The Atom

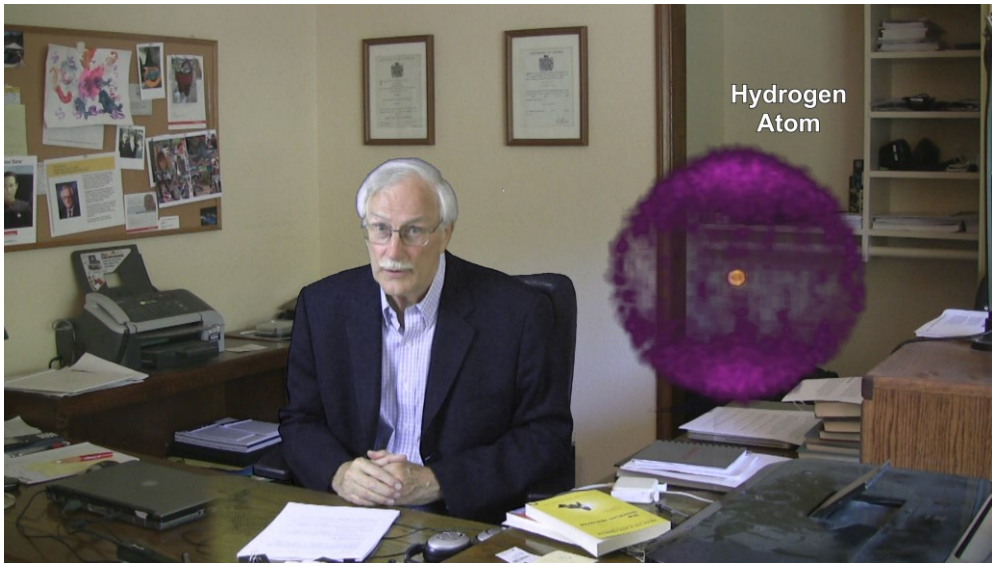
***{Abstract** – In this segment of our "How small is it" video book, we cover the atom.*

*We start with J.J. Thomson's Plum Pudding model of the atom. We then introduce Alpha, Beta and Gamma radiation, and show how Ernst Rutherford used Alpha Particles in his scattering experiments to develop his version of the atom. We then cover Niels Bohr's quantized atomic model along with input from Louis de Broglie who used the wave nature of electrons to show that they take the form of standing waves enveloping the nucleus instead of point particles orbiting the nucleus. We then introduce Schrodinger's Equation, the Heisenberg Uncertainty Principle and the Pauli Exclusion Principle together with electron spin to fill out what we know about electrons around atoms. We explain electron tunneling as an example of these concepts and introduce the Scanning Tunneling Microscope based on these concepts to see atoms by "feeling" them. In closing, we introduce the atomic nucleus and Chadwick's discovery of the neutron, review how small the nucleus is and ask how is possible for protons to hold together in the nucleus when there like charges should push them apart.**}*

Introduction

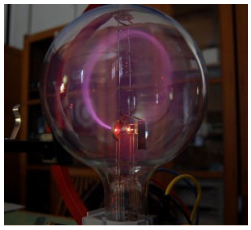
[Music: Joseph Haydn – "*Cello* Concerto No. 2 in D" **–]**

The ancient Greeks wondered how small things can get. One school of thought proposed that a substance such as water could be cut in half infinitely. Others thought that you could only take it to a point and then you would one 'atom' of water. If you split that, you wouldn't have water any more. They were right but they didn't have the tools to prove it. In our first segment, we used photons and electrons to see things down to the size of a carbon atom - 0.14 nanometers. That's small. Atoms are so small, that there are as many atoms in your DNA as there are stars in the Milky Way galaxy. In fact, there are more atoms in the breath of air I just took than there are stars in the visible Universe! The structure of the atom is responsible for nearly all the properties of matter that have shaped the world around us and within us. But what do we actually know about atoms and how small are the particles that combine to create an atom? That's what this segment is all about. We'll start with early guesses about atomic structure and show how we figured out how it actually works. It's a fascinating story and it will put us on the path to understanding elementary particles and the Higgs Boson.

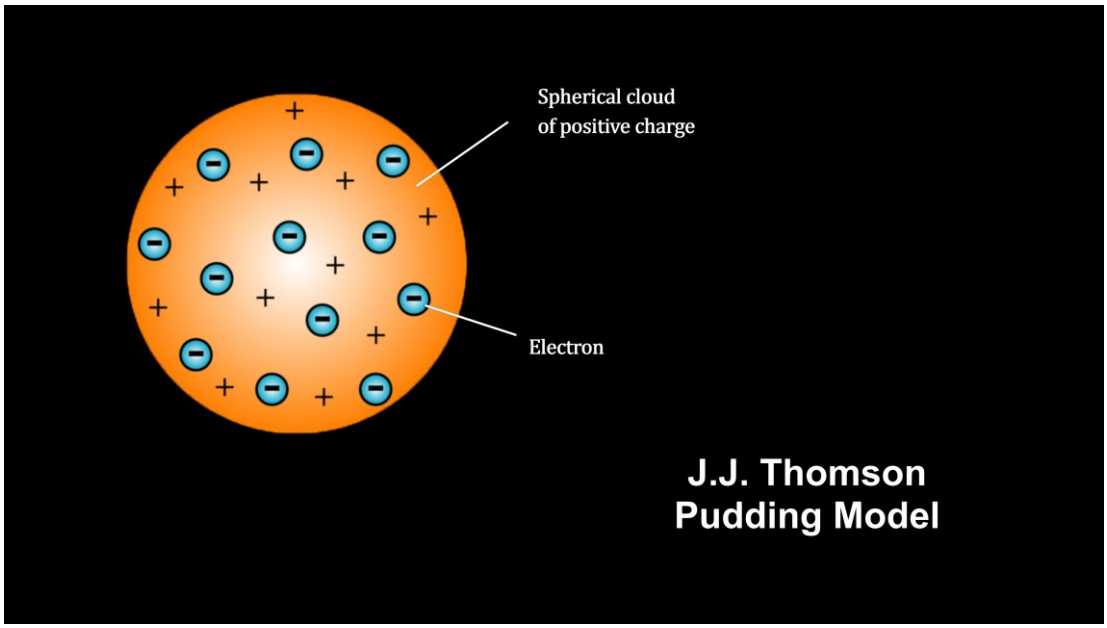


The Thomson Atom

In the 19th century, it was well understood that the chemistry of substances consisted of atoms. But we knew very little about atoms themselves. It was the discovery of the electron by JJ Thomson that first introduced the idea that an atom had parts.



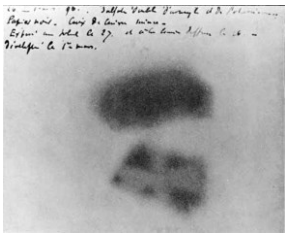
In 1898, with the electron being so light compared to the atom, Thomson suggested what is called the "plum-pudding" model of the atom – with a uniform mass of positively charged matter containing spots of electrons imbedded in it like plumbs in a pudding.



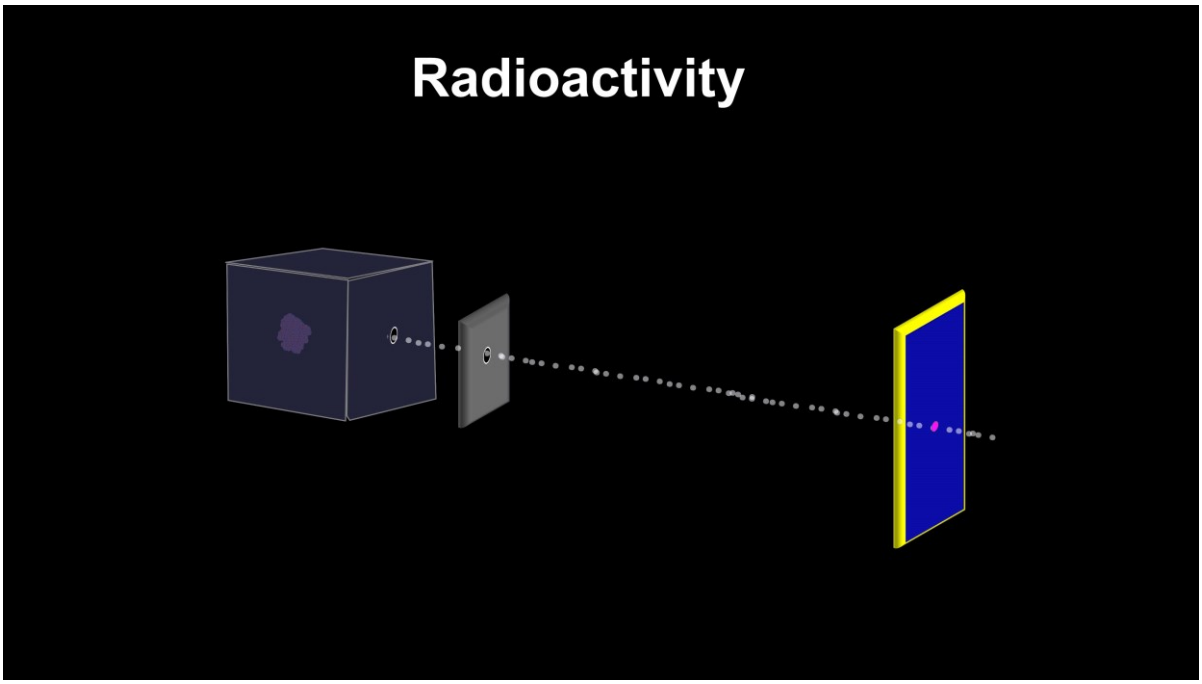
A way to find out if this model is correct or not, is to probe the pudding. But you need to probe with something smaller than the object being probed. For example, you can't probe a grain of sand with your finger. In 1898, there simply wasn't anything smaller than atoms that could be use to probe an atom.

Radioactivity

But around that time, radioactivity was discovered by the French scientist Henri Becquerel. Using uranium salts, he was able to blacken a photographic plate. Here's a photograph of the plate.

