

J. J. Thomson and the Bohr atom

Far from being merely "scientific curiosities," J. J. Thomson's seemingly naive models actually contained some of the fundamental ideas of Niels Bohr's revolutionary quantum theory of the atom.

John L. Heilbron

In 1911 Niels Bohr went to Cambridge, hoping to talk physics with J. J. Thomson; the discoverer of the electron was friendly but uninterested. Two years later Bohr published his epochal three-part paper on the constitution of atoms and molecules, which challenged the program and goal of the Cambridge school. Bohr's new views soon won out; Thomson's quaint atomic models were declared worthless—old lumber fit only, as Ernest Rutherford put it, "for a museum of scientific curiosities." For his part Thomson rejected the advances made by Bohr as meretricious superficialities obtained without, or at the price of, an understanding of the mechanism of atoms.

As in many other instances in the history of science, Bohr's revolutionary theory became such a success that its origins in the views it superseded were all but forgotten. In particular, Thomson's opposition and the quick replacement of his research program by Bohr's obscured the connection between the theory of the quantized atom and the deceptively simple and apparently naive models of the Cambridge school. So has the odd circumstance that the three installments of Bohr's first paper on atomic structure inverted the order of his discoveries. The first installment, the only one now remembered, gives the theory of the Balmer spectrum, which Bohr worked out in a few weeks in February 1913; the other two record Bohr's attempts, beginning in June 1912, to bring Rutherford's nuclear model—itsself a product of Thomson's research program—to bear on the chief problems of atomic theory as Thomson had identified them.

To Thomson the key problem in atomic theory was the explanation of the variation in the periodic properties of the chemical elements represented in Mendeleev's table. Already in 1897, when announcing the discovery of the electron, he intimated that the new particles might well provide this periodicity when they are bound into an atom. Not then knowing how this might be accomplished, he resorted to the sort of analogy characteristic of the Cambridge school of mathematical physics during Thomson's time.

Magnets and a plum pudding

As an analogue to the arrangement of electrons in an atom, Thomson offered Alfred Mayer's floating magnets, which distribute themselves into concentric circles under the influence of a large stationary magnet, as shown in figure 1. In 1903, having secured the electron, measured its charge and mass, and laid the foundation of the electron theory of metals, Thomson took up the question how his favorite corpuscle could play the part of Mayer's magnets.

The first problem was to choose a representation for the positive portion of the atom. The arrangement that is perhaps the most obvious, the nuclear model, had already been proposed and discarded on the ground of mechanical instability: In any Saturnian atom—one with several electrons arranged in a plane ring or rings—there exists at least one unstable mode of oscillation about the equilibrium orbits. The amplitudes of these unstable modes grow until the system flies apart. However, a stable variant can be obtained by allowing the positive charge to fill the entire volume of the atom; the electrons then circulate within the positive charge, subject to a restoring force varying directly as the distance rather than as its

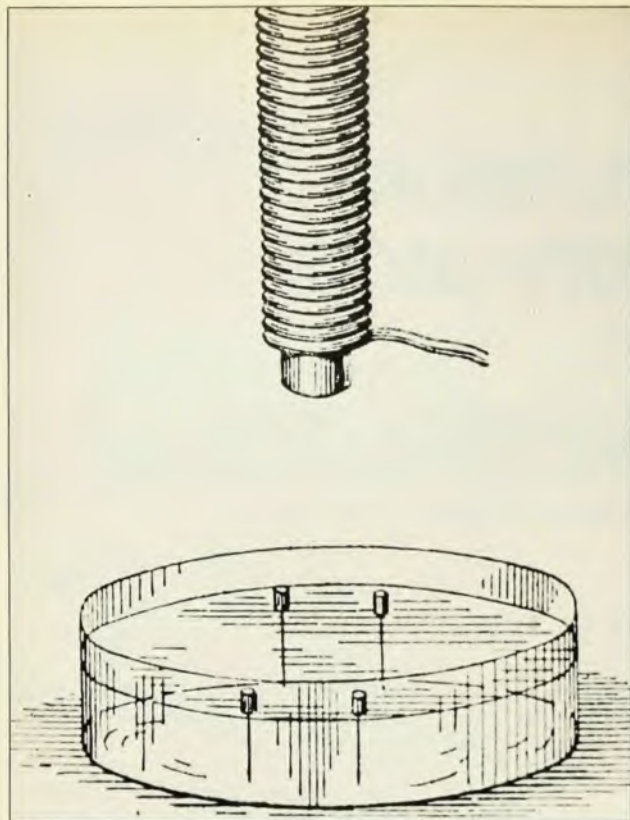
inverse square. This so-called "plum-pudding model" is the one Thomson adopted.

