The Periodic Law and its Interpretation.

MENDELÉEF CENTENARY LECTURE, DELIVERED BEFORE THE CHEMICAL SOCIETY AT THE ROYAL INSTITUTION, LONDON, ON APRIL 19TH, 1934.

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THIS year is the centenary of the birth of Dimitri Ivanovitch Mendeléef (1834—1907), whose name is so indelibly associated with the great advance in our knowledge of the classification of the chemical elements. I have been asked by the Chemical Society to give you a lecture appropriate to the occasion. I feel, however, less fitted than any of my audience to deal adequately with the chemical side of this subject, but I am encouraged in my task by the reflection that I, either directly or indirectly, have been closely associated with several developments which have helped to give a much clearer view of the underlying meaning of this famous law of the elements.

I trust, therefore, that you will excuse me if I should spend but little time on the earlier history of this subject, but try to give you my personal impression of the way in which our knowledge of this subject has widened in later years.

On looking back, we can see that the collection of data on the physical and chemical properties of the elements and particularly of their atomic weights was sufficiently ample about the years 1860—1870 to lead speculative minds to consider whether any order could be traced amongst the great diversity of elements, and particularly whether their properties could be correlated with their atomic weights. While the early suggestions of Prout that the atomic weight of the elements might prove to be integral multiples of the atomic weight of hydrogen had been shown to be inconsistent with the chemical facts, yet this hypothesis had an undoubted and lasting influence in directing the attention of scientific men to the problem of the relations between the elements. A number of tentative suggestions had been put forward from time to time, indicating that there was some sort of repetition of chemical properties when the elements were arranged in the order of their atomic weights. In this connection, mention should specially be made of the work of Newlands in England in 1864 and of Lothar Meyer in Germany.

This aspect of Chemistry had a special attraction to the philosophic mind of Mendeléef, and led him to recognise that there was evidence of a definite kind of order in the variation of chemical properties when the elements were arranged in sequence of their atomic weights. The abstract of his conclusions was published in 1869, and the full paper, published in 1871, contained the classification of the elements into groups or periods essentially the same as those we are familiar with to-day.*

This was a remarkable generalisation which covered not a selection of the elements but all the elements known at that time, although there were inevitably certain exceptions and anomalies that only received an explanation in later times. It is clear that Mendeléef had a firm belief in the general correctness of his classification, for he had the courage to use it to predict the atomic weights and properties of undiscovered elements for which there were vacant places in his ordered system.

The ideas of Mendeléef at first attracted little attention, for the chemist of his day was more occupied in adding to the chemical facts than in speculating on the relation between them. The fate of the Prout hypothesis had a deterring effect in making many minds critical of generalisations in such a difficult field. It was not until the discoveries of the elements gallium in 1875 and scandium in 1879 and the proof that they had the properties of two of the missing elements predicted by Mendeléef that the importance of his discovery was generally recognised. Even then many years passed before the importance and significance of these new ideas percolated to chemists as a whole and found adequate expression in the text books of the time.

* I should recall that Mendeléef was asked to give the Faraday Lecture before the Fellows of the Chemical Society in the theatre of the Royal Institution, June 4th, 1889. His lecture, entitled "The Periodic Law of the Chemical Elements," was published in the Chemical Society's Journal and in the appendix to the English translation of Mendeléef's "Principles of Chemistry."

The next great advance in our knowledge was the discovery of the group of chemically inert gases, of argon by Rayleigh, and of helium, neon, krypton and xenon by Ramsay. It seemed clear that a new group of zero valency must be added to the periodic classification, thus widening rather than altering the scheme of Mendeléef. The discovery of the radioactive emanations added the last member of this group. I look back with some pride to the fact that Soddy and I were able to prove that the radium emanation must belong to the group of inert gases, although the amount of emanation available for our experiments was less than 10^{-11} c.c. This was made possible by the extraordinary sensitivity of the radioactive methods as a means of quantitative analysis. While the rapid discovery of this new group of gases was at the time of extraordinary interest, yet it was not until twenty years had passed that their significance was recognised in providing us with a clue of the utmost value in deciphering the intricate problem presented by Mendeléef's law.