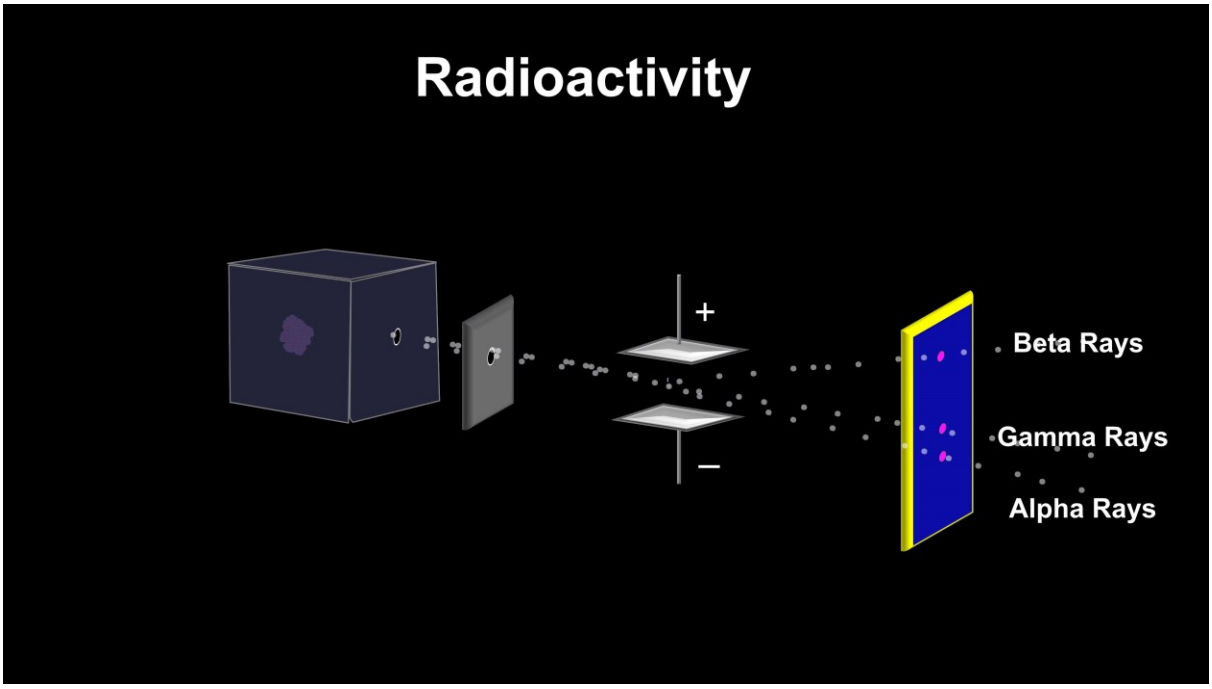


Further research by Becquerel, Ernest Rutherford, Madam Curie, and others discovered 3 types of radiation. Here's how they did it. A radiation source shines on a lead plate with a small hole in it to create a beam. The beam is directed at a florescent screen. The screen flashes when it is struck. Without any electric field present, the beam illuminates a single point on the screen.

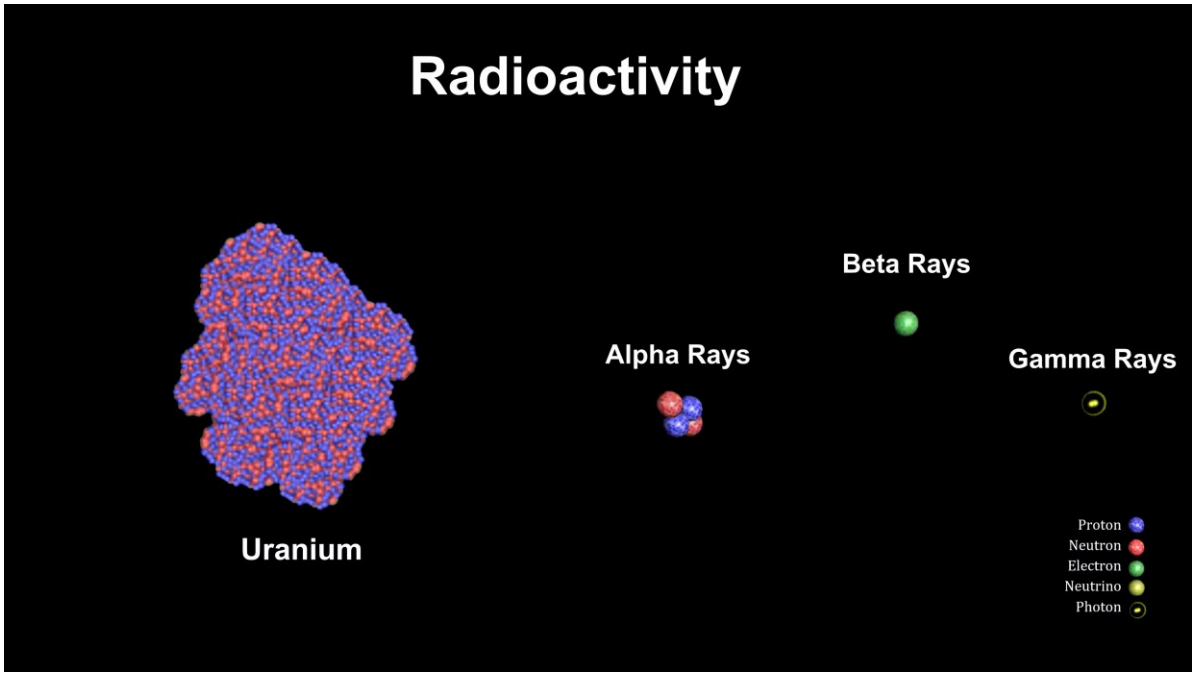


When an electric field is applied, the beam is separated into three components. One is deflected upward by the electric field indicating that it is negatively charged. These were named beta rays. One is deflected downward, indicating that it is positively charged. These were named Alpha rays. The radiation that continued to hit the center was not affected by the electric field and therefore has no charge. These emissions were named gamma rays. It was noted that the Alpha rays were deflected far less than the beta rays. This was because the alpha particles are more massive than the beta

particles. You'll recall the mass spectrometer we used to measure the mass of electrons in our previous segment.



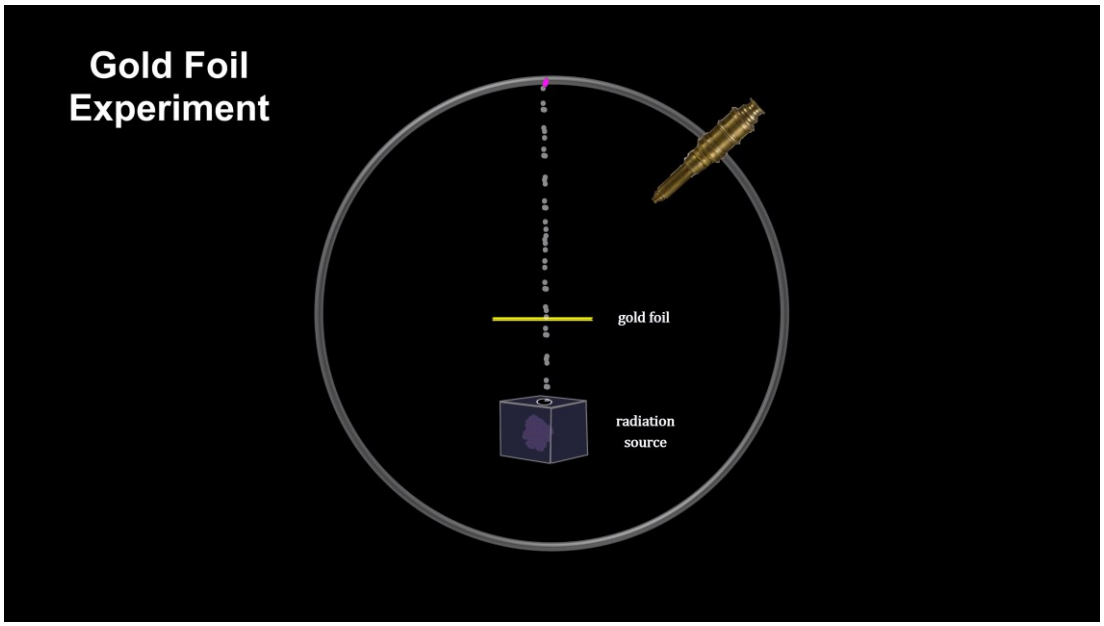
It turned out that beta rays are high speed electrons. The alpha particles were later found to be helium atoms without their electrons. The gamma rays turned out to be high energy photons, more energetic than x-rays.