Note that the instability that led to the initial rejection of the nuclear model was a mechanical one: It did not derive from that drain of energy by radiation that plays so important a role in the standard historical accounts. Indeed, as Thomson showed, the total radiation from a ring of p symmetrically placed electrons describing the same circular orbit decreases very rapidly as p increases; for moderate values of p the ring—and hence the atom—has almost eternal life.

Even the eventual mortality of atoms was no inconvenience to Thomson: He merely associated radioactivity with ancient atoms, the internal motions of which had decayed to the point of instability and explosion. At this time (1904) he thought that the atom contained a great many electrons, perhaps—as the richness of spectral lines and the ratio of the masses of the electron and the hydrogen ion suggest—as many as a thousand times the atomic weight. He did not lack particles to populate his rings and plug the radiation drain.

The urge of individual electrons in an atom to radiate can therefore be curbed by the social pressure of their neighbors. But this pressure can not be driven too far: Electrons are not friendly; they repel one another. When enough of them are assembled in a ring to extinguish their radiation, there may be too many for mechanical stability; a little disturbance to any one of them might cause the ring to fly apart. Thomson conceived the idea that the condition of mechanical stability might be the clue to the periodicity in the electronic arrangements of the atoms. The electrons' need for elbow room might fix their population distribution. In 1904 he put this idea to the test.

John L. Heilbron is professor of history and director of the Office for History of Science and Technology, University of California, Berkeley.



Mayer's magnets—magnetized needles floated on corks, under a large stationary magnet—provided J. J. Thomson with an analogy to the arrangements of electrons in atoms. These diagrams, made by pressing paper against the inked tops of the magnets, displayed stable configurations with a periodicity suggestive of Mendeleev's table. From A. M. Mayer, *Am. J. Science* **116**, 248 (1878).

Figure 1

The heart of Thomson's analysis was the calculation of the frequencies of the perturbed oscillations of the electrons in a single-ring atom as a function of their number p . He hoped to learn from the frequencies how large p might be before mechanical instability set in: The number turned out to be six. To accommodate more electrons in a single ring, the rate at which the restoring force varied with distance had to be greater than that afforded by the diffuse charge alone.

Rings of electrons

Nothing could be simpler than increasing this rate: One needed merely to put one or more electrons (q in all, say) at the atom's center. Thomson calculated the values of q that would result in a stable outer ring of p electrons. It turned out that the inner electrons themselves must be distributed in rings, and that for each value of the total electron population, $n = p + q$, the distribution is unique.

This distribution represents an electronic parallel to Mayer's magnets, but one that is far more suggestive of the physics of atoms. Thomson shows that if $p = 20$, q must lie between 39 and 47, inclusive; his results are presented in Table 1. If q is close to the minimum, the

atom could increase its stability by losing one of its 20 outer electrons; such an atom would act electropositively. If q is near a maximum, the atom would tend to gain an electron, and therefore act electronegatively. The models characterized by $p = 20$ consequently offered a striking analogy to the elements of the second and of the third periods of Mendeleev's table.

It was this elucidation of the periodic table, expanded and translated into German, that brought continental physicists an inkling that something might come from the Cambridge theory of atomic structure. In 1909 Max Born thought Thomson's model sufficiently promising to take it as the subject of his inaugural lecture as *Privatdozent*, and in 1911 Arnold Sommerfeld's physics colloquium studied it with the help of floating magnets.

"If it resembled a little, it was so"

Three points about Thomson's analogy deserve attention:

- ▶ He has introduced the fundamental idea that atoms of successive elements in the periodic table differ from one another by the addition of a single electron.
- ▶ He has, from a modern point of view, interchanged the roles of core and valence

electrons. The atoms of each period are characterized by the same number of external electrons, and differ only in the populations of their inner rings. Chemical and optical properties consequently derive primarily from the deeper-lying electrons; the members of a chemical family have only internal structures in common. Likewise all the electrons in the atom, and not just the deepest, are implicated in radioactivity, and it is therefore difficult to find room in Thomson's scheme for structures with identical chemical and different radioactive properties. The existence of isotopes, as Bohr later emphasized, could not be explained plausibly on the basis of the diffuse-sphere atom.

