CHAPTER NINE
ENTER NIELS BOHR

“When the clock strikes 13, not only do you doubt it, but it also casts doubt on the previous 12 rings.”
Mark Twain

"...a small research of his own."

By the year 1896 Thomson’s Nemesis had already arrived at Cambridge, in the person of a young graduate student from New Zealand named Ernest Rutherford. After completing his doctoral studies with Thomson, Rutherford moved first to McGill University in Montreal, and then back to England, into a new laboratory at The University Of Manchester. Rutherford's personality contrasted radically with that of Thomson. The latter was introverted and egotistical, whereas the former was a generous, warm, gregarious man, even given to bursts of song at moments when he had made one of his many important discoveries. His personal magnetism drew talented people to his laboratory, and made him a pioneer in the employment of research teams. Some of his students even went on to win the Nobel Prize on their own.

It was on a fateful day in the year 1909 that his assistant Hans Geiger appeared in Rutherford’s office, asking if it would be proper for “young Marsden” to be permitted to start “a small research of his own”. This “small research” consisted of bombarding a piece of thin gold foil with alpha particles\(^1\), in order to see if any of them would per chance be deflected through a large angle. The alpha particles, in passing through the foil, were expected to suffer little if any deflection, and to impinge unmolested upon a screen. The latter was coated with zinc sulfide, a substance which luminesces briefly when impacted by particles. The location of the minute flash of light on the screen would then be the measure of the size of the deflection of the alpha particle, which carries a positive charge. Later, in his reminiscences of the event, Rutherford confessed that he had not really expected any unusual result from the experiment: “It was quite the most incredible event that has ever happened to me in my life....” But in retrospect the whole thing was uncanny. What Marsden saw (after his eyes had become sufficiently dark-adapted) was that a few of the alpha particles were indeed being deflected through large angles. But to produce a large angle of deflection requires the application of a correspondingly large force. The latter can only be the result of a positive charge being tightly concentrated in a massive object, localized in a very small region of space. Patiently, Rutherford sketched out the consequences of the presence of a heavy nucleus lying at the heart of the atom, and by 1911 he published his result, showing that the observed distribution of alpha particles on the screen was in perfect agreement with the idea that, at the heart of an atom, lies a tiny, massive nucleus, around which orbit the negatively-charged electrons. Thus, on that day in May 1909 the “nuclear atom” was born, and “Young Marsden” got his unexpected result.

Louis Pasteur once remarked: “Fortune favors the prepared mind”. It was certainly true on that occasion.
Rutherford’s model had a disarming familiarity about it; it was, after all, a miniature model of the solar system: one in which orbiting planets were replaced by orbiting electrons, and the gravitational forces by electrical ones. Thus Rutherford was using the solar system as a metaphor for the atom. But, just as in the case of Thomson’s model, there was a serious, possibly even a fatal, difficulty. This lay in the fact that orbiting electrons are continually deflected by the attractive force field of the nucleus. This deflection implies that the particles, in following their curved paths, are constantly being accelerated. From electromagnetic theory it can be shown that an accelerated charge is constantly losing energy; a fact which leads to the conclusion that the orbiting electrons should have promptly spiraled inward, crashing into their parent nucleus long before any of us were born; and consequently bringing the universe to a premature end.2 Thus both Thomson’s and Rutherford’s models were actually impossible under the accepted laws of physics—as they were understood at that time. And therefore, something was going to have to give way.

Bohr Goes To Work For Rutherford.

At this point, almost as if on cue, there walked onstage one of the most remarkable thinkers of all time: a young Dane named Niels Bohr3. In the fall of the year 1911 he had come to Cambridge, then as now a Mecca for scholars, for the purpose of doing post-doctoral work under the great J. J. Thomson. But weeks had passed, and Thomson had still not gotten around to reading Bohr’s dissertation. Perhaps there was a small matter: the fact that at the first meeting between them, Bohr had tactlessly pointed out the existence of certain errors he had found in Thomson’s work. In any case, it is known that Thomson was something of a loner, a man of reserve, whereas the very soul of Bohr’s style was conversation. He was given to conducting arguments that would last until the small hours of the morning, only to be resumed later in the following day. In short, Bohr learned by conversing with people; and, unable to receive stimulation, he felt himself withering on the vine at Cambridge. But shortly after Christmas, Rutherford came to Cambridge as the invited speaker at the annual Cavendish Dinner. His topic, of course, was his recent discovery of the nuclear atom. And also of course, Bohr was present in the great dining hall that evening to hear Rutherford speak. Totally captivated, he lost little time moving his base of operations to Rutherford's laboratory at Manchester, and so altered the course of human history.