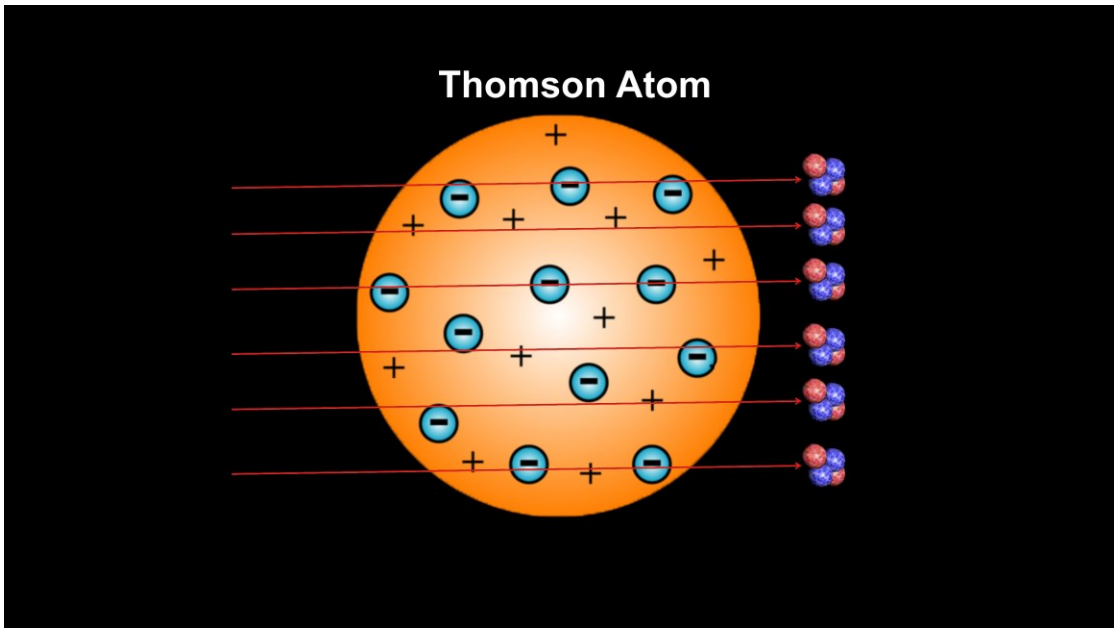
The Rutherford Atom

[Music: Mozart – "Violin Concerto No. 3" – **]**

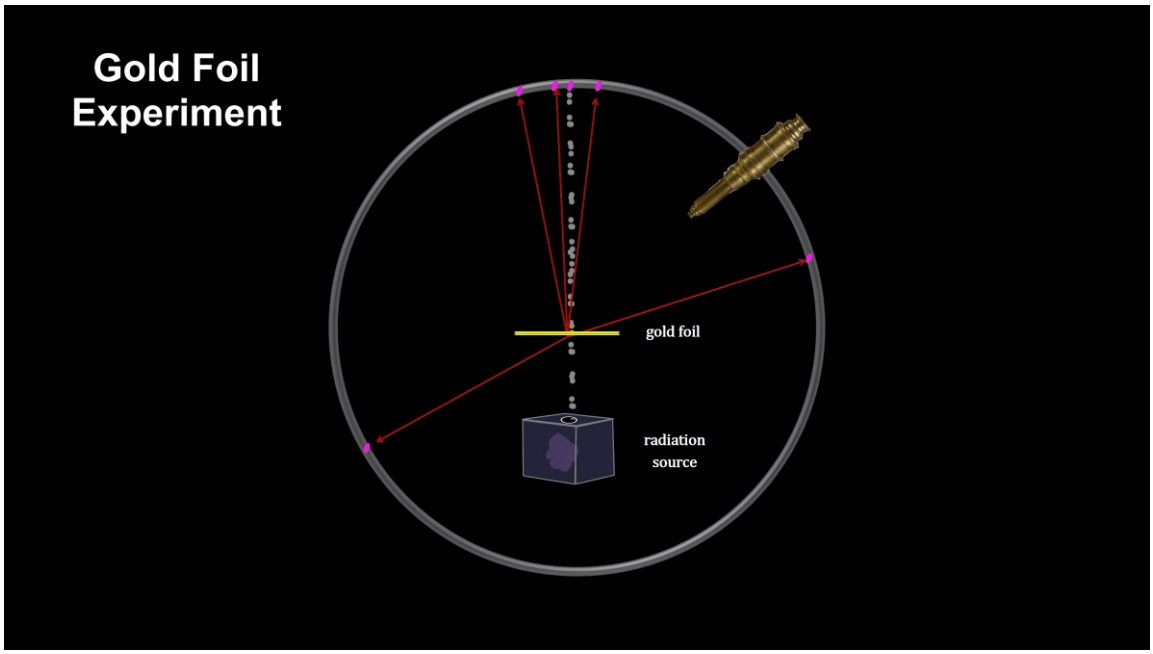
With Alpha particles, Rutherford had something to fire at atoms to see if they were indeed like a positive pudding with imbedded electrons. Here's a graphic of the apparatus used to run the experiment. An alpha particle emitting substance is placed behind a lead screen with a small hole in it to enable a narrow beam of particles to flow through. This beam is directed at a very thin gold foil. A movable zinc sulfide screen is placed on the other side of the foil. Zinc sulfide flashes when hit by an alpha particle. A microscope swivels to view all scattering angles.



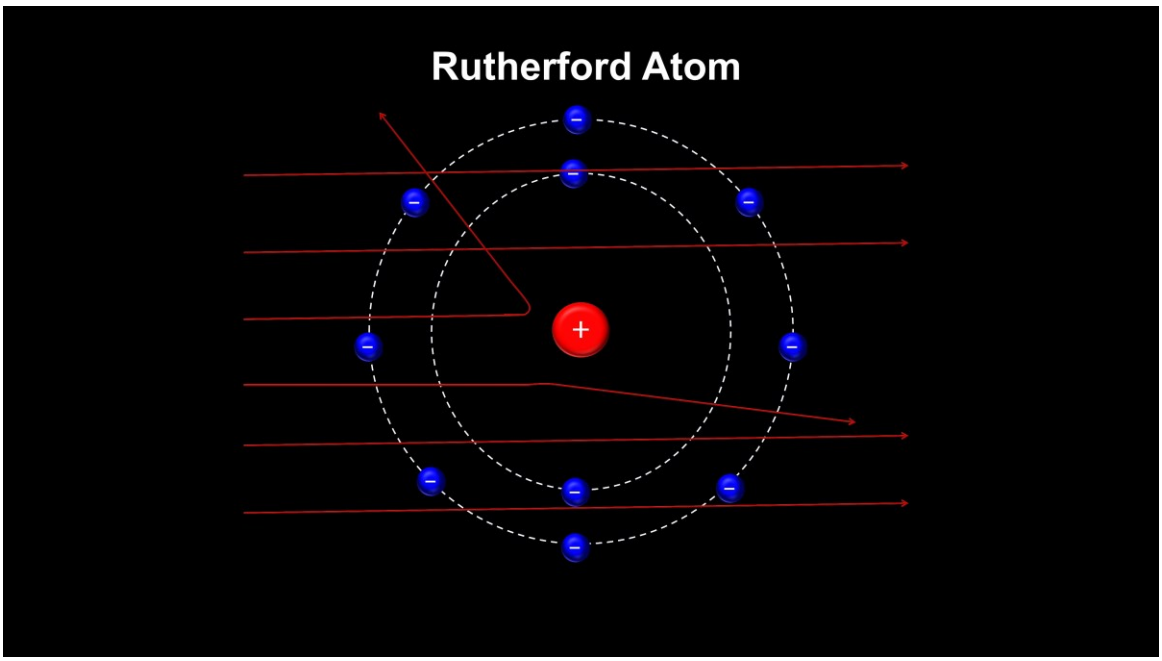
If the Thomson model was correct, the positively charged alpha particles would pass through the distributed and therefore diluted positive charge in the gold atoms with little or no deflections.



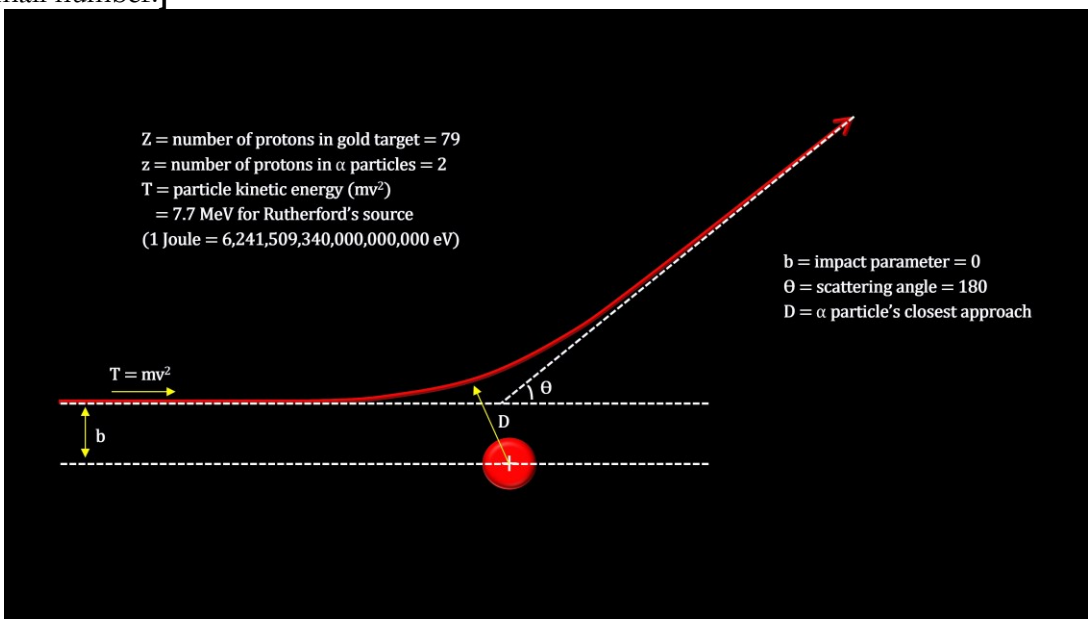
But after days of observations here's what they found. While most of the alpha particles do go right through with only minor deflections, some were scattered through very large angles. A few were even scattered in the backward direction! Rutherford described it as "almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back at you.'



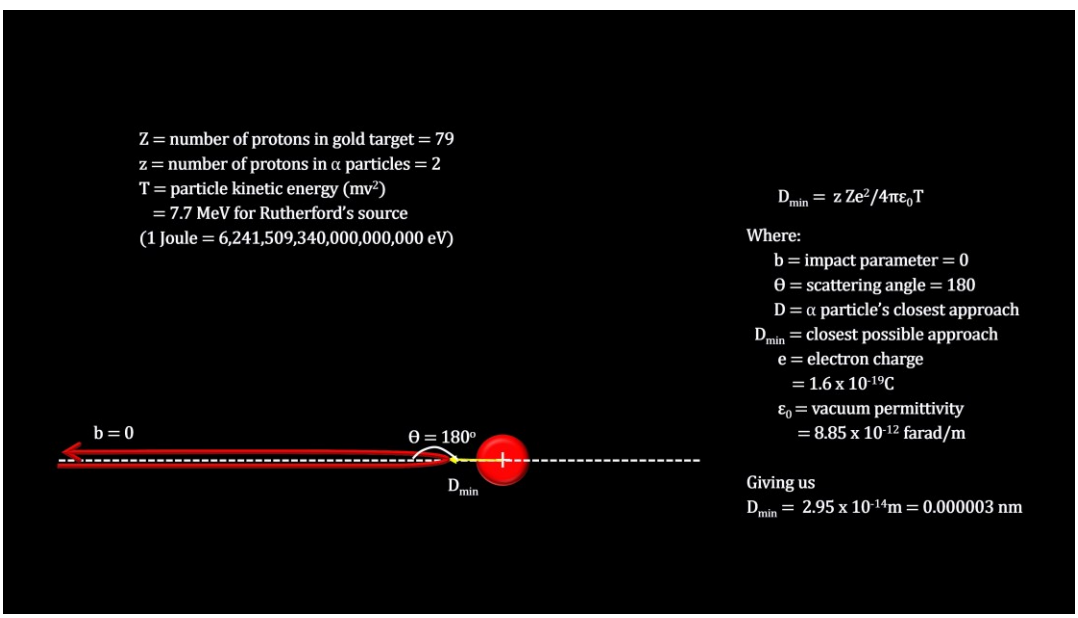
To explain these results, Rutherford was forced to picture an atom as being composed of a tiny nucleus where the positive charge and nearly all of its mass are concentrated, with electrons some distance away. Note that the closer the alpha particle is to the nucleus the greater the angle of the deflection. We can use this angle to measure the maximum possible size of the nucleus.



