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The Periodic Table

How the Elements Are Organized

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Chapter 1: The Discovery of The Periodic Table

In the 5th century BCE, Leucippus and his pupil Democritus proposed that all matter was composed of small indivisible particles which they called "atoms". Around 330 BCE, the Greek philosopher Aristotle proposed that everything is made up of a mixture of one or more roots, an idea originally suggested by the Sicilian philosopher Empedocles. The four roots, which the Athenian philosopher Plato called elements, were earth, water, air and fire. Similar ideas existed in other ancient traditions, such as Indian philosophy with five elements: Earth, water, fire, air and aether collectively called 'pañca bhūta'. In 1661, Boyle defined elements as "those primitive and simple Bodies of which the mixt ones are said to be composed, and into which they are ultimately resolved."

The first person in recorded history to discover a new element was Hennig Brand, a bankrupt German merchant. Brand tried to discover the philosopher's stone—a mythical object that was supposed to turn inexpensive base metals into gold. In 1669, or later, his experiments with distilled human urine resulted in the production of a glowing white substance, which he called "cold fire" (kaltes Feuer). He kept his discovery secret until 1680, when Anglo-Irish chemist Robert Boyle rediscovered phosphorus and published his findings. The discovery of phosphorus helped to raise the question of what it meant for a substance (any given variety of matter) to be an element, in a world where versions of atomic theory were only speculative and later understandings of the nature of substances were only beginning to become possible.

In 1718, Étienne François Geoffroy's Affinity Table made use of several aspects — (1) tabular grouping and (2) correlation with chemical affinity — that would later be reprised. In 1789, French chemist Antoine Lavoisier wrote Traité Élémentaire de Chimie (Elementary Treatise of Chemistry), which is considered to be the first modern textbook about chemistry. Lavoisier defined an element as a substance whose smallest units cannot be broken down into a simpler substance. Lavoisier's book contained a list of "simple substances" that Lavoisier believed could not be broken down further, which included oxygen, nitrogen, hydrogen, phosphorus, mercury, zinc and sulfur. This formed the basis for the modern list of elements.

Lavoisier's list also included "light" and "caloric", which at the time were believed to be material substances. While many leading chemists refused to believe Lavoisier's new revelations, the Elementary Treatise was written well enough to convince the younger generation. However, Lavoisier's descriptions of his elements lack completeness, as he only classified them as metals and non-metals. In 1808–10, British natural philosopher John Dalton published a method by which to arrive at provisional atomic weights for the elements known in his day, from stoichiometric measurements and reasonable inferences. Dalton's atomic theory was adopted by many chemists during the 1810s and 1820s. In 1815, British physician and chemist William Prout noticed that atomic weights seemed to be multiples of that of hydrogen.

In 1817, German physicist Johann Wolfgang Döbereiner began one of the earliest attempts to classify the elements. In 1829, he found that he could form some of the elements into groups of three, with the members of each group having related properties. Chlorine, bromine, and iodine formed a triad; as did calcium, strontium, and barium; lithium, sodium, and potassium; and sulfur, selenium, and tellurium. Various chemists continued his work and were able to identify more and more relationships between small groups of elements.

However, they could not build one scheme that encompassed them all. In 1843, building on work done by Döbereiner, Leopold Gmelin developed a forerunner of the modern periodic table that listed 55 chemical elements grouped by common characteristics. Even though Amedeo Avogadro and, independently of him, André-Marie Ampère, proposed the solution in the form of diatomic molecules and Avogadro's law already in the 1810s, it was not until after Stanislao Cannizzaro's publications in late 1850s when the theory began to be widely considered. In 1860, the modern scientific consensus emerged at the first international chemical conference, the Karlsruhe Congress, and a revised list of elements and atomic masses was adopted. It helped spur creation of more extensive systems.