The Key Assumptions

To start with, it was Bohr who showed that Rutherford's model was closer to the truth than was Thomson's. Exactly how this came about is one of the great tales in the history of science. During the years 1912-13, whether in Manchester or at home in Copenhagen, Bohr remained continually in close intellectual contact with Rutherford and his co-workers, sketching in the rough outlines of the physics of the atom, intuitively anticipating discoveries that would not be made until years afterward. But if we want to come to an understanding of the behavior of complex nuclei as well as of molecules, it’s necessary to begin by addressing the problem of why the “orbiting” electrons continue to retain their energy, even while being accelerated: this is the problem of the stability of the Universe—a consummation devoutly to be wished.

Bohr’s extraordinary empirical approach to the problem can be summarized in the following way:
• The “orbiting” electron, obedient to the laws of classical physics, radiates away its energy—except when it doesn’t, (a radical idea, that one). An electron that does not radiate energy is said to be in a “stationary state”.
• When the “orbiting” electron moves from one stationary state to another, that’s when the radiation of energy happens; and that energy is radiated in quanta— in "lumps”—following Planck’s Law!

"As soon as I saw Balmer's formula, it was all clear to me."

But how does the quantum enter the picture? After all, there has to be some kind of rationale for bringing it onstage. It was at this point in his career Bohr had the good fortune to encounter his old friend, H. M. Hansen. Hansen, a fellow Dane who had become an expert in the science of spectroscopy, had one very important piece of information to share with him: it was about spectra. When the light of the sun is passed through a prism we get the familiar solar spectrum: a smear of colored light, its visible portion spanning the range of colors from red to violet. But if instead we place a blob of table salt upon the end of a long wire and heat it in a gas flame, the flame suddenly takes on a deep yellow-orange hue. When its light is passed through a prism there is little contribution to the spectrum save from the color yellow, a result of the presence in table salt of the element sodium. By now the significance of this phenomenon was clear: by Planck’s formula, only one color means: only one energy. What we expect to see when we study a spectrum of an element is a display of various energies. But if we wish to understand the spectra of elements, it is wise to start first with the simplest example. And of all the elements in the Universe the one having the simplest, most fascinating spectrum is hydrogen. See figure 9-1 (below).

![Figure 9-1](image)

When viewed through a spectroscope, hydrogen displays a set of sharply delineated, thin bands of colors, principally: red, blue-green and violet. When the frequencies associated with these colors are measured, they follow a certain cryptic but suggestively orderly sequence: a kind of Rosetta Stone for scientists--a tantalizing and cryptic message that cries aloud for somebody to decipher it. And it was from Hansen in 1913 that Bohr learned of the first key to that cipher.

It had come in 1885 from the work of a Swiss high school teacher named Johann Balmer, whose
discovery can be described in the following fashion. The frequencies associated with the bands possessed certain numerical regularities: They are proportional to the following numbers: $1/2^2 - 1/3^2$, $1/2^2 - 1/4^2$, $1/2^2 - 1/5^2$, and so on. Significantly, Bohr had already numbered the stationary states in his model of hydrogen: 1, 2, 3, etc. Later Bohr was to say: “As soon as I saw Balmer’s formula, it was all clear to me.” Bohr instantly saw that the atom’s emission of light could be seen as the result of the energy change between any two stable orbits, as the electron jumped from one to another.