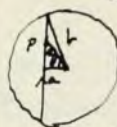
▶ Lastly, despite the mathematical labor that secured it, Thomson's analogy was essentially qualitative. Here we reach a perplexing and perennial characteristic of Thomson's physics. At the very beginning of his career, in 1882, he had won the prestigious Adams Prize at Cambridge for a lengthy essay on Kelvin's vortex atoms. To describe encounters between such atoms, which resembled smoke rings in air, required severe and rigorous calculations, the application of which to physical or chemical phenomena proved all but impossible. Already then Thomson had

Theory of structure of atoms

Suppose atom consists of $+ve$ charge Ne at centre & $-ve$ charge e distributed throughout sphere of radius a .

$$\text{Force at } p \text{ on electron} = Ne^2 \left\{ \frac{1}{a^2} - \frac{1}{a^3} \frac{1}{a} \right\} = Ne^2 \left\{ \frac{1}{a^2} - \frac{1}{a^3} \right\} = \frac{Ne^2}{a^3} \left(a - 1 \right)$$

Suppose charged particles e move in circles through atom so that displacement is small but ω distance from centre = a



Deflection force ω directed towards centre at p

$$= Ne^2 \left\{ \frac{1}{a^2} - \frac{1}{a^3} \right\} \cos \theta$$

$$\text{Actual } \omega \text{ directed } \omega \text{ within} = d\omega = \frac{Ne^2}{m} \left(\frac{1}{a^2} - \frac{1}{a^3} \right) \frac{a}{2}$$

∴ Force ω required to keep things stable ω directed

$$u = \int d\omega \cdot d\omega = \frac{Ne^2}{m} \left(\frac{1}{a^2} - \frac{1}{a^3} \right) \frac{a}{2}$$

$$= \frac{Ne^2}{m} \left(\frac{1}{a^2} - \frac{1}{a^3} \right) \frac{a}{2} = \frac{Ne^2}{m} \left(\frac{1}{a} - \frac{1}{a^2} \right)$$

$$= \frac{2Ne^2}{m} \left(\frac{1}{a} - \frac{1}{a^2} \right) = \frac{2Ne^2}{m} \left(\frac{1}{a} - \frac{1}{a^2} \right)$$

$$= \frac{2Ne^2}{m} \left(\frac{1}{a} - \frac{1}{a^2} \right) = \frac{2Ne^2}{m} \left(\frac{1}{a} - \frac{1}{a^2} \right)$$

Rutherford's first calculations on the passage of alpha particles through atoms. In his "theory of structure of atoms," Rutherford used a nuclear atom that was a variant of Thomson's model, of electrons in a sphere of positive charge: It had a positive central nucleus of charge ne surrounded by a diffuse sphere of negative electricity. From the Rutherford Papers, Cambridge University Library.

Figure 2

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Margrethe Norlund and Niels Bohr announce their engagement in 1911. That year Bohr defended his thesis at the University of Copenhagen and left for the Cavendish Laboratory. Figure 3

to content himself with the sort of qualitative and suggestive connections he was later to make with his electronic atom. He never identified particular chemical atoms with definite models, whether vortical or electronic. "Things needed not to be very exact for Thomson," Bohr used to say, "and if it resembled a little, it was so."

The most important undetermined parameter in Thomson's model was the total electron number n . On its magnitude depended not only the security of the atom against radiation collapse, but also inferences about the nature of positive charge and the process of spectral emission. Thomson worked on the problem for five or six years, bringing to bear his powerful mathematics and the experimental resources of the Cavendish Laboratory. He was the first to explore the atom by shooting charged particles through it, and the first to work out formulas, including probability considerations where appropriate, for the scat-

tering of x rays and beta rays.

The chief result of comparing the experiments to the formulas was that n was about equal to the atomic weight A . The outcome of Rutherford's variant of Thomson's scattering theory—alpha scattering elucidated by the nuclear atom—was an n of about $A/2$. That the nuclear atom was an outgrowth of Thomson's research program appears plainly from the first page of Rutherford's first calculations on the "theory of the structure of atoms," reproduced in figure 2. Note the depiction of the scatterer as a tiny positive nucleus of charge ne , surrounded by a *diffuse sphere of negative electricity of fixed radius*.

The thousandfold reduction of the atomic population brought Thomson and his co-workers very close to the doctrine of atomic number. It also made acute the problem of the radiation collapse of light atoms. It was quite characteristic of Thomson to acknowledge this unpleasantness and move on; he considered

spectra too complicated to reveal anything useful about atomic structure—and in this opinion too he was followed by Bohr.

Bohr's approach

Bohr came to the problem of atomic structure almost by chance. His subject had been the electron theory of metals, on which he had written a thesis defended at the University of Copenhagen in the spring of 1911. He then went to the Cavendish Laboratory, intending to rework his thesis for English publication. And why the Cavendish?