The first such system emerged in two years. Properties of the elements, and thus properties of light and heavy bodies formed by them, are in a periodic dependence on their atomic weight. French geologist Alexandre-Émile Béguyer de Chancourtois noticed that the elements, when ordered by their atomic weights, displayed similar properties at regular intervals. In 1862, he devised a three-dimensional chart, named the "telluric helix", after the element tellurium, which fell near the center of his diagram. With the elements arranged in a spiral on a cylinder by order of increasing atomic weight, de Chancourtois saw that elements with similar properties lined up vertically.

The original paper from Chancourtois in Comptes rendus de l'Académie des Sciences did not include a chart and used geological rather than chemical terms. In 1863, he extended his work by including a chart and adding ions and compounds. John Newlands published a letter in the Chemical News in February 1863 on the periodicity among the chemical elements. The next attempt was made in 1864. This last option has nonetheless been criticized by the chemist and philosopher of science Eric Scerri on the grounds that it appears to imply that hydrogen is above the periodic law altogether, unlike all the other elements.

The first periodic table to become generally accepted was that of the Russian chemist Dmitri Mendeleev in 1869; he formulated the periodic law as a dependence of chemical properties on atomic mass. In a paper dated December 1869 which appeared early in 1870, Meyer published a new periodic table of 55 elements, in which the series of periods are ended by an element of the alkaline earth metal group. The paper also included a line chart of relative atomic volumes, which illustrated periodic relationships of physical characteristics of the elements, and which assisted Meyer in deciding where elements should appear in his periodic table. By this time he had already seen the publication of Mendeleev's first periodic table, but his work appears

to have been largely independent. In 1869, Russian chemist Dmitri Mendeleev arranged 63 elements by increasing atomic weight in several columns, noting recurring chemical properties across them.

Mendeleev used the trends he saw to suggest that atomic weights of some elements were incorrect, and accordingly changed their placements: for instance, he figured there was no place for a trivalent beryllium with the atomic weight of 14 in his work, and he cut both the atomic weight and valency of beryllium by a third, suggesting it was a divalent element with the atomic weight of 9.4. Mendeleev widely distributed printed broadsheets of the table to various chemists in Russia and abroad. Mendeleev argued in 1869 there were seven types of highest oxides. Compared to the rest of the work, Mendeleev's 1869 list misplaces seven then known elements: indium, thorium, and five rare-earth metals: yttrium, cerium, lanthanum, erbium, and didymium. The last two were later found to be mixtures of two different elements; ignoring those would allow him to restore the logic of increasing atomic weight.

These elements (all thought to be divalent at the time) puzzled Mendeleev as they did not show a regular increase in valency despite their seemingly consequential atomic weights. He measured the heat capacity of indium, uranium, and cerium to demonstrate their higher assumed valency, which was soon confirmed by Prussian chemist Robert Bunsen. Mendeleev grouped them together, thinking of them as of a particular kind of series. Within 100 years of the appearance of Mendeleev's table in 1869, Edward G. Mazurs had collected an estimated 700 different published versions of the periodic table. Many forms retain the rectangular structure, including Charles Janet's left-step periodic table (pictured below), and the modernised form of Mendeleev's original 8-column layout that is still common in Russia. Mendeleev continued to improve his ordering; in 1870, it gained a tabular shape, and each column was given its own highest oxide, and in 1871, he further developed it and formulated what he termed the "law of periodicity".

Some changes also occurred with new revisions, with some elements changing positions. In 1870, he first tried to characterize the yet undiscovered elements, and he gave detailed predictions for three elements, which he termed eka-boron, eka-aluminium, and eka-silicium; he also more briefly noted a few other expectations. In early 1870, he decided that the weights for these elements must be wrong and that the rare-earth metals should be trivalent (which accordingly increased their predicted atomic weights by half). In 1871, Mendeleev expanded his predictions further. Having had no success in that, he abandoned his attempts to incorporate the rare-earth metals in late 1871 and embarked on his grand idea of luminiferous ether.