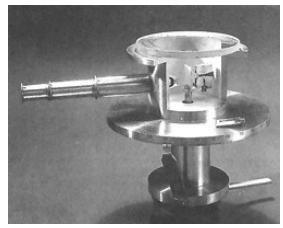
Here we have an alpha particle trajectory with an 'impact parameter' b that scatters at the angle theta and reaches a closest distance labeled D. The number of protons in the alpha particle is 2 and the number of protons in gold is 79. The energy of the naturally occurring alpha particles used by Rutherford where 7.7 million electron volts. An electron volt is the energy it takes to move one electron across one volt. **[**There are 6,240 trillion electron volts in one joule. So, you can see it is a very small number.**]**



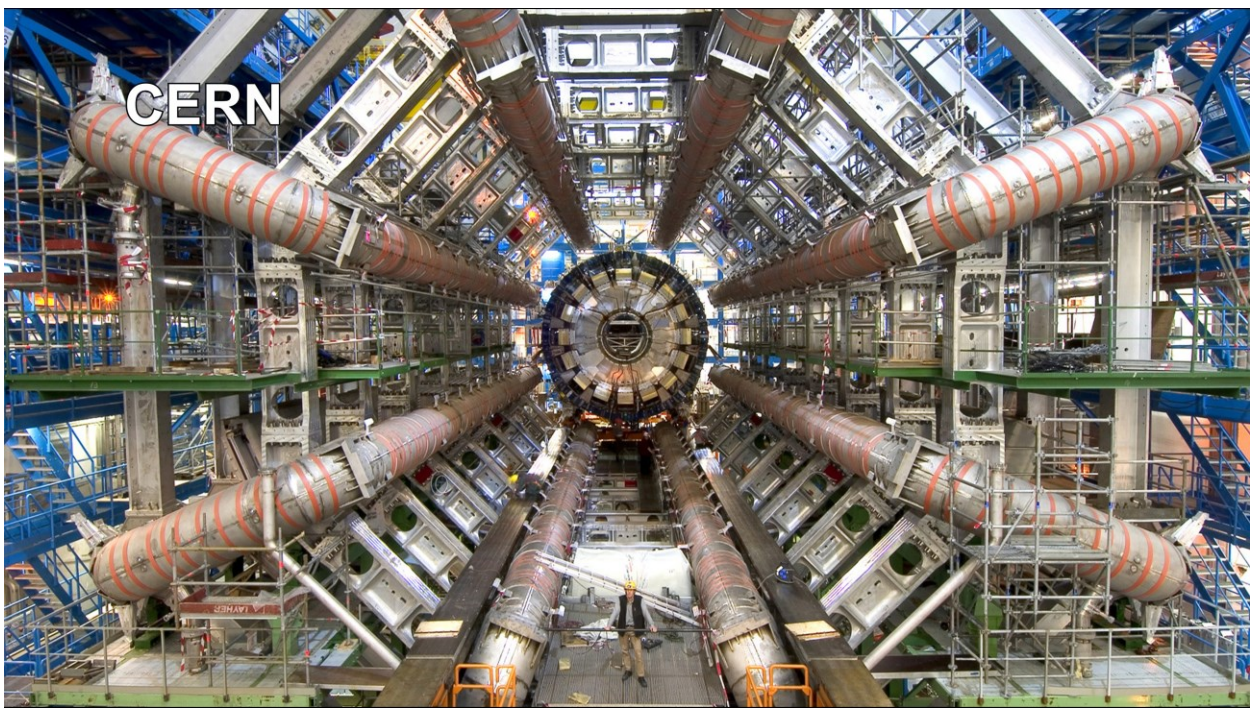
If we consider the direct hit trajectory, the initial kinetic energy of the alpha particle will drop to zero when it reaches its closest possible distance to the nucleus. All its kinetic energy would have been converted to electric field potential energy. Conservation of energy tells us that these two numbers must be equal. Rutherford's calculations showed that the radius of a gold atom nucleus cannot be any larger than 0.00003 nanometers. That was ten thousand times smaller than the size of a gold atom.



Here's a picture of the test apparatus Hans Geiger (of Geiger counter fame) and Ernest Marsden built to look for scattered alpha particles from every angle. The microscope could be swiveled all the way around the gold foil.



This is the first experiment that fired a beam of particles at a target to detect the scattering effects and deduce what is going on. That was around a hundred years ago. This is exactly what we are doing today at the European Center for Nuclear Research (CERN) to analyze the Higgs Boson. We'll return for a deeper look at Rutherford scattering when we get to particle accelerators.

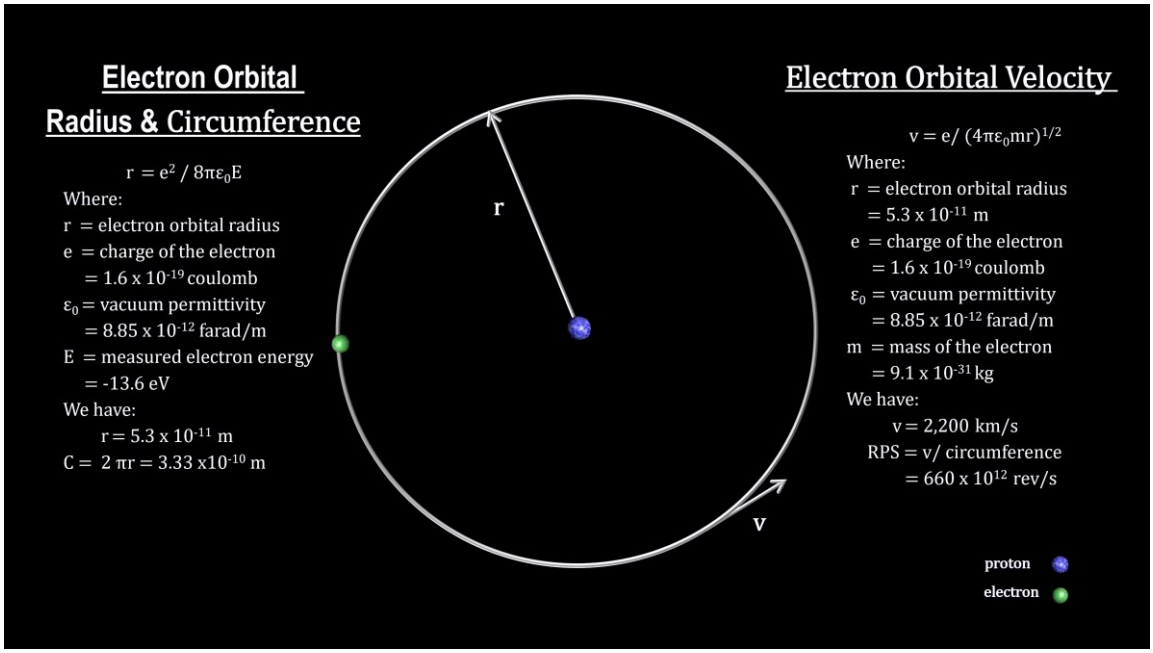


The Bohr Atom

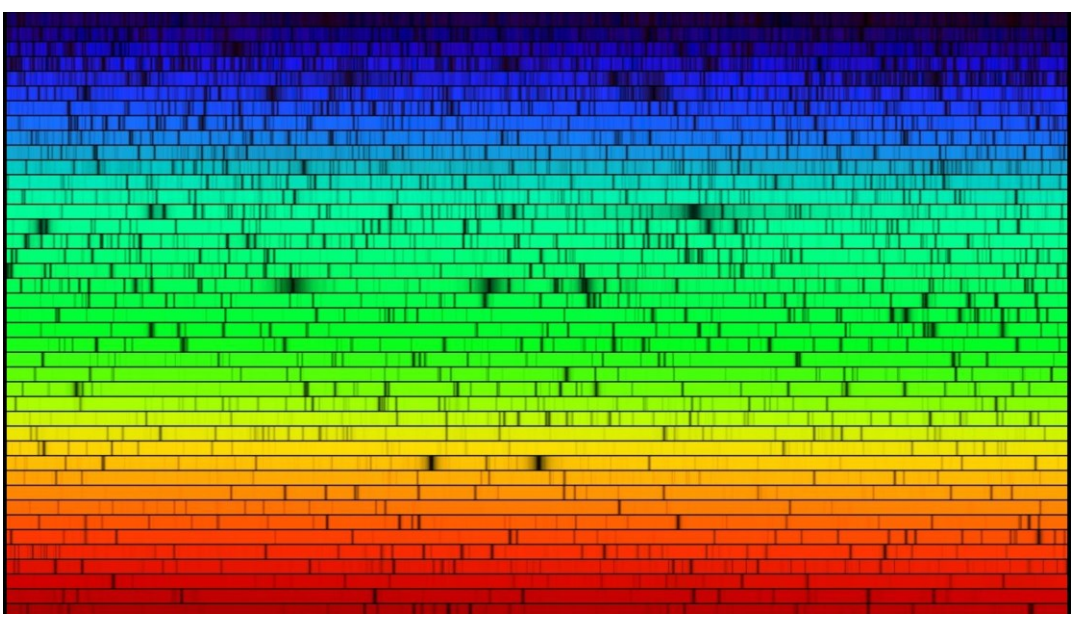
[Music: Tomaso Albinoni – "Adagio in G minor" –**]**

The Rutherford model of the atom left one outstanding problem. In the Thomson model, the electrons were stationary in the positively charged pudding. But what keeps a negatively charged electron from falling into the positively charged nucleus - given that opposite charges attract each other. The first proposed solution was to assume that the electron is in orbit around the nucleus like the Earth around the Sun. Just as we can use gravitational and centripetal forces to calculate the radius and velocity of a planet around the Sun, we can use electric and centripetal forces to calculate the radius, circumference, velocity, and revolutions per second of an electron around the nucleus.