"I considered first of all Cambridge as the center of physics," Bohr later said of his decision to study there, "and Thomson as a most wonderful man . . . , a genius who showed the way for everybody." Thomson received him politely and promised to read the rough translation of his thesis that Bohr had brought him.

"I have just talked to J. J. Thomson," Bohr wrote his brother after his first interview, "and I explained to him as well as I could my views on radiation, magnetism, etc. You should know what it was for me to talk to such a man. He was so very kind to me; we talked about so many things; and I think he thought there was something in what I said. He has promised to read my thesis, and he invited me to have dinner with him next Sunday at Trinity College, when he will talk to me about it . . ."

The exchange of views Bohr desired did not take place. Thomson, who had long before given up active cultivation of the electron theory, probably never read Bohr's thesis; in any case he did not enjoy having his ancient errors rehearsed by a tenacious foreigner whose English he could scarcely understand. But even had language and divergent interests not been barriers, one doubts that the intellectual communion that Bohr sought could have developed.

For one thing, the imprecise and contradictory analogies Thomson fancied were inadequate for Bohr, who sought coherent, consistent models from which quantitative predictions about experimental results might be drawn. For another, Thomson, though friendly and receptive to questions, worked alone; he seldom solicited his students' views on scientific questions, nor did he develop his own through extended conversations with others. Bohr's life-long practice, on the other hand, was to refine his ideas in lengthy discussions, which often became monologues, with informed individuals. Whether his colloquist was a full collaborator, a sounding board or an amanuensis, he required some human contact at almost every stage of his work.

It is perhaps not too fanciful to see a reflection of their styles in their photographs. Figure 3 shows Bohr about 1911, aged 26, boyish, callow, soft-featured and gentle. With him is one of his aman-

uenses, his future wife Margrethe, who wrote out his first papers on atomic structure. Figure 4 portrays Thomson at about the same time, aged 53. He had not changed much since his discovery of the electron.

"You ask whether J.J. is an old man," Rutherford had written his fiancée in 1896. "He is just 40 and looks quite young, small, rather straggling moustache, short, wears his hair (black) rather long, but has a very clever-looking face, and a very fine forehead and a most radiating smile, or grin as some call it, when he is scoring off anyone."

A little piece of reality

Thomson's indifference by no means deflected Bohr from the pursuit of the electron theory. It was the chief subject of his research throughout the eight months he spent at Cambridge, and it remained so during the first three months of his stay at Manchester, where he moved in March 1912, to learn something of the experimental side of radioactivity. It is important to recognize that Bohr did not go to Manchester, Rutherford's citadel, to help develop the consequences of the nuclear atom. He went to take a six-week course on experimental technique, a standard service of the laboratory for beginners in radioactivity, after which they usually began a small research task proposed by Rutherford. Figure 5 shows a page of Bohr's carefully kept laboratory notebook.

It was not that Bohr wished to become an experimentalist: His object was to capitalize on his time in England, and to make contact with Rutherford, evidently the coming power in English physics. After finishing the laboratory work for the day he would return to the electron theory of metals.

Bohr came to atomic physics in a casual way. The research topic Rutherford had assigned him was interrupted for want of radium emanation (radon). While waiting for more to grow he studied a paper on the absorption of alpha particles that had just been published by C. G. Darwin, the only mathematical physicist besides himself in Rutherford's group. Bohr found that Darwin's treatment rested on an unsatisfactory assumption about the interaction between alpha particles and atomic electrons: Darwin had ignored the binding forces. Bohr, following a technique used by Thomson, proposed to take the forces into account by treating the interaction as a resonance phenomenon depending on the ratio of $1/\nu'$, the natural period of the electrons' vibrations about equilibrium, to the time required by an alpha particle to pass the atom.

Bohr expected to make an easy calculation, which would quickly furnish a short note for the *Philosophical Magazine*; that was in early June, 1912. By the middle of the month he had abandoned the laboratory, shelved the electron



J. J. Thomson in 1909. In 1896 Rutherford had written of his "most radiating smile . . . when scoring off anyone." (Photo in G. P. Thomson, *J. J. Thomson*, Doubleday, 1965) **Figure 4**

theory and given himself up entirely to the design of atomic models. A letter from Bohr to his brother Harald, dated 12 June 1912, gives a clue to what happened:

"It could be that I've perhaps found out a little bit about the structure of atoms. You must not tell anyone anything about it; otherwise I certainly could not write you this soon. If I'm right, it would not be an indication of the nature of a possibility (like J. J. Thomson's theory) but perhaps a little piece of reality. It has all grown out of a little piece of information I obtained from the absorption of alpha particles . . . You can imagine how anxious I am to finish quickly and I've stopped going to the laboratory for a couple of days to do so (that's also a secret)."