His idea was carried on by Austrian-Hungarian chemist Bohuslav Brauner, who sought to find a place in the periodic table for the rare-earth metals; Mendeleev later referred to him as to "one of the true consolidators of the periodic law". In addition to the predictions of scandium, gallium, and germanium that were quickly realized, Mendeleev's 1871 table left many more spaces for undiscovered elements, though he did not provide detailed predictions of their properties. In total, he predicted eighteen elements, though only half corresponded to elements that were later discovered. In

the years following Mendeleev's publication, the Indian chemist Prafulla Chandra Ray began working with English geologist Thomas Henry Holland to search for "certain rare Indian minerals in the expectation that one or two new elements might turn up and thus fill up the gap in Mendeleev's periodic table." None of the proposals were accepted immediately, and many contemporary chemists found it too abstract to have any meaningful value.

Of those chemists that proposed their categorizations, Mendeleev strove to back his work and promote his vision of periodicity, Meyer did not promote his work very actively, and Newlands did not make a single attempt to gain recognition abroad. Both Mendeleev and Meyer created their respective tables for their pedagogical needs; the difference between their tables is well explained by the fact that the two chemists sought to use a formalized system to solve different problems. Mendeleev's intent was to aid composition of his textbook, Foundations of Chemistry, whereas Meyer was rather concerned with presentation of theories. Mendeleev's predictions emerged outside of the pedagogical scope in the realm of journal science, while Meyer made no predictions at all and explicitly stated his table and his textbook it was contained in, Modern Theories, should not be used for prediction in order to make the point to his students to not make too many purely theoretically constructed projections. Mendeleev and Meyer differed in temperament, at least when it came to promotion of their respective works.

Boldness of Mendeleev's predictions was noted by some contemporary chemists, however skeptical they may have been. Meyer referred to Mendeleev's "boldness" in an edition of Modern Theories, whereas Mendeleev mocked Meyer's indecisiveness to predict in an edition of Foundations of Chemistry. Eventually, the periodic table was appreciated for its descriptive power and for finally systematizing the relationship between the elements, although such appreciation was not universal. Mendeleev's eka-aluminium was discovered in 1875 and became known as gallium; eka-boron and eka-silicium were discovered in 1879 and 1886, respectively, and were named scandium and germanium. Mendeleev was even able to correct some initial measurements with his predictions, including the first prediction of gallium, which matched eka-aluminium fairly closely but had a different density.

Mendeleev advised the discoverer, French chemist Paul-Émile Lecoq de Boisbaudran, to measure the density again; de Boisbaudran was initially skeptical (not least because he thought Mendeleev was trying to take credit from him) but eventually admitted the correctness of the prediction. Mendeleev contacted all three discoverers; all three noted the close similarity of their discovered elements with Mendeleev's predictions, with the last of them, German chemist Clemens Winkler, admitting this suggestion was not first made by Mendeleev or himself after the correspondence with him, but by a different person, German chemist Hieronymous Theodor Richter. Some contemporary chemists were not convinced by these discoveries, noting the dissimilarities between the new elements and the predictions or claiming those similarities that did exist were coincidental. However, success of Mendeleev's predictions helped spread the word about his periodic table. Later, chemists used the successes of these Mendeleev's predictions to justify his table.

In 1881, Mendeleev claimed priority over Meyer for the periodic table as follows: That person is rightly regarded as the creator of a particular scientific idea who perceives not merely its philosophical, but its real aspect, and who understands so to illustrate the matter so that everyone can become convinced of its truth. Then alone the idea, like matter, becomes indestructible. Even as Mendeleev corrected positions of some elements, he thought that some relationships that he could find in his grand scheme of periodicity could not be found because some elements were still undiscovered, and that the properties of such undiscovered elements could be deduced from their expected relationships with other elements. In 1881, Mendeleev and Meyer had an argument via an exchange of articles in British journal Chemical News over priority of the periodic table, which included an article from Mendeleev, one from Meyer, one of critique of the notion of periodicity, and many more. In 1882, the Royal Society in London awarded the Davy Medal to both Mendeleev and Meyer for their work to classify the elements; although two of Mendeleev's predicted elements had been discovered by then, Mendeleev's predictions were not at all mentioned in the prize rationale.