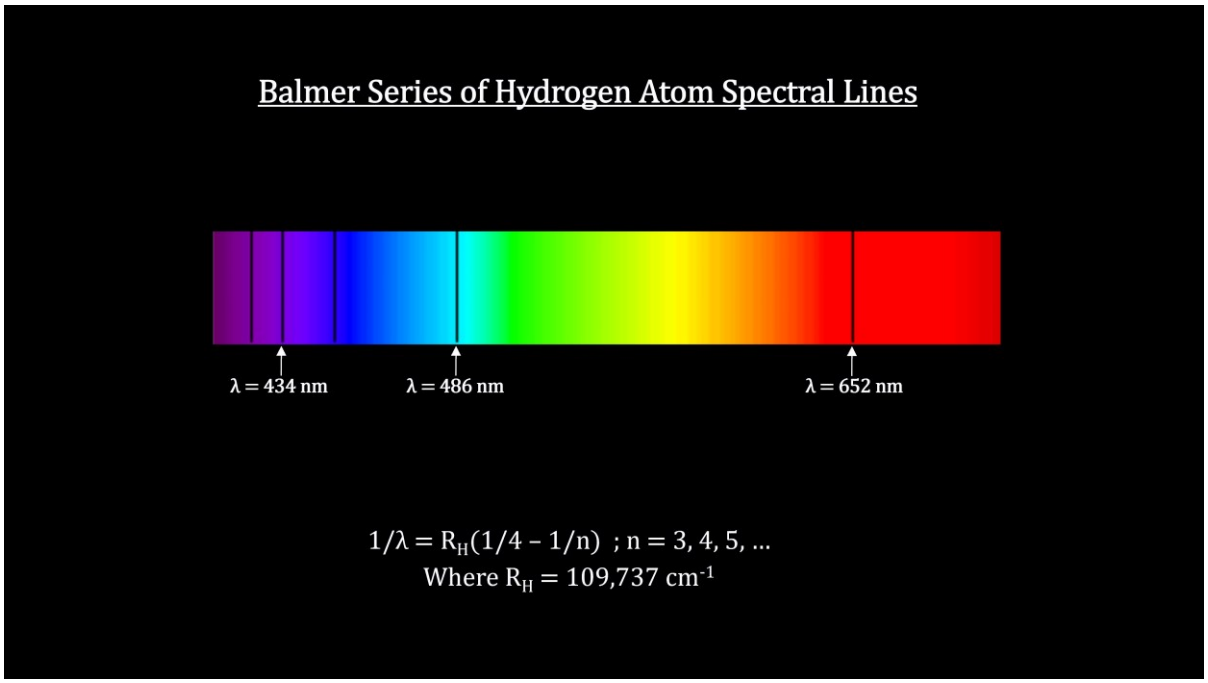
For Hydrogen, we get a very small circumference of around a third of a nanometer and a very large velocity of around 1% of the speed of light. That combination gives us a fantastically large 660 trillion revolutions every second.



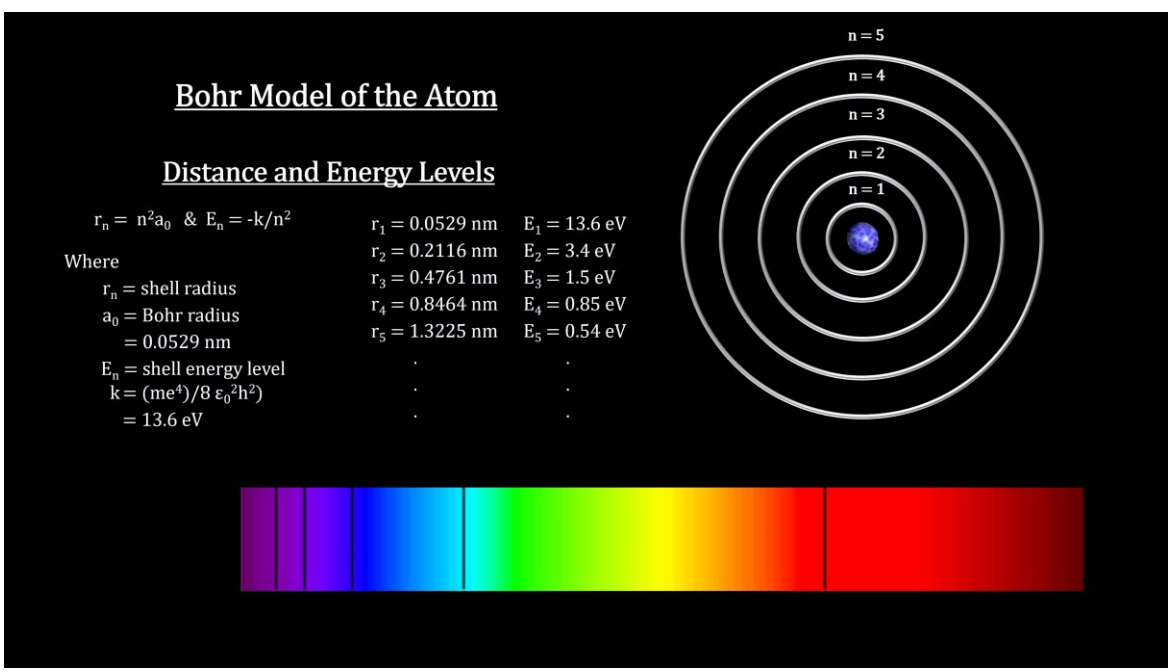
This would create a stable atom if the electron didn't have a charge. But classical electromagnetic theory points out that an accelerating charge radiates energy. Theoretically, the electron should collapse into the nucleus in less than a trillionth of a second. And yet, we see that it does not collapse. You'll recall from our "How far away is it" segment on "Distant Stars", that the light spectrum from stars was covered by thousands of dark lines called Fraunhofer lines or spectral lines. Although these lines had been studied for over a hundred years, no one understood what they were.



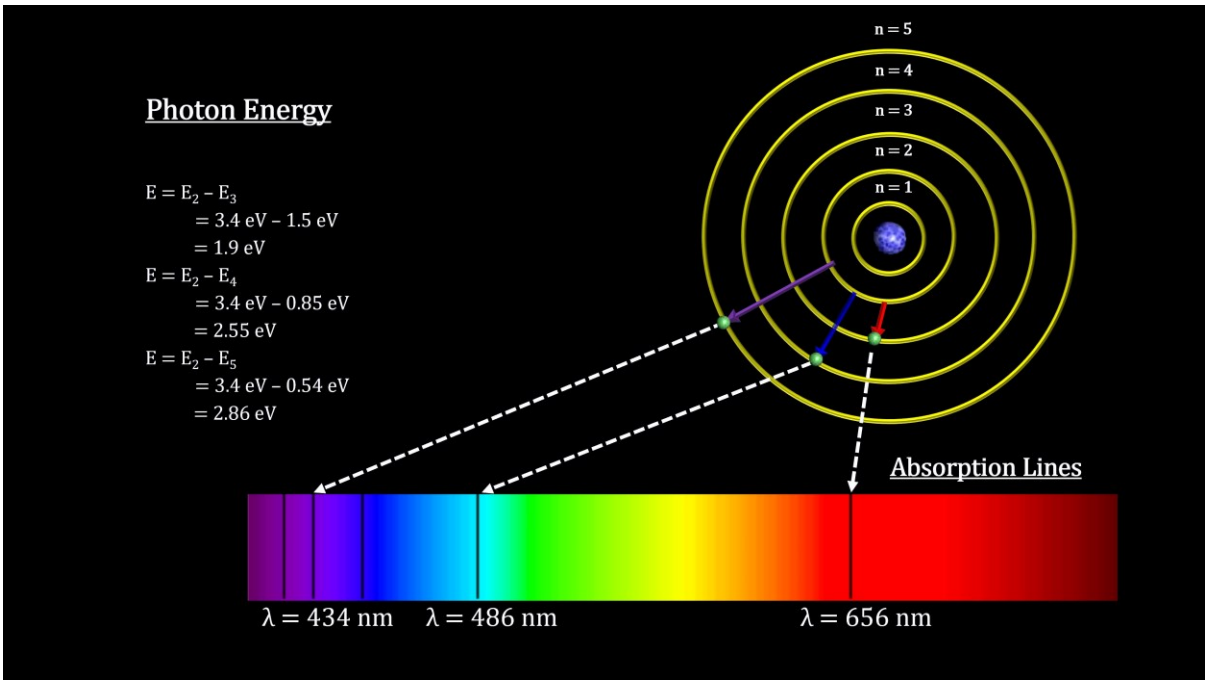
In 1885, Johann Balmer broke out a subset of these lines for Hydrogen and developed some mathematical interrelationships between them.



Then, almost 30 years later, Niels Bohr developed a quantized momentum theory for the atom that partially explained these lines. His model still had the electrons orbiting the nucleus, but they could only orbit at certain specific distances from the nucleus called shells. Each shell had its own unique energy level n, where n was a positive integer = 1, 2, 3, etc. These were called the atom's quantum numbers.

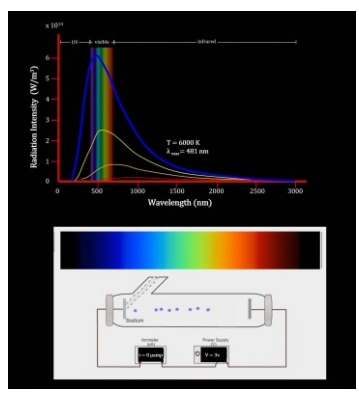


Electrons radiated or absorb energy when they change energy levels. The emitted or absorbed light has the energy difference between the levels. This energy is equal to Planck's Constant times the frequency of the emitted light. Here's how it works. When a photon with an energy E, hits an electron in a shell around a nucleus that has a higher shell it can reach with this same exact energy the photon's entire energy is transferred to the electron instantaneously. This jumps the electron to a higher energy level with a larger quantum number. The photon is eliminated. This creates 'absorption lines' in a star's spectrum as light from the star travels through the star's atmosphere. When the electron drops from this excited state back to a lower energy level, a photon with the exact difference between energy levels is emitted. This creates emission lines that we can see in the lab.