To understand this, one needs to observe that Planck’s quantum constant has the characteristics of a certain physical property called angular momentum; hence, said Bohr, what harm could there be in starting out by assuming that it is the angular momentum of the orbital electron that is quantized? Next, Bohr wrote out the expression for the energy of the orbital electron in the electric field of the hydrogen nucleus. Finally, he wrote the expression for the centripetal force on the electron. Extraneous variables could be eliminated, and there it was: a mathematical expression for the frequencies of the hydrogen spectrum, (see Appendix B).

Out in front of this expression stood a complicated constant, involving the electron mass, its charge, Planck’s constant and the speed of light. When Bohr had calculated its value: it came to:

$$R = 109,500 \text{ cm}^{-1}.$$ 

In the years after Balmer’s original discovery, the Swedish spectroscopist J. R. Rydberg had built on the work of the former,

![Figure 9-2](image)

**Figure 9-2** Bohr radii and allowed energies of hydrogen. The first few lines of several of the spectral series of hydrogen are shown at the left.

and had developed an empirical equation to account for the frequencies of the lines in the
hydrogen spectrum. Bohr's value, obtained from his daring quantum hypothesis, differed from the one measured by Rydberg by less than 2 parts in a thousand! This meant that henceforth any description of the behavior of the atom was going to have to be a quantum description. And however outrageous his assumptions had been, Bohr had placed his key in the lock of nature, and the key had turned!

A useful guide to Bohr's model is seen in the diagram of fig. 9-2. On the left we can see the quantum jumps that produce the spectral lines visible in figure 9-1. On the right side of the diagram are represented the corresponding orbits—the ones between which the jumps occur. The orbits and their energies are fixed by the quantization of the angular momentum. On the left side are the corresponding energies.

**Discontents And Inadequacies**

It must be admitted that Bohr was conscious of the inadequacies of his model. To talk about it, one was required to visualize a kind of circular amphitheater, with electrons rolling about the center, as if along the benches. The electrons could only receive energy by being struck like billiard balls, by other electrons or by electromagnetic waves. The metaphor of electrons visualized as tiny balls of charged matter, engaged in jumping from one classical orbit to another, seemed entirely out of character with the outlandish, madcap idea of the quantum. Frankly, the whole theory had a disturbing jerry-built quality to it; leaving one feeling unsatisfied—especially if one had been trained in the methods of classical physics, with its emphasis upon mechanical models. Moreover, the electron was supposed to be an actual particle of matter; and the approved model for matter in motion, the classical paradigm, was the smooth, continuous trajectory of a projectile, as visualized earlier by Galileo. Bohr knew that sooner or later he would be called to account for these awkward gaps in his theory. Sure enough, his first taste of what was to come occurred at his first meeting with Einstein in 1920. At that time the latter let it be known that he objected to any theory that “left to chance the time and direction of the elementary processes.” What Einstein meant by this was that for Bohr's theory to gain his approval, determinism would have to prevail in quantum processes. Further sources of difficulty were the methods—those that had worked so well in the case of hydrogen; they were ineffective when applied to the other, more complex elements in the Periodic Table. Bohr knew that he was going to have to revolutionize his thinking. But the job of sketching in the broad outlines of atomic structure for others to complete, and using them to explain the significance of the Periodic Table, had kept him busy for several years, while the rest of the European continent was convulsed by war.

**A French Prince Changes His Studies From Medieval History To Physics.**

While The Great War was turning Europe into a vast slaughterhouse, Louis de Broglie, a descendant of a noble French family, his studies of medieval history having been interrupted by his draft board, found himself thrust into the signal corps, and sent off to study radio transmission. In this manner he acquired a decided taste for science, and after the war he entered into graduate study in physics at The University Of Paris. For his dissertation de Broglie chose a very interesting topic, one that was inspired by a deep mystery associated with the Photoelectric Effect.
You will recall that James Clerk Maxwell and Heinrich Hertz had demonstrated that light is propagated in the form of an electromagnetic wave. Therefore, when we think of light striking a metal surface, we can conjure up an image of ocean waves crashing down upon a beach. We might also be tempted to stretch this analogy further, and to conclude that the amount of energy in the wave is proportional to the intensity of the light—roughly analogous to the square of the height of the ocean wave. But such is not the case. Instead, Einstein found that the energy of the ejected electrons is proportional only to the color, (the frequency), of the incident light.