Eka-silicon was found in 1886 by German chemist Clemens Winkler, who named it germanium. The properties of gallium, scandium, and germanium matched what Mendeleev had predicted. Mendeleev believed it to be a very light gas, with an atomic weight several orders of magnitude smaller than that of hydrogen. He also postulated that it would rarely interact with other elements, similar to the noble gases of his group zero, and instead permeate substances at a velocity of 2,250 kilometers (1,400 mi) per second. Mendeleev was not satisfied with the lack of understanding of the nature of this periodicity; this would only be possible through the understanding of the composition of the atom.

In 1889, Mendeleev noted at the Faraday Lecture to the Royal Institution in London that he had not expected to live long enough "to mention their discovery to the Chemical Society of Great Britain as a confirmation of the exactitude and generality of the periodic law". In the same lecture he allowed that "we already feel that we have neared the moment when this [periodic] law begins to change, and change fast". Even the discovery of the noble gases at the close of the 19th century, which Mendeleev had not predicted, fitted neatly into his scheme as an eighth main group. Mendeleev nevertheless had some trouble fitting the known lanthanides into his scheme, as they did not exhibit the periodic change in valencies that the other elements did. By 1890, Mendeleev's periodic table had been universally recognized as a piece of basic chemical knowledge. Apart from his own correct predictions, a number of aspects may have contributed to this, such as the correct accommodation of many elements whose atomic weights were thought to have wrong values but were later corrected.

The debate on the position of the rare-earth metals helped spur the discussion about the table as well. The great value of Newland's, Mendeleef's, and Lothar Meyer's generalisation, known as the periodic arrangement of the elements, is universally acknowledged. Although Mendeleev knew of lanthanum, cerium, and erbium, they were previously unaccounted for in the table because their total number and exact order were not known; Mendeleev still could not fit them in his table by 1901. This was

in part a consequence of their similar chemistry and the imprecise determination of their atomic masses. Mendeleev himself added these elements to the table as group 0 in 1902, without disturbing the basic concept of the periodic table.

However, Mendeleev firmly believed that future would only develop the notion rather than challenge it and reaffirmed his belief in writing in 1902. After much investigation, the Czech chemist Bohuslav Brauner suggested in 1902 that the lanthanides could all be placed together in one group on the periodic table. Mendeleev treated the change by assessing each element to an individual place in his system of the elements rather than continuing to treat them as a series. Mendeleev noticed that there was a significant difference in atomic mass between cerium and tantalum with no element between them; his consideration was that between them, there was a row of yet undiscovered elements, which would display similar properties to those elements which were to be found above and below them: for instance, an eka-molybdenum would behave as a heavier homolog of molybdenum and a lighter homolog of wolfram (the name under which Mendeleev knew tungsten). However, the higher valency for didymium had not been established, and Mendeleev tried to do so himself.

By 1904, Mendeleev's table rearranged several elements, and included the noble gases along with most other newly discovered elements. He determined that the rare-earth elements (lanthanides), 13 of which were known, lay within that gap. The contemporarily accepted discovery of element 75 came in 1925, when Walter Noddack, Ida Tacke, and Otto Berg independently rediscovered it and gave it its present name, rhenium.) British chemist John Newlands presented in Chemical News a classification of the 62 known elements. Newlands noticed recurring trends in physical properties of the elements at recurring intervals of multiples of eight in order of mass number; based on this observation, he produced a classification of these elements into eight groups.