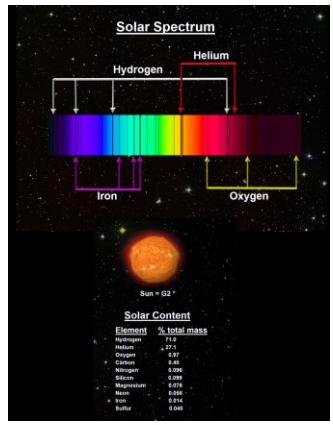


Bohr's model explained the Balmer series for Hydrogen spectra.

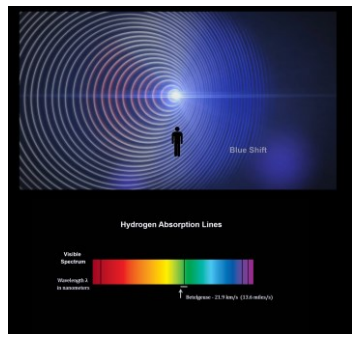
In addition, it provided the physical mechanism for Planck's quantized emission Blackbody Radiation and Einstein's quantized absorption Photelectric Effect.



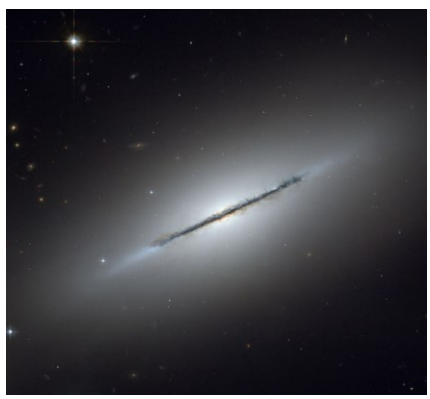
It was also momentous for astronomy. Every atom and molecule have their own unique spectral line signature. So, by observing the absorption lines in a star's spectrum, we can tell what the star is made of!

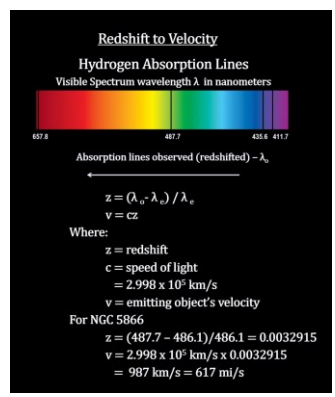


And not only that, by analyzing how these lines shift, we can calculate star radial velocities via the doppler effect,



and even use them to measure the expansion of the Universe.



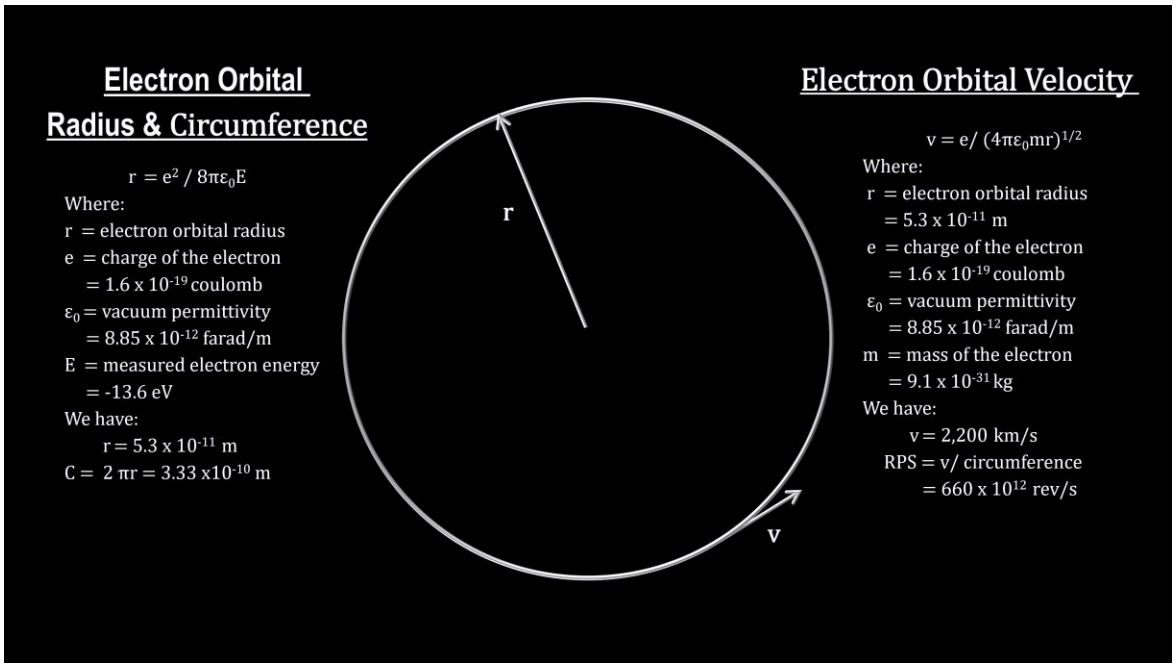


Indeed, Bohr's model explained a great deal. But there was no explanation for why the shell distances from the nucleus were as described, and there was no explanation for why the orbiting electrons didn't radiate away their energy and collapse into the nucleus.

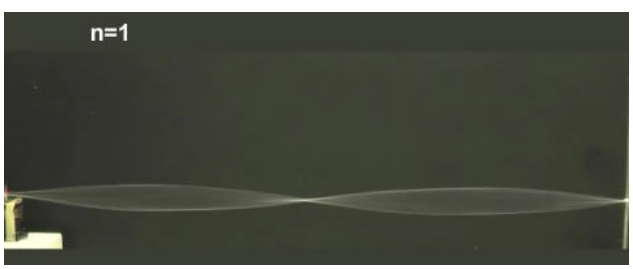
The de Broglie Atom

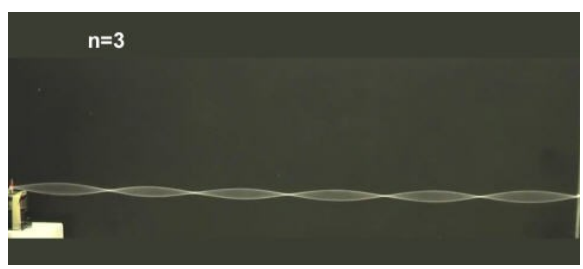
In 1925, Louis de Broglie came up with the model that explained how electrons avoid falling into the nucleus. Earlier, we calculated the circumference and velocity of the electron, so like we did for

electron microscopes in the previous segment, he calculated its wavelength. He found that it was exactly the length of the electron orbit's circumference as enumerated by Bohr! In other words, the wavelength of the electron is exactly the length of one revolution. This would create a standing wave!



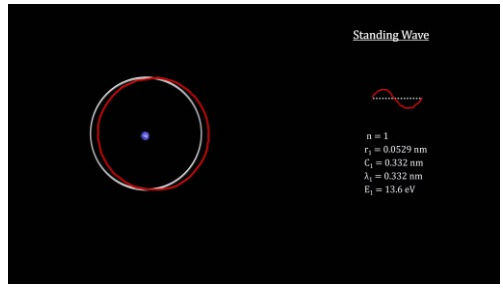
Here are a couple of standing waves on a string.



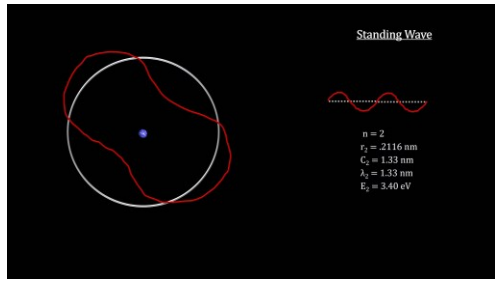


Here's a water standing wave. A standing wave is a wave constrained to vibrate in a distance that's an exact multiple of its wavelength. Anything more or less would create destructive interference and the wave would collapse.

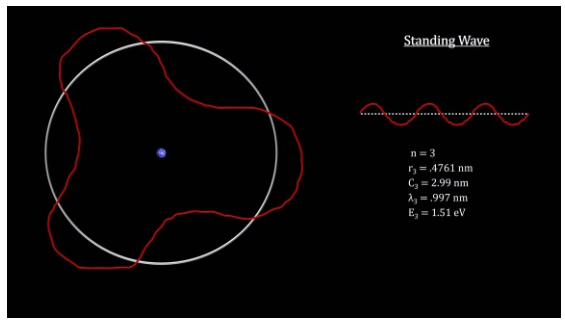
So, the first energy shell would have to have the radius that creates the circumference that exactly fits one wave.



The second shell would have to have the radius that creates the circumference that exactly fits two wavelengths.

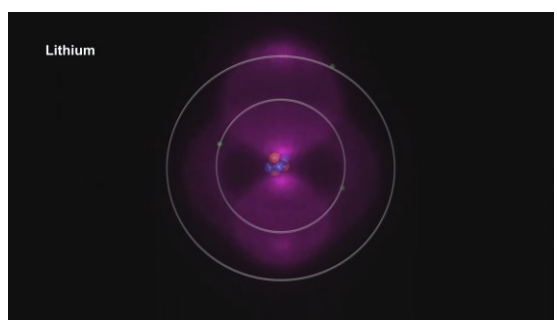


The third shell would have to have the radius that creates the circumference that exactly fits three wavelengths, and so on.



So, the proposed answer to the question "How can an electron sit way outside the nucleus without orbiting away its energy?" is that the electrons exist as standing waves that envelop the nucleus. No orbital motion is required and therefore, no radiation is emitted. Remember that: electrons in an atom do not 'orbit' the nucleus like planets around the sun – they exist as standing waves.