We can turn down the light’s intensity until it is nearly zero, but the ejected electrons will still possess the same energy as before. There was no ducking it; the effectiveness of the light, in giving energy to electrons, depended only upon the color.\(^9\)

A really crucial observation was made when someone measured the total amount of energy per second arriving at the surface of the metal, and divided this quantity by the number of surface atoms. It was found that if the light energy striking the metal were to arrive in the form of a wave, and if it were uniformly distributed over the surface of the metal, it would take a fairly long time before enough energy would be absorbed to dislodge even one single electron. Amazingly, however, in the laboratory the electrons were emitted immediately, just as soon as the light was switched on. The only way to explain this phenomenon was to say that the quantum of light energy was arriving at the metal in the form of particles!

The game became even more interesting in the year 1922, when a young American named Arthur Compton caused X-rays to impinge upon electrons. Compton measured the energy of the X-rays and the electrons, both before and after the collision. He found to his surprise that the phenomenon could only be explained if he assumed that the X-ray photons also struck the electrons in the form of particles. There was no question about it; light seemed to travel as a wave, but when it arrived, it arrived in the form of a particle, confined to a very small region of space. Therefore, to represent our experiences in a faithful way, we must treat light as if it has a dual nature. Sometimes we can use the wave metaphor; and at other times we must use the particle metaphor. It was this last notion that started de Broglie on a long, fateful train of thought.

The Princely musings went on as follows. This mysterious Being Of Light, the photon, could manifest itself under one line of questioning in the form of a wave, and under another in the form of a particle. The electron on the other hand, had hitherto manifested itself only as a particle. Suppose however, that the electron, too, had a wavelike nature, one that had been hitherto undetected; if so, voilà! This would have fascinating consequences for the atom. For if we were to confine an electron to one of the Bohr orbits, it would be similar to confining a wave to a string whose length was the circumference of the orbit! And as the orbit gets larger, the number of wavelengths it can contain will also increase—in jumps, by one
quantum at a time. What De Broglie envisioned was an arrangement that looked like the one in
we have in figure 9-3. (above).

By comparing the Bohr quantum condition with the stipulation that the orbital circumference can
accomodate only integral numbers of waves, de Broglie discovered that the wavelength
associated with a particle is inversely related to its momentum: that is, as the momentum
increases, the wavelength gets smaller. The most important part of this result is: the more
accurately we know the wavelength of a quantum object, the more accurately we will be able to
determine its momentum: the one being an inverse measure of the other. De Broglie himself
went on to visualize the electron as being related to its wave in a manner analogous to the way an
ocean liner is related to the tugboats that push it about in the harbor. Therefore, the wave, for de
Broglie, was a kind of “pilot wave”, while the electron itself remained, for him, a classical,
 discrete, Newtonian object. In his heart, De Broglie was to remain a Newtonian to the very end
of his long life.

In November 1924, when he submitted his dissertation, the reaction of de Broglie’s committee
was a blend of stony Cartesian skepticism and blank, total non-comprehension. It had all been
too bizarre. Paul Langevin, the committee chair, finally decided that the matter should be
submitted to Albert Einstein, a notoriously eccentric genius, for his opinion. Einstein took a
liking to de Broglie’s theory immediately; for he saw in it a way out of the dilemma he had
helped to create in 1905 with his unorthodox interpretation of the photoelectric effect. Perhaps de
Broglie could light the way back from this abominable discontinuity of energy levels and all this
unpredictable quantum jumping, back to the One True Faith: that of continuity and
determinism—the Newtonian Myth. When one of the members of de Broglie’s dissertation
committee had asked him how one would go about determining the wave nature of the electron,
the Prince replied tersely: “By focusing a beam of electrons onto a crystal of metal”. Curiously
enough, during the following year two researchers at the Bell Telephone Laboratory in the
United States, Davison and Germer, men who apparently had never heard of Louis de Broglie, happened by coincidence to be working with an electron beam; and when they focussed the beam upon a mass of powdered nickel, they found themselves staring at an interference pattern, that tell-tale badge of wave motion. There it was! The electron really behaved like a wave!