It is to be borne in mind that, while at the close of the 19th century the periodic law of the elements was everywhere recognised as a relation of outstanding importance, and had been incorporated in the teaching of chemistry, yet it was in a sense an empirical generalisation for which no explanation of any kind could be offered. There was no rational court of appeal to decide between the numerous contestants who differed as to the correct arrangement of the elements in the classification. All was a matter of opinion, with no theory or principle to serve as a guide. It is true that to the philosophic mind, the Mendeléef rule suggested that the atoms of the elements were not, so to speak, separate creations but must be built out of some more elementary form of primordial matter. The discovery of the electron as a universal constituent of all atoms of matter by Sir J. J. Thomson in 1897 lent much weight to these ideas. It is to him that we owe the first suggestions of the electrical structure of the atom and an explanation along general lines of the variation of chemical properties of an atom with increasing number of electrons in its structure. Before, however, such ideas could be made definite and directly applicable to the atom of the chemists, it was necessary to await a better understanding of the actual structure of the atoms and of the relation of the elements with one another.

From consideration of the scattering of α-particles by heavier atoms, I was led in 1911 to suggest that all atoms had a similar nuclear structure, the main mass of the atom being concentrated in a central nucleus of minute dimensions which carried an excess positive charge of electricity. The experimental results of Geiger and Marsden, made to test the theory, indicated that the nuclear charge of an element was approximately proportional to its atomic weight and in fundamental units equal in number to about half the atomic weight. It was clear that this rule held only for the heavier elements and broke down for the lighter elements. For example, it was recognised that the hydrogen nucleus had a charge 1, and the helium nucleus a charge 2. The idea that the nuclear charge of an element might be given by its ordinal or atomic number was first suggested and used by Bohr in developing his theory of spectra. By a strange oversight, Bohr himself gave the credit of this suggestion to van den Broek, who later discussed the applicability of this conception to the elements in general.

The Order of the Elements.

Before discussing the next great advance, it may be of interest to summarise the general mental attitude towards these atomic problems in my laboratory in Manchester, where Bohr and Moseley were both working, at the period 1912—1913. There was a firm belief in the general correctness of the nuclear theory of the atom, and it was recognised that the properties of the atom were defined by a whole number which represented the number of units of positive charge carried by the nucleus. It seemed probable, too, that the charge on the nucleus might prove to be equal to the ordinal number of the element. The number and motions of the external electrons were seen to be dependent on the nuclear charge, and in the hands of Bohr the first beginning had been made to explain the spectra of the lighter atoms by the application of the quantum theory and thus to throw light on the distribution and motions of the outer electrons. Bohr, too, had pointed out that there must be a clear-cut distinction between the properties to be ascribed to the nucleus and the properties belonging to the outer distribution of electrons. For example, it seemed clear that the essential radioactive properties must be ascribed to the nucleus itself, and the ordinary light and X-ray spectra to the disturbance of the motion of the electron satellites.

It was at this stage that Moseley began his famous experiments that were to throw such a clear light on the problem of the number and order of the elements. A few months after the discovery by Laue, Friedrich, and Knipping that X-rays could be diffracted by their passage through crystals, Moseley and Charles Darwin at my suggestion had begun experiments to study the diffraction of X-rays by crystals, using an electric method. In the course of their experiments, Sir William Bragg announced that he had found that a bright line spectrum could be obtained by the reflection of X-rays at the surface of a crystal. This was confirmed by Moseley and Darwin, who mapped out in detail the complicated " L " spectrum given when platinum was bombarded by swift electrons. After this investigation was completed, I remember well Moseley coming to discuss with me the next research he should take up. He outlined four or five possible investigations in which he was interested and amongst them was one that at once attracted me. This was to test directly whether the properties of an element depended on its atomic weight, as had been generally supposed in the periodic law of Mendeléef, or whether it was governed by its ordinal or atomic number as seemed probable on the nuclear theory. He proposed to examine this question by studying the changes in the X-ray spectra of a definite group of elements, using the photographic method for recording and measuring the spectra. It is to be remembered that at this stage very little was known about the X-ray spectra of the elements, and no work had been done on the spectra given by the light elements. On examination, the " K " spectrum of the light elements was found by him to be unexpectedly simple, consisting essentially of two strong lines in all cases. The frequency of vibration of corresponding lines was found to increase according to a simple law on passing from one element to the next in order, the frequency being proportional to $(N - a)^2$, where N was the ordinal number of the element and a a constant. For the heavier elements, it was more convenient to use the " L " spectra, and a similar law connecting frequency and ordinal number was found to hold for the stronger lines in these spectra. In interpreting this relation, Moseley assumed that N represented not only the ordinal number of the element but also the magnitude of its nuclear charge. In making this deduction, he was influenced by Bohr's theory of spectra which had just then been published; for Bohr had shown that the frequency of vibration of an electron in its innermost orbit must be proportional to the square of the nuclear charge; a was assumed to be a constant arising from the presence of one or more electrons in reducing the effect of the central nuclear charge on the vibrating electron. We now know that Moseley was essentially correct in all these deductions. It was only later, in 1923, that Chadwick was able to prove by direct scattering experiments that the nuclear charge was equal to the ordinal number of the element as defined by Moseley. Although Moseley only examined a selection of the elements in his pioneer experiments, the general laws he deduced have been found by subsequent observers, and particularly by Siegbahn, to hold for all the elements which have been examined.