And what was the "little piece of information"? It may well have been the discovery that the nuclear atom is mechanically unstable.

Thomson and the Cambridge school had rejected the nuclear model on account

of its mechanical instability; Bohr welcomed it precisely because it needed a nonmechanical force to exist. Already in his Copenhagen dissertation he had pointed to certain phenomena—heat radiation and paramagnetism in particular—that eluded the electron theory and appeared to require the ascription of a nonmechanical rigidity to the paths of atomic electrons. He was drawn to the nuclear model as a possible representation or reification of the sorts of difficulties he had encountered in his earlier studies.

Bohr's fiat

To make further progress possible he exempted, by fiat, electrons that describe closed orbits satisfying the condition

$$T = K\nu' \quad (1)$$

(where T is the electron's kinetic energy, ν' its orbital frequency and K a constant) from the ordinary necessities of their existence: They did not radiate energy and they did not respond to small perturbations. Electrons so characterized, electrons in their ground or permanent state,

Table 1. A Thomson atom with twenty external electrons

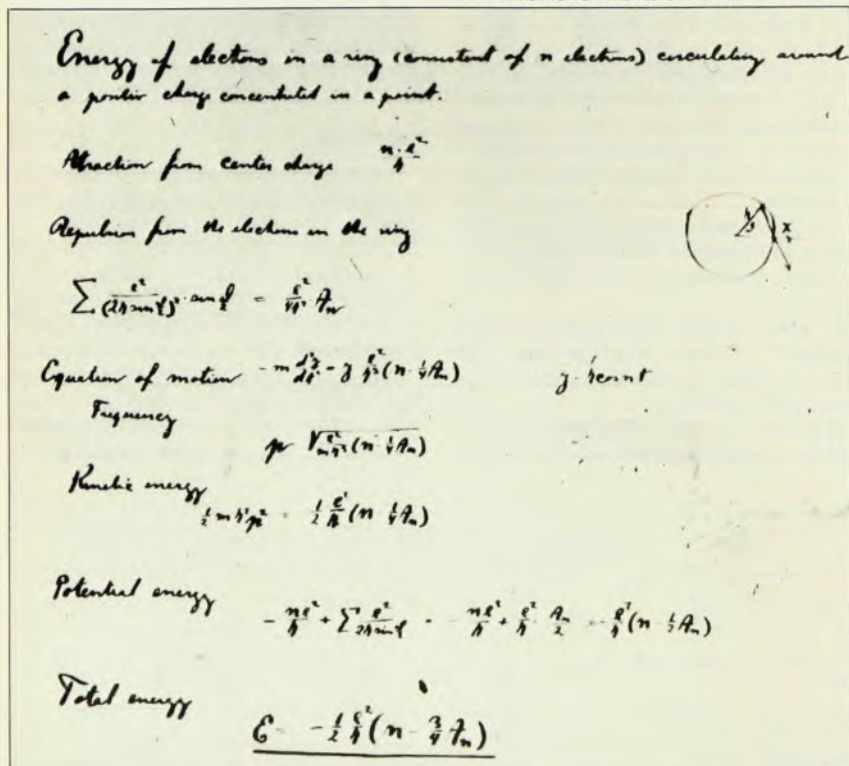
Total number of atomic electrons n	59	60	61	62	63	64	65	66	67
Number in outermost ring p	20	20	20	20	20	20	20	20	20
Number of electrons in successive rings q :	16	16	16	17	17	17	17	17	17
innermost ring at the bottom	13	13	13	13	13	13	14	14	15
	8	8	9	9	10	10	10	10	10
	2	3	3	3	3	4	4	5	5

Adapted from J. J. Thomson, *Phil. Mag.* 7, 237 (1904).



Bohr's laboratory notebook at Manchester 1912. During one experiment he ran out of radon and read a paper that launched him into the problem of atomic structure. Figure 5

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Bohr's calculation of the energy of an n -electron ring, from his "Manchester Memorandum." It contains an error in the potential energy, and hence also in the total energy. Figure 6

are more like beads on a wire than like freely orbiting particles.

There is no doubt that Bohr's introduction of the stability condition marked a fundamental departure from Thomson's program. The form of the condition was chosen in analogy to Max Planck's quantum theory, and with the expectation that K might be a submultiple of Planck's constant h . It turned out that $K = h/2$, a fact Bohr discovered in February 1913, when at last he came to examine Balmer's formula. The resulting account of the Balmer lines and the concept of stationary states forced him to conclude that the frequencies of spectral lines are not the mechanical frequencies of the atoms that emit them.