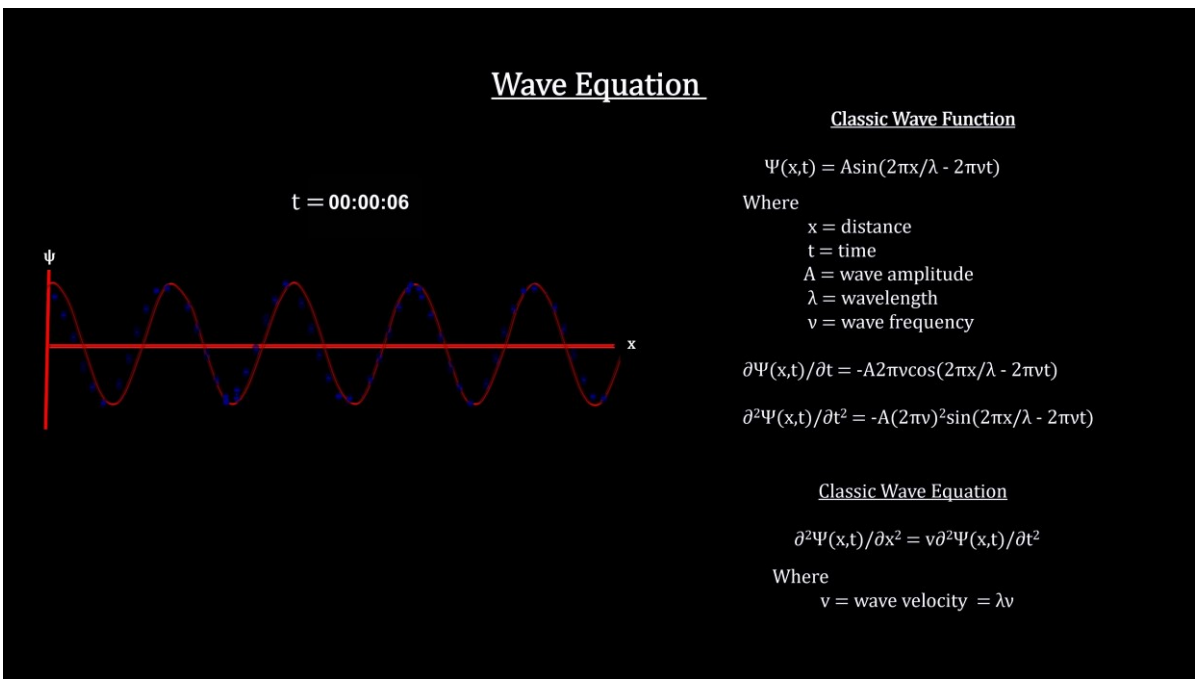
de Broglie's simple geometry elegantly explained the reason for each energy shell's distance from the center and its corresponding energy. But it didn't scale to explain the spectra of more complex atoms that have more electrons. [And it could not explain how individual atoms interact with one another to produce the physical and chemical properties that we observe in everyday life.]



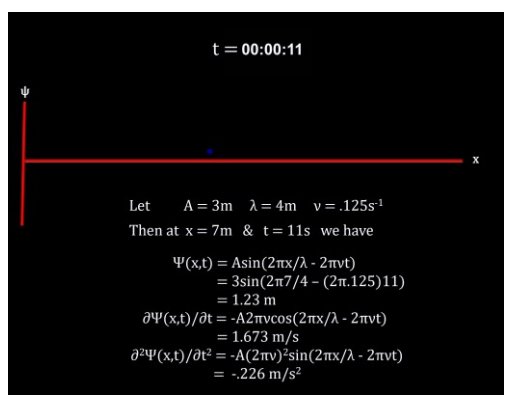
Schrodinger's Equation

Given that particles travel as waves, and are confined in atoms as standing waves, it followed that a generalized wave equation was needed to describe them. Building on the works of Planck, Einstein, Rutherford, Bohr, de Broglie and others, Erwin Schrodinger an Austrian physicist developed just such an equation, now bearing his name.

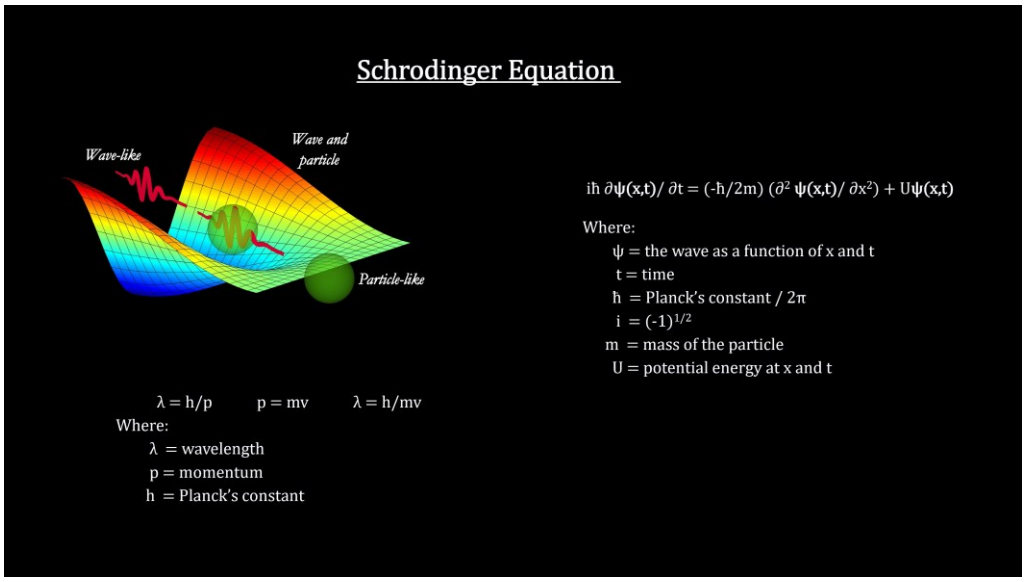
Here's a simple sine wave in water. It's described by a wave function. The function tells us the displacement of every water particle in the wave at any time t. If we take the change in a particle's displacement with respect to time, we get its velocity. And if we take the change in a particle's velocity with respect to time, we get its acceleration. A generalization called the wave equation describes how a wave function evolves over time.



Here's an example. If we take a look at the particle 7 meters down the line and take a snapshot at the 11 second mark, we see that it is just above the line, heading up rapidly and slowing down slightly.

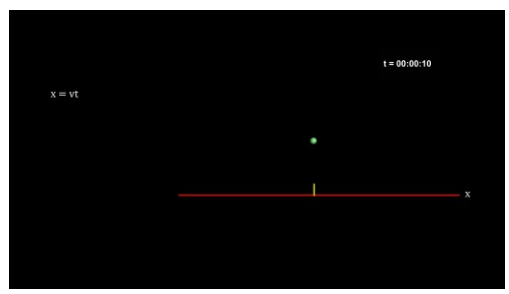


Schrodinger used the fundamental relationships between energy and wave frequency and quantized momentum to develop a quantum mechanical equivalent of the wave equation. Importantly, it had one critical difference with the classical model - it did not produce a location for a particle. In fact, it did not represent an observable physical quantity at all. Instead, it produced a probability curve for particle location.

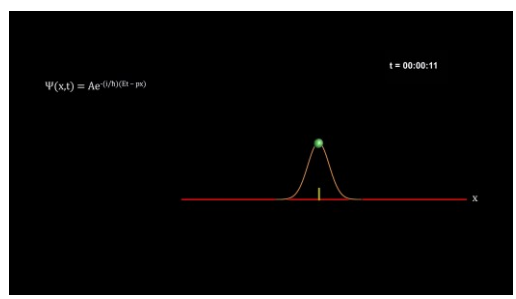


For free particles, the square of the wave function gives us the probability of experimentally finding the particle at a particular location at a particular time.

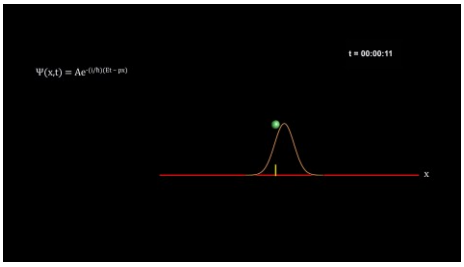
For example, suppose we had a particle moving from left to right at a specific speed. From Newton's equations, the distance x is equal to the speed v times the time t. After 24 seconds, we would say that the particle is here.



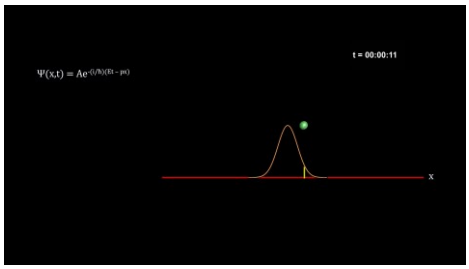
But, because the particle moves as a matter wave, we need to use Schrodinger's equations. So, when you touch the wave at time t, it collapses into a particle. Where classical physics says it is here, quantum mechanics says that 'here' is the most likely place.



But there is a smaller probability that it is here.

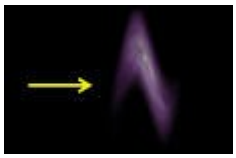


Or an even smaller probability that it is here. In fact, there is a chance that it may be anywhere along this probability curve, with the probabilities dropping rapidly as we move away from the most probable point.

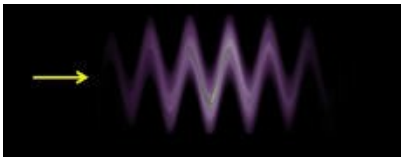


Heisenberg Uncertainty Principle **[Music:** Dave Porter – "Breaking Bad" –**]**

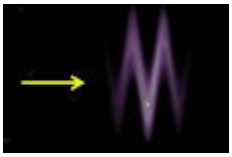
This brings us to the Heisenberg Uncertainty Principle. In our quest to understand how small things can get, we need to know if there is a measure of size below which we can't go. We see from our little thought experiment that, as a wave, a particle's location is not fixed. The wave is spread out. Here we see three different wave packets for an electron.



The wave packet at the top is narrow and therefore easier to locate, but it is less than one wavelength, so its momentum is impossible to figure out.