A relation of extreme simplicity is thus seen to hold for all the elements, for their properties are defined by a whole number which represents the ordinal number of the element and at the same time the number of units of charge on the nucleus.

This discovery of Moseley represents an outstanding landmark in the history of our knowledge of the elements, for it fixed once for all the true order of the elements and showed that only 92 elements were possible from hydrogen 1 to the heaviest element uranium 92. It fixed at once also the number of elements possible in the group of rare earths and afforded a simple method of testing the purity of the preparations.

Many of the anomalies in the Periodic Law at once disappeared when the atomic numbers of the elements were substituted for their atomic weights; for example, the elements cobalt and nickel, numbers 27 and 28, fell into their proper place in the new system, although according to their atomic weights 58·97, 58·15, their positions were reversed. It became clear that the atomic weight of an element, though often running nearly parallel with the atomic number, was in a sense a secondary property, and this is borne out by our later knowledge. Moseley's law not only fixed the ordinal number of the missing elements, but allowed us to predict with considerable accuracy the X-ray spectra to be expected from them, while the use of the X-ray spectra provided a new and powerful method of qualitative and quantitative analysis in the search for missing elements. The application of these new ideas soon led to the discovery of hafnium 72, of masurium 43, and of rhenium 75. Apart from the rare-earth element 61, named illinium, of which we have heard little, there remain only two missing elements, 85 and 87, and 85 has been detected by its X-ray spectrum but not yet isolated. The search for undiscovered elements thus seems nearly to have reached an end unless elements of numbers higher than 92 exist in the earth.

It has always been a matter of surprise to me that practically all possible nuclear charges from 1 to 92 are represented by actual elements in this earth. Apart from the radioactive elements, all these atoms appear to be permanently stable under normal conditions on the earth's surface. If we regard the nucleus as a complicated quantum-dynamical system, it might almost have been anticipated that in some cases the nuclear structure corresponding to certain nuclear charges might be unstable and break up, leaving some gaps in our list of elements. The absence of any such gaps in thus a clear indication that stable elements are possible for all values of the nuclear charge, disregarding of course the radioactive elements. This stability of the elements is still more remarkable when we recall that many of the elements consist of a mixture of atoms of different masses and thus of different nuclear constitution.

Discovery of Isotopes.

While Moseley's law fixes definitely the number of elements defined by their nuclear charge, it does not of itself throw any light on the number of species of atoms of different mass that may exist with the same nuclear charge. The proof of the complexity of the chemical elements, to which attention was first drawn by Soddy, arose from a study of the radioactive substances, and is a discovery of the first importance. Certain radioactive elements, although they had different atomic weights and showed distinctive radioactive properties, were found to be chemically inseparable. Soddy concluded that these distinctive species of elements must occupy the same place in the periodic table and for that reason named them "isotopes." We now know that the series of transformations in the radioactive bodies give rise to several isotopes of the elements thorium, radium, bismuth and lead. For example, there are produced four radioactive isotopes of lead (82), varying in atomic mass between 210 and 214, and also three non-radioactive isotopes—the end products of the series—of atomic masses 206, 207, 208. All of these radioactive isotopes of lead emit β-rays, but with very different energies, while their stability varies over a wide range, the average life of the elements varying from 39 minutes to 35 years. The end products of the series on the other hand appear to be as permanently stable as any of the ordinary elements.

These ideas, originating from a study of radioactive transformations, received a great extension when it was shown that the ordinary elements were also in general not simple but composed of two or more isotopes of different masses. This development, which we owe largely to the fine experimental work of Aston, depended upon the production of a stream of charged ions of the elements and their deflexions in a combined magnetic and electric field. In the ingenious type of mass-spectrograph devised by Aston, the relative masses of the isotopes can be measured with great accuracy and their relative abundance determined.