As we now know, there followed a progressive relaxation of the dominion of mechanics in the microphysical world, culminating in the invention of quantum mechanics and the principles of uncertainty and complementarity. Nothing so radical was in Bohr's mind in June 1912, however. Having taken a step that was to have revolutionary consequences, he immediately turned back to the problems of the Thomsonian atomist.

In June or July of 1912 Bohr drew up the notes now known as the "Manchester Memorandum" for discussion with Rutherford. The Memorandum opens with a definition of the nuclear atom and an acknowledgment of its mechanical instability, which can be demonstrated, as Bohr put it, "by an analysis similar to the one used by Sir J. J. Thomson in his theory about the constitution of an atom." How then can one account for periodicity? This was a pressing problem: No atomic model unable to elucidate Mendeleev's table could decisively defeat Thomson's.

Bohr thought he had a simple solution. He computed the total energy W of each electron in a ring of n electrons, and discovered that W was negative for $n \leq 7$, but positive for $n > 7$. Evidently for $n > 7$ the electrons leave the atom; for $n \leq 7$ they may be bound securely if their motions satisfy condition 1. For an atom with more than seven electrons, several rings will be required; but, in marked contrast to Thomson's model, the additional rings will be formed outside the first, and the population of the outermost will determine the valence of the atom.

"This," said Bohr, "seems to offer a very strong indication of a possible explanation of the periodic law of the chemical properties of the elements."

An error

What is particularly interesting about this analysis—other than the fact that it addresses, as its first order of business, Thomson's central problem—is that it is altogether wrong. Figure 6 shows Bohr's calculations. From the equation of motion, which is correct, it follows that T , the kinetic energy of each electron, is $Q/2r$,

where $Q = e^2(n - A_n/4)$ and $A_n = \sum_{i=1}^{n-1} \csc(\pi i/n)$. (The balance of forces makes $mv^2/r = Q/r^2$.) Bohr's computation of the potential energy U is, however, incorrect; because U belongs to both of the interacting particles, the sum in Bohr's expression for U should be divided by two. Then we have $U = -Q/r$ and

$$W = U + T = -Q/2r = -T$$

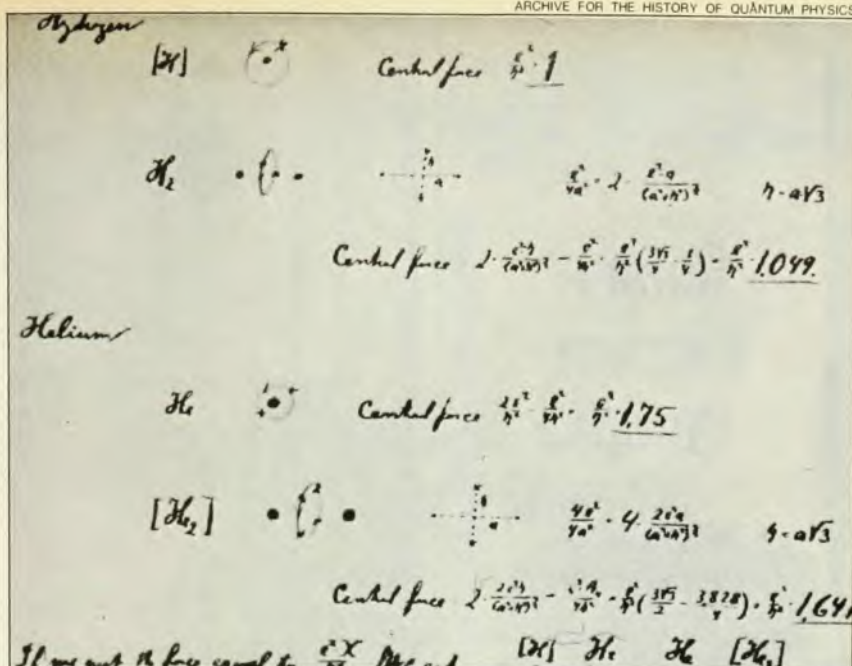
The total energy is the negative of the kinetic energy, and consequently can never change sign. Bohr's error is the more remarkable because his value for the potential energy conflicts both with his expression for the equation of motion and with a theorem proved later in the Memorandum, namely that any particle bound into an orbit by an inverse-square force has a potential energy twice the negative of its kinetic energy. Bohr's slip may betray his anxiety to solve Thomson's problem of periodicity.