An Austrian Professor Has A Romantic Adventure And Discovers The Correct Equation.

After reading de Broglie's dissertation, Einstein relayed the news to Erwin Schroedinger, a 38-year-old Viennese who held a professorship at Zurich. Surprisingly, Einstein was never really fluent in mathematics, (if you are a genius, it's not necessary). But he knew that Schroedinger, a mathematical virtuoso, would be just the man for the job—the job of creating a mathematical description of de Broglie's waves. Schroedinger realized that he had just been given the ingredients for describing the motion of those elusive “matter waves”, as the electrons were then called. When, in the winter of 1926 a colleague extended to him the use of a chalet in the Swiss Alps, he accepted the offer with some alacrity, for the Herr Professor also had a secret romance on his agenda, and that secluded chalet would serve his purposes to perfection. After some time had elapsed, a somewhat disheveled Schroedinger emerged from the Chalet, carrying in his hand the famous equation that bears his name. In the course of his voluptuous holiday, he had found a way to show how de Broglie’s “waves of matter” could move in a strictly deterministic fashion, under the influence of various forces.

This result was extremely pleasing to Einstein, a believer in Continuity and Determinism, a man who had found quantum jumps to be very unsettling. Einstein had desperately wanted the electron to be something smooth and continuous, just like a wave, and Schroedinger's result seemed almost too good to be true. But actually, there were implications concealed in Schroedinger's work that would continue to haunt Einstein for the rest of his life.

Two Theories: Too Much Of A Good Thing?

Word of Schroedinger’s accomplishment was swiftly transmitted to Niels Bohr at his institute in Copenhagen, where the news of Schroedinger's theory had an electrifying effect upon everyone, for Bohr had already been talking with a young German named Werner Heisenberg, who had also been working on the problem of the quantum. Heisenberg had produced an esoteric theory, which predicted the results of quantum measurements by means of manipulating certain creatures from the world of abstract algebra—devices called “matrices”. Before the work of Heisenberg there had been no coherent theory of the quantum; but now there were two theories—ones which did not seem to resemble each other in the slightest. What could one make of this?

As the year 1926 wore on, new developments were occurring on an almost weekly basis. Soon Erwin Schroedinger was able to demonstrate that his version of “quantum mechanics” was really equivalent to Heisenberg’s. Heisenberg, working in Goettingen "at his end of the tunnel," with the help of his mentor, Max Born, showed that his matrices could be thought of as representing (standing for) various measurement processes, such as the measurement of the position or momentum of a particle. It was at this point that the famous Heisenberg Uncertainty Principle was discovered. This Principle says the following: if we wish to observe, say, the location of an
electron, we have to shine light upon it. A low energy photon will not disturb the electron much, but it will also not permit us to see the electron very clearly. A high-energy photon will enable us see the position of the electron in sharper resolution, but it will also impart an indeterminate quantity of momentum in colliding with the electron. Thus, according to Heisenberg’s original formulation of the Principle, a measurement of the one quantity precludes the receipt of information about the other, but as he saw it, this was simply a result of the disturbance of the electron’s motion. The new discovery, however brilliant it was, turned out to be insufficient; it was only the tip of a conceptual iceberg. For when Heisenberg paid a visit to Copenhagen, Niels Bohr was waiting for him with news of the utmost importance.

What Bohr had to say was this:
• It is the act of measurement of a property of a quantum object that actually bestows that property upon the object; there is no way to separate the property itself from the measurement of that property! This idea sounds the death-knell for radical atomism.
• No measurement, no property. How can we talk meaningfully about a property that we never measure? Position and momentum are intrinsically incompatible properties, because they are obtained by methods of measurement, ones that are physically incompatible with each other.
• You can’t obtain a measurement of position, once you have at the same time, measured the momentum, and visa-versa. This result has become known as:

The Principle Of Complementarity.
• Position and momentum are “complementary” quantities. Such quantities are mutually incompatible, but both are essential, for neither one nor the other, taken alone, gives a complete description of the state of a quantum system—thereby giving a description of the world.
• Therefore it makes no sense to ascribe a position or a momentum to a quantum object if an appropriate measurement has not taken place.