These important results received a simple explanation on the nuclear theory. The nuclei of the isotopes of a given element must all have the same charge but have different atomic masses. The arrangements of the outer electrons in the isotopes of a given element are so nearly identical that there is only a second-order effect even on the optical or X-ray spectra, which are only slightly dependent on the mass of the nucleus. Even in the case of the heavy isotope of hydrogen, mass 2, which has recently been discovered, the ordinary spectrum is not only similar in form to that of hydrogen of mass 1, but the corresponding Balmer lines differ in wave-length by only 1 part in 3700. We should thus expect that the properties of an element which depend on the arrangement of the outer electrons should be practically identical for all isotopes. On the other hand, properties, for example, diffusion, which depend on the mass of the isotope should be altered. While such differences are in general not marked for the isotopes of the heavier elements where the fractional change of mass may be small, it may become very pronounced for large relative changes of mass such as occur in the case of the hydrogen isotopes of masses 1 and 2. It is for this reason that the isotope H^2 can be readily separated from H^1 either by diffusion or by the processes occurring in electrolysis which depend upon the mass.

During the last few months, successful methods have been devised to separate weighable quantities of the isotopes of lithium 6 and 7 in an almost pure state by methods depending on the deflexion of the ions of lithium in a magnetic or electric field, and no doubt it will be possible by similar methods to separate the isotopes of a number of the lighter elements. Oliphant used some of this material to examine the distinctive transformations produced by each of the separated isotopes of lithium when bombarded by the ions of ordinary and of heavy hydrogen.

For experiments in transmutation, where the effects of bombardment may be entirely different for the individual isotopes of the same element, it is essential for progress to have reasonable quantities of each isotope for examination. The chemist and the physicist of the future will undoubtedly require for many investigations to have samples of each isotope in as pure a state as possible.

It is now well established that the masses of many of the isotopes of the elements differ by very nearly unity, indicating that either the proton or neutron or both are components in the structure of all nuclei. We are not yet certain whether these two units of nearly equal atomic mass—1·0073 for the proton in the free state and 1·0067 for the neutron as deduced by Chadwick—are independent or whether one is derived from the other. While each of these particles is expelled in many nuclear transformations, there is at present not sufficient evidence to give a definite answer to this question. It is of interest to note, as suggested by Harkins, that the neutron, if fundamental, may be regarded as the first element in the series of nuclear charge 0 and mass nearly 1.

It is thus clear that in a sense we have found that a modification of the Prout hypothesis is applicable to the masses of the individual atoms, but it must be borne in mind that probably two units, the proton and the neutron, are involved and not one alone. While in general the atomic masses of all the isotopes are given by integral numbers in terms of $O = 16$, yet this rule, while useful as a working guide, is only approximate. The relative masses of a number of the atoms are now known with sufficient accuracy to show the general trend of the departures from this whole-number rule with increasing weight of the atom. It is of great importance that the relative masses of all existing types of atoms should be known with the highest possible precision. In this connection it may be remarked that there is now good evidence that the generalised law of the conservation of energy, where changes of mass are taken into account, holds for nuclear transformations. The measurement of the energy supplied and released in these transformations provides us with data to determine the masses of some of the nuclei with even greater accuracy than is yet possible by the methods used by Aston and Bainbridge. There can be no doubt that a close study of nuclear transformations brought about by various agencies will not only provide important information on the structure of nuclei but will at the same time fix the masses of some of the lighter elements with great precision.

Time does not allow me to discuss with any detail the extraordinarily interesting data which have been accumulated by Aston and others on the isotopic constitution of the elements. I can only refer to the remarkable difference as to numbers and masses of the isotopes shown by even- and odd-numbered elements and the existence of a number of atoms of about the same mass—isobares—but of different nuclear charge. In general, isotopes are much more abundant for even-numbered elements about the middle part of the periodic table. For example, 11 isotopes have been found for tin, the masses varying from 112 to 124 with only 113 and 123 missing. The relative abundance of these isotopes varies, however, in an apparently capricious manner for which we have as yet no kind of explanation. In the last few years, optical methods have proved of great value in disclosing the presence of isotopes existing in very small amount compared with the main isotope, for example, in hydrogen, carbon, nitrogen and oxygen.

This raises the very interesting question whether an increase in the sensitivity of our methods for the detection of isotopes might not add largely to their number. I have already drawn attention to the remarkable stability of the nuclear structure disclosed by the presence of numerous isotopes of widely different masses. In the case of tin, for example, to which I have previously referred, it is natural to suppose on our present ideas that the isotopes differ from one another by the addition or subtraction of one or more neutrons to the nuclear structure. This will alter the mass by multiples of nearly unity but will not change the nuclear charge. It may be that, in addition to the isotopes found by Aston, others also exist in small amount over a much wider range of mass. Further information of this kind would be very valuable in extending our understanding of the stability of nuclear structures of even atomic number.