Each group displayed a similar progression; Newlands likened these progressions to the progression of notes within a musical scale. Newlands's table left no gaps for possible future elements, and in some cases had two elements at the same position in the same octave. Newlands's table was ignored or ridiculed by some of his contemporaries. The Chemical Society refused to publish his work. The president of the society, William Odling, defended the society's decision by saying that such "theoretical" topics might be controversial; there was even harsher opposition from within the society, suggesting the elements could have been just as well listed alphabetically.

Later that year, Odling suggested a table of his own but failed to get recognition following his role in opposing Newlands's table. German chemist Lothar Meyer also noted the sequences of similar chemical and physical properties repeated at periodic intervals. But a study of this arrangement, it must be allowed, is a somewhat tantalising pleasure; for, although the properties of elements do undoubtedly vary qualitatively, and, indeed, show approximate quantitative relations to their position in the periodic table, yet there are inexplicable deviations from regularity, which hold forth hopes of the discovery of a still more far-reaching generalisation. What that

generalisation may be is not yet to be divined; but that it must underlie what is known, and must furnish a clue to the explanation of irregularities, cannot be disputed.

Chapter 2: Understanding The Periodic Table

Hence the periodic table is usually drawn to begin each row (often called a period) with the filling of a new s orbital, which corresponds to the beginning of a new shell. Thus, with the exception of the first row, each period length appears twice: 2, 8, 8, 18, 18, 32, 32,... Furthermore, national variations are sometimes encountered: in Japan, alkaline earth metals often do not include beryllium and magnesium as their behaviour is different from the heavier group 2 metals. Similar anomalies arise for the 1s, 2p, 3d, 4f, and the hypothetical 5g elements: the degree of this first-row anomaly is highest for the s-block, is moderate for the p-block, and is less pronounced for the d- and f-blocks.

For example, all elements in group (column) 18 are noble gases that are largely—though not completely—unreactive. There are some other relationships throughout the periodic table between elements that are not in the same group, such as the diagonal relationships between elements that are diagonally adjacent (e.g. lithium and magnesium). Many more categorizations exist and are used according to certain disciplines. A few terms are widely used, but without any very formal definition, such as "heavy metal", which has been given such a wide range of definitions that it has been criticized as "effectively meaningless". The scope of terms varies significantly between authors.

Chapter 3: Key Characteristics of The Periodic Table

Metallic character increases going down a group and from right to left across a period. Nonmetallic character increases going from the bottom left of the periodic table to the top right. Atomic radii (the size of atoms) are dependent on the sizes of their outermost orbitals. They generally decrease going left to right along the main-group elements, because the nuclear charge increases but the outer electrons are still in the same shell. However, going down a column, the radii generally increase, because the outermost electrons are in higher shells that are thus further away from the nucleus. For example, the 2p orbitals do not experience strong repulsion from the 1s and 2s orbitals, which have quite different angular charge distributions, and hence are not very large; but the 3p orbitals experience strong repulsion from the 2p orbitals, which have similar angular charge distributions. Thus higher s-, p-, d-, and f-subshells experience strong repulsion from their inner analogues, which have approximately the same angular distribution of charge, and must expand to avoid this. This makes significant differences arise between the small 2p elements, which prefer multiple bonding, and the larger 3p and higher p-elements, which do not. In the transition elements, an inner shell is filling, but the size of the atom is still determined by the outer electrons. The trend that atomic radii decrease from left to right is also present in ionic radii, though it is more difficult to examine because the most common ions of

consecutive elements normally differ in charge.

An element's electronegativity varies with the identity and number of the atoms it is bonded to, as well as how many electrons it has already lost: an atom becomes more electronegative when it has lost more electrons. This sometimes makes a large difference: lead in the +2 oxidation state has electronegativity 1.87 on the Pauling scale, while lead in the +4 oxidation state has electronegativity 2.33. The simple substances of the more electronegative atoms tend to share electrons (form covalent bonds) with each other. Thus uranium somewhat resembles chromium and tungsten in group 6, as all three have six valence electrons. Relationships between elements with the same number of valence electrons but different types of valence orbital have been called secondary or isodonor relationships: they usually have the same maximum oxidation states, but not the same minimum oxidation states.