For the rest, the Memorandum concerns the structure of simple molecules such as those illustrated in figure 7. Bohr aimed to show, among other things, why the H_2 molecule occurs and He_2 does not, and to demonstrate that no charge is transferred in the combination of identical atoms. He probably took the problem of charge distribution in symmetric diatomic molecules from Thomson's *Corpuscular Theory of Matter*, which gave perhaps the earliest useful explanation of chemical bonding via electron exchange.

Thomson had decided that charge transfer occurs in the formation of H_2 and O_2 because identical plum-pudding atoms can not remain in stable equilibrium. For say they are symmetrically combined, by interpenetration of their positive spheres; any subsequent jostling would create a flow of electrons from one sphere to the other, and a permanent polar bond. Thomson made this conclusion plausible by a characteristic analogy. This system, one of identical water-filled jars suspended from identical springs and connected with a siphon, is unstable; for any relative vertical displacement of the jars will grow with the flow of water through the siphon. Thomson thought the evidence favored asymmetric H_2 and O_2 ; Bohr thought the case for symmetry stronger; hence the considerable attention given to the structure of simple molecules in the Memorandum.

The second and third parts of Bohr's paper of 1913 remain within the set of problems posed by the Memorandum. Part II concerns the problem of the distribution of electrons into rings. Bohr takes for granted the chief result of Thomson's program, the doctrine of atomic number. He then lays down two principles:

► In the ground state of an atom every electron, regardless of its distance from the nucleus, has just one quantum of angular momentum.



The structures of simple molecules, according to Bohr's Memorandum. The earliest useful explanation of chemical bonding by electron exchange was probably that of Thomson. Figure 7

► The ground-state configuration is the one with the lowest possible potential energy consistent with the principle of angular momentum.

Alas! these directions do not suffice, for they point to structures—such as a single-ring lithium atom—in obvious disagreement with atomic volumes and chemical data. So Bohr assigned distributions more by intuition than by principle, with the curious result given in Table 2. Note particularly the confluence of inner rings at neon ($Z = 10$) and argon ($Z = 18$), brought about, Bohr thought, by the demands of the usual laws of mechanics. Bohr's care and trouble in constructing Table 2 may be indicated by the alternative distributions of figure 8.

Part II of Bohr's paper of 1913 also resolves—or rather shelves—the problem of radioactivity by tucking it into the nucleus. As for Part III, it argues the merits of Bohr's hydrogen molecules.

Thomson's response

Thomson did not salute Bohr's work as the capstone of his own. To him, setting down an arbitrary condition like $T = K\nu'$, pretending that it had dynamical signifi-

cance, was not doing physics; it was a screen of ignorance, a cowardly substitute for "a knowledge of the structure of the atom." Nothing could be easier, or so Thomson told the British Association for the Advancement of Science in September 1913, than to obtain quantum theoretical results in an orthodox mechanical manner.

Take Einstein's formula for the photoelectric effect, $mv^2/2 = h\nu$, for example. (For simplicity Thomson omitted the work function.) Assume, he said, that the usual Coulomb attraction A/r^2 operates only in a few separated, pie-shaped regions in the atom and that, in addition, an inverse-cube repulsion B/r^3 exists everywhere. An electron can sit in stable equilibrium within the pie-shaped regions at a distance a from the atom's center, where $a = B/A$. The frequency of small vibrations about this position is

$$\nu' = \frac{1}{2\pi} \frac{B}{a^2} \left(\frac{1}{mB} \right)^{1/2}$$

Assume that a passing light wave of frequency $\nu = \nu'$ strikes the electron, and gives it enough energy to cross from the pie-shaped region into one of uncom-

Table 2. Bohr's electronic distribution, 1913

1(1)	7(4,3)	13(8,2,3)	19(8,8,2,1)
2(2)	8(4,2,2)	14(8,2,4)	20(8,8,2,2)
3(2,1)	9(4,4,1)	15(8,4,3)	21(8,8,2,3)
4(2,2)	10(8,2)	16(8,4,2,2)	22(8,8,2,4)
5(2,3)	11(8,2,1)	17(8,4,4,1)	23(8,8,4,3)
6(2,4)	12(8,2,2)	18(8,8,2)	24(8,8,4,2,2)

From Phil. Mag. 26, 476 (1913). The symbol $N(n_1, n_2, \dots)$ indicates the total number of electrons and their distribution counting outward from the nucleus.