I may also refer in passing to the information gained by a study of the isotopic constitution in throwing light on anomalies of atomic weights of some of the elements. For example, argon has two isotopes 36 and 40, the latter being far more abundant, so that the atomic weight is 39·9. It is of interest to note here that Professor Zeeman has recently given good evidence, using the original positive ray method of Sir J. J. Thomson, that a weak isotope of mass 38 is also present. On older views of the relation between the atomic weights of the inert gases, it was to be anticipated that the atomic weight of argon should be about 36 and not 39·9. The abundance of the isotope 40 compared with 36 accounts in a sense for this anomaly, although no explanation can be offered why such a distribution of isotopes occurs.

From a study of isotopes we now know that more than 200 species of atoms exist in this earth, and no doubt this number will be added to in the future. The recent discovery of M. and Mme. Curie-Joliot that unstable atoms can be formed by bombardment of the elements boron, magnesium and aluminium by fast α-particles is a noteworthy event. The new atoms formed break up according to the same law as a radioactive body but, strange to say, emit positive electrons in the process. There seems to be little doubt that the number of these transition elements will be greatly extended by using projectiles like protons, diplons, and neutrons, as well as α-particles to bombard the atoms. This discovery opens up a new and important field of investigation which allows us to produce in the laboratory new types of unstable atoms which may have existed at the time of formation of our earth but have long ago disappeared.

Arrangement of Electrons outside the Atom.

I now come to the most important and at the same time the most difficult part of my subject, namely, the explanation of the periodic classification of the elements when the atomic number is substituted for the atomic weight. As we have seen, the number of external electrons surrounding a nucleus is normally equal to its nuclear charge and thus varies from 1 in hydrogen to 92 in uranium. Since the ordinary chemical properties to be ascribed to an atom must depend solely on the number and arrangement of these electron satellites, an explanation of the periodic law must be derivable immediately from the way in which the electrons are arranged round the nucleus. The first attack on this problem was made by Bohr in 1913 by applying the quantum theory to the explanation of spectra; but it was not until ten years had passed that the main features of the arrangement of the electrons in the outer atom became clear. The problem for solution appeared to be of the most formidable and forbidding character, for it was necessary not only to account in a general way for the periodic properties of the elements, but also to explain in detail the complex light and X-ray spectra given by all the elements under different conditions of excitation.

As a result of the work of Bohr and his collaborators, there has been a complete revolution in our knowledge of spectra and the main features of the spectra of all the elements are now well understood and in most cases the spectra have been worked out in detail. This is a veritable triumph of mind over matter and represents one of the outstanding scientific achievements of our age. I remember that, in the early days, when I reviewed the enormous complexity of the spectra of the elements and the apparent lack of any relation between them, it seemed to me that a solution of such a formidable problem would be long delayed, but yet, within a period of less than 20 years, the spectra of all the elements have been reduced from chaos into a definite and orderly system.

It would take much too long to discuss in detail the various considerations which led Bohr to fix the modes of binding and the energy states of the electrons surrounding the nuclei of the elements. In his first tentative attack on this problem, he supposed that the electrons circulated round the nucleus in certain orbits defined by definite quantum conditions. He concluded that the electrons were arranged in more or less definite groups of orbits round the nucleus. While the introduction of quasi-planetary orbits served, and continues to serve, a very useful purpose in assisting us to visualise the arrangement of the electrons in the atom, there has grown out of it a revised and improved version, the modern or wave-mechanics, with the development of which the names of L. de Broglie, Heisenberg, Schrödinger and Dirac are so intimately associated. The new theory enables accurate calculations to be made of atomic properties with, so far as we know, complete success. It is found, however, that the earlier idea of precise orbits must be abandoned and replaced for the present by electron distributions much less definite. Nevertheless, the old orbits remain a good first approximation to many features of the true distribution and we shall for that reason continue to use them here for descriptive purposes.