For example, chlorine and manganese both have +7 as their maximum oxidation state (e.g. Cl2O7 and Mn2O7), but their respective minimum oxidation states are −1 (e.g. HCl) and −3 (K3[Mn(CO)4]). Thus in period 6, Cs–Ba have 1–2 valence electrons; La–Yb have 3–16; Lu–Hg have 3–12; and Tl–Rn have 3–8. For example, magnesium forms Mg2+ rather than Mg+ cations when dissolved in water, because the latter would spontaneously disproportionate into Mg0 and Mg2+ cations. This is because the enthalpy of hydration (surrounding the cation with water molecules) increases in magnitude with the charge and radius of the ion.

For similar reasons, the common oxidation states of the heavier p-block elements (where the ns electrons become lower in energy than the np) tend to vary by steps of 2, because that is necessary to uncover an inner subshell and decrease the ionic radius (e.g. Analogous arguments based on orbital hybridization can be used for the less electronegative p-block elements. For transition metals, common oxidation states are nearly always at least +2 for similar reasons (uncovering the next subshell); this holds even for the metals with anomalous dx+1s1 or dx+2s0 configurations (except for silver), because repulsion between d-electrons means that the movement of the second electron from the s- to the d-subshell does not appreciably change its ionisation energy. Because ionizing the transition metals further does not uncover any new inner subshells, their oxidation states tend to vary by steps of 1 instead. The common and maximum oxidation states of the d- and f-block elements tend to depend on the ionisation energies.

Thus, the early transition metal groups tend to prefer higher oxidation states, but the +2 oxidation state becomes more stable for the late transition metal groups. The highest formal oxidation state thus increases from +3 at the beginning of each d-block row, to +7 or +8 in the middle (e.g. OsO4), and then decrease to +2 at the end. The more electronegative atom will tend to attract the electron pair more, and the less electronegative (or more electropositive) one will attract it less. In extreme cases, the more electronegative atom takes the shared pair entirely to itself, and the resulting link is an ionic rather than a covalent one.

Chapter 4: The Science Behind The Periodic Table

Four numbers describe an orbital in an atom completely: the principal quantum number n, the azimuthal quantum number ℓ (the orbital type), the orbital magnetic quantum number $m\ell$, and the spin magnetic quantum number ms. The sequence in which the subshells are filled is given in most cases by the Aufbau principle, also known as the Madelung or Klechkovsky rule (after Erwin Madelung and Vsevolod Klechkovsky respectively). This rule was first observed empirically by Madelung, and Klechkovsky and later authors gave it theoretical justification. Phrased differently, electrons enter orbitals in order of increasing $n + \ell$, and if two orbitals are available with the same value of $n + \ell$, the one with lower n is occupied first. In general, orbitals with the same value of $n + \ell$ are similar in energy, but in the case of the s orbitals (with $\ell = 0$), quantum effects raise their energy to approach that of the next $n + \ell$ group.

In hydrogen, there is only one electron, which must go in the lowest-energy orbital 1s. This electron configuration is written 1s1, where the superscript indicates the number of electrons in the subshell. Helium adds a second electron, which also goes into 1s, completely filling the first shell and giving the configuration 1s2. Starting from the third element, lithium, the first shell is full, so its third electron occupies a 2s orbital, giving a 1s2 2s1 configuration. The 2s electron is lithium's only valence electron, as the 1s subshell is now too tightly bound to the nucleus to participate in chemical bonding to other atoms: such a shell is called a "core shell".

