and calcium (atomic numbers 18, 19 and 20) the still incomplete eight-electron third shell already showed the properties of a closed shell, as witnessed by the chemical and physical properties of the Ar atom and the K^+ and Ca^{++} ions.

This and similar facts can be explained as follows. In passing from an element with the atomic number N to the next heavier element N + 1, i.e. with an increase in the nuclear charge and the number of electrons by unity, the "(N + 1)-th" electron occupies the lowest possible energy level. The observed sequence of filling the electron orbitals of the first 20 elements (ending with Ca) indicates that the 19th electron in potassium and the 20th electron in calcium are not d-, but s-electrons, i.e. they are placed in the four-quantum shell which corresponds to the lowest energy; this also follows from the theoretical model of the K and Ca atoms. The d-orbital of the three-quantum shell only begins to fill starting with the next element, scandium (atomic number 21), which is complete in copper, as has already been mentioned. The irregularity with which this orbital is filled, determined each time by minimizing the energy of the atom, explains all those anomalies which are observed in the properties of the elements from scandium to copper. A similar irregularity in the filling of electron orbitals underlies the anomalies in the elements from yttrium to palladium, in the rare earth elements—from the 58th element, cerium to the 70th, ytterbium in which the 32-electron shell with $n = 4$ is first completed, and in the group of heavy elements starting with actinium (atomic number 89).

The discovery of isotopes and the insight into the structure of atomic nuclei played a significant part in the understanding of the structure of the system of the elements and the relationship between the atomic weight and the atomic number of the elements. The problem of the isotopes first arose after the discovery of radioactive elements—the number of products of the radioactive decay of uranium and thorium exceeded the number of free places in Mendeleev's Table. The only course open was to place the elements with similar chemical properties but different atomic weights into one space. Soddy, in 1910, called such variants of the element isotopes. It is interesting to note that the thought of the possible existence of isotopes had been voiced, in passing, by a number of nineteenth century scientists. Thus, for example, Butlerow, addressing the Russian Physico-Chemical Society in 1881 said: "... Most atomic weights approximate whole numbers to such an extent that it can hardly be considered fortuitous. Could it not be supposed that under certain conditions atomic weights might actually be expressed by whole numbers (related to H = 1), i.e. that the atomic weights represent values which, under certain conditions, can change within certain limits?" Isotopes proved not to be an exclusive feature of radioactive elements only. Already in 1912 J. J. Thomson found that there were two varieties of neon atoms, with atomic weights 20 and 22. Now we know that each element consists of a certain number of stable and radioactive isotopes.

It is also known that in addition to isotopes there are also isobars—isotopes of different elements which, however, have the same atomic weight. Their existence demonstrates particularly clearly that it is not the atomic weight but the atomic number of an element, i.e. the nuclear charge and hence the number of electrons in the atom, which determines the chemical and other properties of the element. In particular, the chemical properties of the elements, according to the atomic theory, are determined by the outer electrons of the atom and their energetics. The periodicity of the chemical and other properties of the elements is a direct consequence of the periodicity in the arrangement of the orbitals and sub-orbitals in the electron shells of the atoms.

Isotopes also played a big part in the creation of the theory of the atomic nucleus which, in turn, was of the greatest significance in interpreting the structure of the periodic system of the elements, in particular the relationship between the atomic weight and the atomic number of the elements. Nuclear forces acting between the nucleons—the component parts of the nucleus, protons and neutrons, which at a certain ratio between the number of protons and neutrons bind them into a whole—the atomic nucleus. However, besides the forces of attraction between the protons in the nucleus there are also forces of electrostatic repulsion, which make the nucleus less stable. When the number of protons in the nucleus is small the nucleus may remain stable even with an excess of protons since the nuclear forces are several times stronger than the repulsive forces. The light isotope of helium ^3He has such a nucleus, which consists of one neutron and two protons. The isotopes of all the elements following helium, however, are unstable when the number of protons exceeds the number of neutrons and this explains the fact that the higher the atomic number of the element, i.e. the larger the number of protons, the more does the number of neutrons in the nucleus of its stable isotopes exceed the number of protons.

We can thus see that the theory of atomic structure, of its electron shells and its nucleus, proved to be the physical basis upon which the periodic system of the elements rests. All the difficulties in arranging the elements within the system and the contradictions which arose during the first stages of constructing the system disappeared in the light of the atomic theory. At present the periodic system of the chemical elements forms the foundations on which rests all our science, the guiding light illuminating the progress of science in the future.

The role of the Periodic Law in the advancement of science was utterly clear to Mendeleev from the first years

of the existence of his periodic system of the elements. Addressing the Chemical Society in England in 1889 he referred to the Periodic Law as an "instrument of enlarging our mental horizon" and said that "the law of periodicity enabled us to perceive undiscovered elements at a distance which formerly was unaccessible to chemical vision".

The Periodic Law gave an impetus to the advance of inorganic chemistry. It had an enormous and decisive significance for the development of the atomic theory and the atomic nucleus, for nuclear physics and nuclear chemistry. It is exceptionally important for the development of geochemistry, geology, crystallography, mineralogy and metallurgy. It is a guiding light in catalysis and the consideration of the rates of chemical reactions. There is no branch of science where the Periodic Law does not exert its influence.

The whole world celebrates gratefully the centenary of Mendeleev's achievement and recalls the names of all those whose endeavours and genius have put before us the Periodic Law in all its majesty and brilliance.