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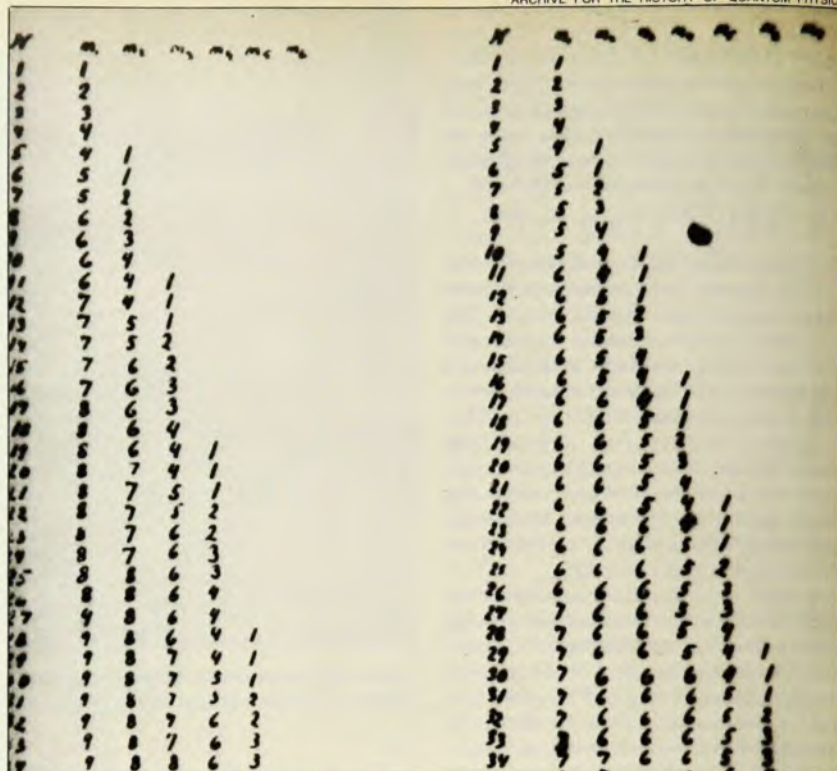
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Tentative electron distributions, 1912–13. This is a part of Bohr's manuscript with two sets of ring populations (the n 's) for atoms with electron number N up to 40. Figure 8

compensated repulsion. It will be pushed out into the world with kinetic energy

$$\frac{1}{2}mv^2 = \int_a^{\infty} \frac{Bdr}{r^3} = \frac{B}{2a^2} \\ = \pi(mB)^{1/2}v' = \pi(mB)^{1/2}v$$

Now set $\pi(mB)^{1/2} = h$: Einstein's formula emerges, and h discloses its true nature, a shorthand for the product of certain electronic parameters.

This *tour de force* was widely applauded by Thomson's school. *Nature* called it a "brilliant attempt" not soon to be forgotten. Other sympathizers rushed to reinterpret Bohr's fundamental contribution, the elucidation of the Balmer lines. One likened the plum pudding to a rotating, pulsating sphere of gas, and imagined that the Balmer lines were emitted by electrons running around on nodal surfaces. Another made what he called a "spherical counterpart" to Thomson's sectioned atom, a baroque structure with many niches of stable equilibrium about which an electron could vibrate at one or another of the Balmer frequencies.

Thomson himself contributed to this curious literature. "If [the Bohr theory] is true," he said, "it must be the result of forces whose existence has not been demonstrated." He set out to find these forces, and to represent them in "the working of a model"; and so, for a time, he occupied himself in reinterpreting Bohr — as Bohr had been reinterpreting him.

He ended by appealing to a force varying sinusoidally with the distance between the radiating electron and what he coyly called the "positive center" of the atom.

These rearguard actions did nothing to divert the progress of the quantum theory of the atom. When academic physics resumed after World War I, Thomson recognized that he was out of date and resigned the Cavendish professorship in favor of Rutherford. Not that he gave up physics; but he could never be persuaded that quantum theory was a fundamental one.

In his *Recollections and Reflections*, an autobiography published in 1937, Thomson allowed that Bohr's papers had "changed chaos into order" in certain branches of spectroscopy. And that, he thought, was "the most valuable contribution which the quantum theory has ever made to physical science."

Further reading

- For Thomson: *Dictionary of Scientific Biography*, XIII, Scribners, New York (1976), page 362; "The Scattering of α and β Particles and Rutherford's Atom," *Archive for History of Exact Science* 4, 247 (1968).
- For Bohr: J. L. Heilbron, T. S. Kuhn, "The Genesis of the Bohr Atom," *Historical Studies in the Physical Sciences* 1, 211 (1969).
- For the Archive for History of Quantum Physics: T. S. Kuhn, J. L. Heilbron, P. Forman, L. Allen, *Sources for History of Quantum Physics*, American Philosophical Society, Philadelphia (1967). □