I am informed on good authority that the success of the new wave-mechanics is so complete in its own domain that the periodic law of the elements, as we know it to-day, could be entirely reconstructed from first principles. Given a competent mathematician—and he would require to be very competent—acquainted with the laws of the quantum mechanics of an electron and with one additional postulate—the Pauli exclusion principle—he should be able, even if he had never heard of the periodic law, to reconstruct the periodic classification of the elements. Naturally he would require plenty of time and help of computers, and unless he was exceptionally cautious might have occasionally to consider his steps very carefully at certain critical points of the classification. I do not mean to imply, and with that my informant would agree, that it would not be quicker and easier to formulate the periodic law from experiment rather than from calculation.

In describing the theory of the periodic law in the form most suitable for the science of chemistry, it is best to take as the starting point the peculiar and privileged group of inert gases, He, Ne, A, Kr, Xe and Rn, with 2, 10, 18, 36, 54, and 86 electrons surrounding their respective nuclei. These numbers of electrons are in each case just, and only just, sufficient to fill up completely a group of orbits which are distinguished from the next possible orbit but are far more tightly bound. As a further result of the group of orbits being complete, the electrons of the group present every possible orientation of spin and orbital angular momentum, with the result that the complete group is spherically symmetrical to the outside world and possesses only a very small residual field outside the atom. The result is that such structures must be extremely inert and non-reactive, as they are in fact observed to be. The theory also requires that these exceptional inert structures shall occur for these and only these atomic numbers.

When now we consider the atom of the next higher atomic number than one of these inert gases, all the electrons but the last one go into orbits similar to those of an inert gas and form a positive univalent ion of inert-gas type, with a highly symmetrical tightly bound structure but now with one residual charge. The last electron cannot go into any one of these groups (by Pauli's exclusion principle), for they are full. It must therefore fill the most tightly bound orbit of the next group, which is in fact only lightly bound, giving the atoms in the solid the properties of a metal. We must thus get, theoretically, after each inert gas a univalent metal, forming a group of elements of closely similar properties. These are, of course, the alkalis. By following the process of adding further electrons, one by one, we see how the properties of the elements, ordered by the atomic number, must arrange themselves in a periodic pattern, repeating after each inert gas.

But more than this, we can see clearly why the periodicity, though regular, is a periodicity of increasing length of period, with groups of transition elements like the nickel and palladium and platinum groups, and further, in the last complete period, with the super-transition group of the rare earths. These transition elements must occur and are found to occur whenever the relative order of binding of two groups of orbits changes with increasing nuclear charge, but I cannot spare time to describe the details of this extremely fundamental process. I must be content to say that the existence of the transition and rare-earth groups is a necessary consequence of the theory and is predicted by the theory in their actual places. I have so far said nothing about the application of the quantum mechanics to the theory of valency, but since its success has already been so great I should not pass it over in silence. Naturally, the problems presented by a study of chemical combinations are extremely complicated and one cannot expect exact solutions of any but the simplest. But if the theory is right and adequate to cover the field of observed facts, one should at least be able to say theoretically that such and such things can happen—have a natural home in the theory—even if we cannot say that they *will* happen in given circumstances. Judged in this way, so far as we can see, the quantum theory of valency is already a marked success and the applications of the theory are continually being extended to more and more complicated cases.

In conclusion, I should say a word about the construction of atomic nuclei, some of the properties of which I have already mentioned. Detailed information on this subject is very scanty and we have to rely in the main on the data obtained from a study of the radioactive bodies, and the artificial transformation of elements brought about by bombarding particles. By these artificial methods, we have not only been enabled to break up some of the atoms into lighter components but also to form atoms of higher atomic number and mass than the bombarded atom. As we have seen, some of these newly formed elements have a limited life and show a special type of radioactivity. The use of the ions (D^+) of heavy hydrogen for bombardment purposes has given us a wealth of new results of great interest. It leads in many cases to the liberation of groups of fast protons, neutrons and also α-particles from a number of light elements. The effect of bombarding heavy hydrogen by its own ions is especially noteworthy. It appears that several new types of atoms are formed, including a new isotope of hydrogen of mass 3 and also a new isotope of helium of mass 3, which have not hitherto been identified in Nature. It seems not unlikely that in the near future we may be able on a small scale to synthesise some of the lighter nuclei from more elementary nuclei.

We have so far little definite information on the detailed composition of nuclei, although it is believed that protons and neutrons are the main building units. When we know more of the structure of nuclei, it may be that they also will show some sort of periodicity in their structure with increasing nuclear charge. Since, however, at least two units and not one are involved, the periodic pattern may be much more complicated and require to be illustrated by a surface rather than a plane as in the case of the periodic law. It may be that a Mendeléef of the future may address the Fellows of this Society on the " Natural Order of Atomic Nuclei " and history may repeat itself.