The 1s subshell is a core shell for all elements from lithium onward. The 2s subshell is completed by the next element beryllium (1s2 2s2). The following elements then proceed to fill the 2p subshell. Boron (1s2 2s2 2p1) puts its new electron in a 2p orbital; carbon (1s2 2s2 2p2) fills a second 2p orbital; and with nitrogen (1s2 2s2 2p3) all three 2p orbitals become singly occupied. This is consistent with Hund's rule, which states that atoms usually prefer to singly occupy each orbital of the same type before filling them with the second electron.

Starting the next row, for potassium and calcium the 4s subshell is the lowest in energy, and therefore it fills next. Potassium adds one electron to the 4s shell ([Ar] 4s1), and calcium then completes it ([Ar] 4s2). However, starting from scandium ([Ar] 3d1 4s2) the 3d subshell becomes the next highest in energy. Such anomalies, however, do not have any chemical significance: most chemistry is not about isolated gaseous atoms, and the various configurations are so close in energy to each other that the presence of a nearby atom can shift the balance. Therefore, the periodic table ignores them and considers only idealized configurations.

At zinc ([Ar] 3d10 4s2), the 3d orbitals are completely filled with a total of ten electrons. Next come the 4p orbitals, completing the row, which are filled progressively by gallium ([Ar] 3d10 4s2 4p1) through krypton ([Ar] 3d10 4s2 4p6), in a manner analogous to the previous p-block elements. From gallium onwards, the 3d orbitals form part of the electronic core, and no longer participate in chemistry. The overlaps get quite close at the point where the d orbitals enter the picture, and the order can shift slightly with atomic number and atomic charge. Starting from the simplest atom, this lets us build up the periodic table one at a time in order of atomic

number, by considering the cases of single atoms.

In the transition series, the outer electrons are preferentially lost even though the inner orbitals are filling. For example, in the 3d series, the 4s electrons are lost first even though the 3d orbitals are being filled. A full explanation requires considering the energy that would be released in forming compounds with different valences rather than simply considering electron configurations alone. Tl+ uncovers 6s, and Tl3+ uncovers 5d, so once thallium loses two electrons it tends to lose the third one as well). As the energy difference between the (n−1)d and ns orbitals rises along each transition series, it becomes less energetically favourable to ionize further electrons.

The bond then binds two ions, one positive (having given up the electron) and one negative (having accepted it), and is termed an ionic bond. In 2023, Smits, Düllmann, Indelicato, Nazarewicz, and Schwerdtfeger made another attempt to place elements from 119 to 170 in the periodic table based on their electron configurations.

Chapter 5: Observation and Measurement of The Periodic Table

The problem of placing isotopes in the periodic table had arisen beginning in 1900 when four radioactive elements were known: radium, actinium, thorium, and uranium. These radioactive elements (termed "radioelements") were accordingly placed at the bottom of the periodic table, as they were known to have greater atomic weights than stable elements, although their exact order was not known. Researchers believed there were still more radioactive elements yet to be discovered, and during the next decade, the decay chains of thorium and uranium were extensively studied. Following British physicist Charles Glover Barkla's discovery of characteristic X-rays emitted from metals in 1906, British physicist Henry Moseley considered a possible correlation between X-ray emissions and physical properties of elements.

In 1907 it was discovered that thorium and radiothorium, products of radioactive decay, were physically different but chemically identical; this led Frederick Soddy to propose in 1910 that they were the same element but with different atomic weights. Soddy later proposed to call these elements with complete chemical identity "isotopes". The significance of these atomic properties was determined in the Geiger–Marsden experiments, in which the atomic nucleus and its charge were discovered, conducted between 1908 and 1913. By 1912, almost 50 different radioactive substances had been found in the decay chains of thorium and uranium. American chemist Bertram Boltwood proposed several decay chains linking these radioelements between uranium and lead.

These were thought at the time to be new chemical elements, substantially increasing the number of known "elements" and leading to speculations that their discoveries would undermine the concept of the periodic table which had long been established to obey the octet rule. For example, there was not enough room between lead and uranium to accommodate these discoveries, even assuming that some discoveries

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