To say that Heisenberg found this idea upsetting is an understatement; he was actually in tears. In fact, he and Bohr nearly had a falling-out over it. In later years Heisenberg described his feelings at the time:

I remember discussions with Bohr which went through many hours till very late at night and ended almost in despair, and when at the end of the discussion I went alone for a walk in the neighboring park I repeated to myself again and again the question: “Can nature possibly be as absurd as it seemed to us in these atomic experiments?”

Later discoveries revealed that there was no way to construct a Schroedinger wave that could be said to represent directly a particle such as an electron, for instance. It had long been known that an appropriate superposition (a “bunching up”) of Schroedinger waves could indeed be constructed, something that looked like a sharp spike in space. But it became disappointingly clear to Einstein and the wave enthusiasts that this spike would quickly dissipate, spreading out through all space. However, a radical suggestion was to come from Professor Max Born from his laboratory at the University of Goettingen. Just what this suggestion was, is something that we shall learn presently.

It was now the year 1927. Fourteen years after his historic paper on the hydrogen atom, Bohr’s
quantum philosophy had finally jelled, and none too soon, either. In October of that year there would be a time of testing; for then would come the first great confrontation with Einstein, leader of the Continuist, Determinist Faction--The Old Guard. The year 1927 was really the year of the Storm—a Storm that eventually washed away the old intellectual order.

And by the end of the year 1927 the outlines of quantum physics were just beginning to be clear. In the words of physicist-philosopher Bernard d’Espagnat:

...whoever tries to form an idea of the world, and of man’s position in the world,—has to take the findings of quantum physics seriously into account.

Notes:
(1) An alpha particle is the charged nucleus of a helium atom. These occur in the course of the radioactive decay of large atoms. It just turned out that Rutherford was the world authority on the subject of radioactive decay. Thus he was ideally placed as an actor in the play.
(2) If we were to put a hydrogen nucleus together with its orbital electron at 9 AM, according to the classical theory the entire structure would last for a mere $10^{-9}$ seconds, by which time the electron would have spiraled irrevocably into the nucleus. The world would never make it to 9:01 AM on the day of creation; nor would there ever be anything left to record its demise. That this doesn’t happen is well, self-evident.
(4) These thin bands are really called “lines”. They occur because the light from the glowing gas is passed through a narrow slit. The “lines” are really images of the slit, either captured on a photographic emulsion, or just displayed upon the wall of the laboratory.
(5) Actually Rydberg’s constant is not given in terms of the frequencies of the light, but rather in terms of something called “the inverse wavelength”, a quantity dear to spectroscopists.
(6) As the respected scientist Sir James Jeans had remarked, the only justification for Bohr’s assumptions was “the very weighty one of success”. But the phone call that came through from Stockholm was real enough; Bohr was duly awarded The Prize in 1922.
(7) There is even a Latin proverb: Natura non facit saltum: “nature makes no leaps”. Little did they know…. Now is a good time to look back at Galileo’s analysis of projectile motion, in Chapter Two.
(8) World War I, called the Great War, broke out on 3 August 1914, and lasted until an armistice was declared in 1918. World War II was essentially a continuation of World War I.
(9) Actually, the problem is somewhat more complicated than this. Both Einstein and Compton treated the problem as if the photons were arriving one at a time. In fact, at that time it was impossible to control the number of photons being produced, but no one was aware of it. 80 years would pass before a completely rigorous demonstration of the quantum nature of the photoelectric and Compton effects would be possible. Fortunately, no one realized this at that time. Strictly speaking, Einstein did not really demonstrate the existence of the photon, but his guess was still right.
(10) As Malory coyly remarked about the tryst between Lancelot and Guinevere, (in La Morte d’Arthur 14th cent.) “I know not what they did there, for then their customs were different from ours”. The woman who made such a distinguished contribution to science seems destined to remain forever anonymous.
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(11) It is this idea, central to the Copenhagen Interpretation, that led to Heisenberg’s terse dictum: “An atom is not a thing”.