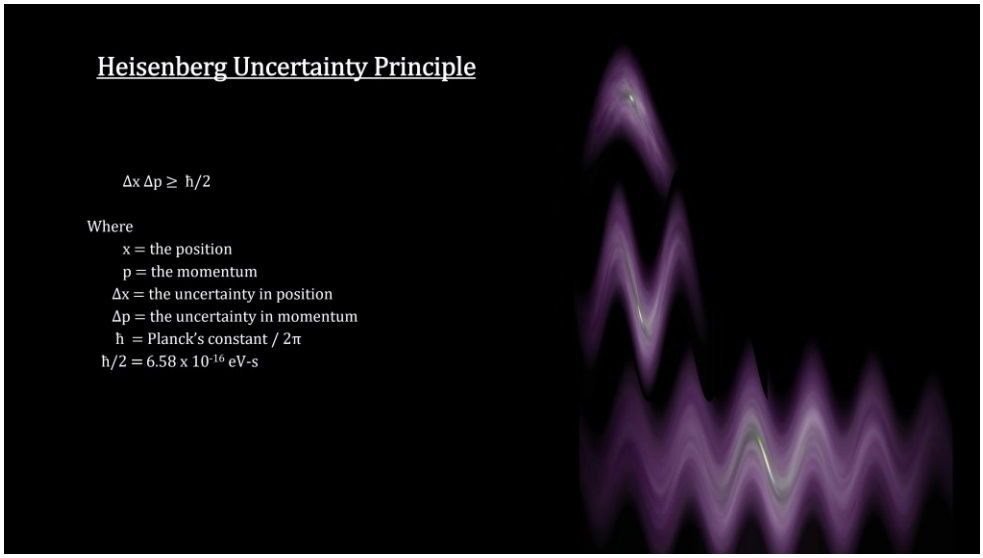


The bottom wave packet contains plenty of wavelength information, but it is quite spread out and its location is more uncertain.

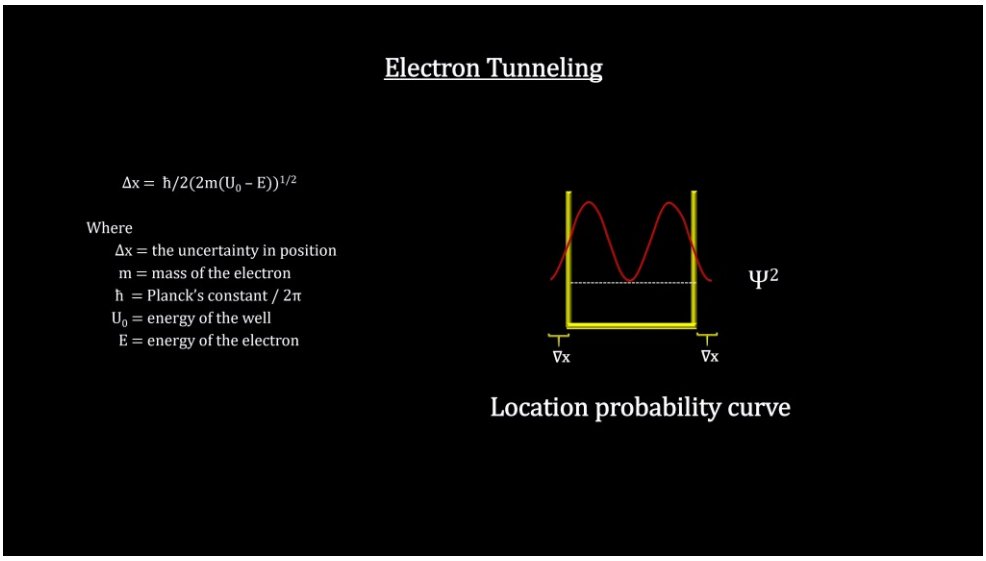


The wave packet in the middle has enough wavelength information to make its momentum less uncertain, and it is less spread out than the one on the bottom making its location less uncertain.

But, due to the spread-out nature of matter waves, we still can't know both the location and momentum at the same time. Mathematically stated, the uncertainty in position times the uncertainty in the momentum is always greater than or equal to Planck's constant divided by 4 π. This is the Heisenberg Uncertainty Principle. It has nothing to do with the accuracy of our instruments and everything to do with the wave nature of matter.

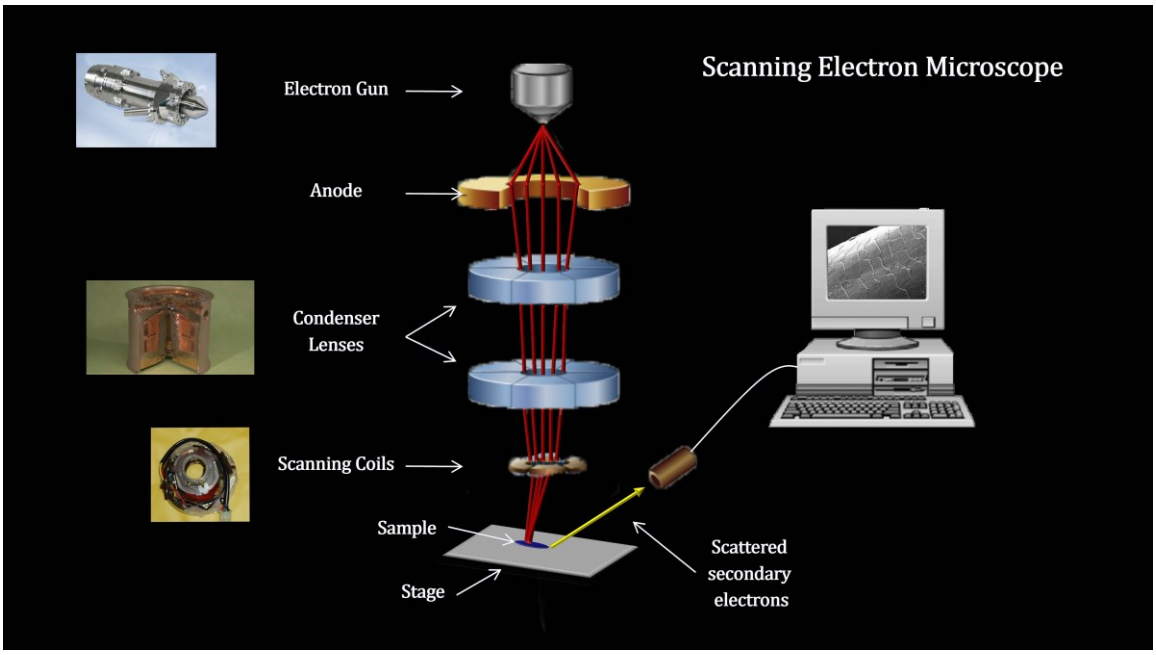


A good way to illustrate this is to look at an electron in an energy well too deep for it to get out. But remembering that the electron has a wave function that gives the probability of finding it at any given point, and some of these points (admittedly with very very low probability) can be found outside the walls of the well – as if it had tunneled through the wall when in fact it did not.

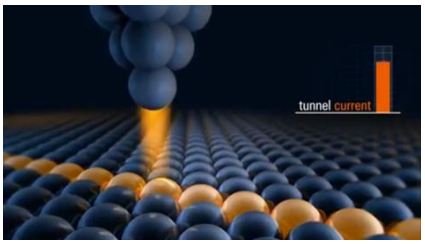


Scanning Tunneling Electron Microscope (STM) **[Music:** Georges Bizet – "L'Arlésienne" –**]**

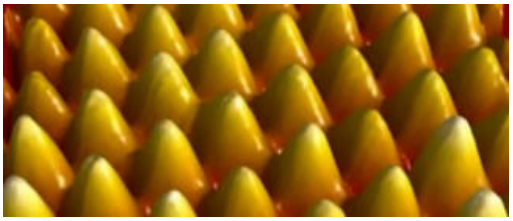
In our 'How Small Is It' chapter on the microscopic, we covered scanning electron microscopes that mapped the surface of an object by using the wave nature of electrons and analyzing their scattering properties.



Here we will cover Scanning Tunneling Microscopes or STM for short, that use the quantum mechanical tunneling property. Here's an STM at the Max Planck Institute.



It has a small pin head that is actually one single atom at its tip. The tip is brought close enough to the object for electrons to tunnel across the space exactly in accordance with Schrodinger's equation. This creates an electric current.

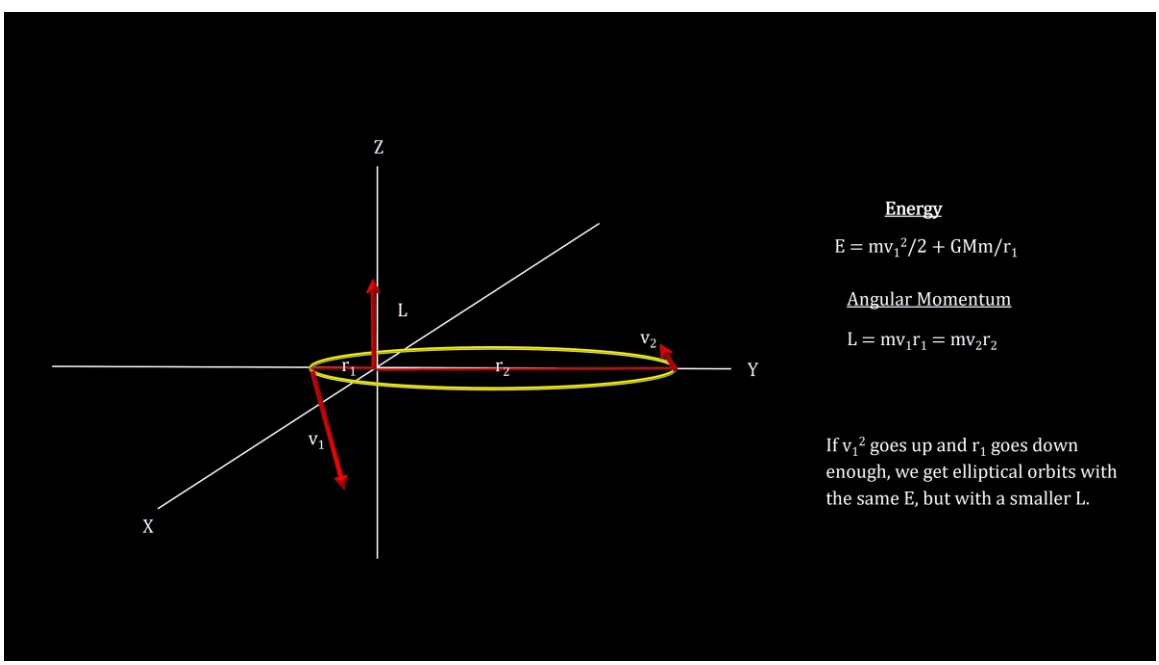


As the tip scans across the object, the current will go up or down depending on weather an atom is under the tip or not. This is repeated over and over till the entire surface is mapped. What we are doing is actually feeling the surface of the object to see and measure the atoms. The resolution reaches 0.01 nm. That's about 1/100th the diameter of an atom.

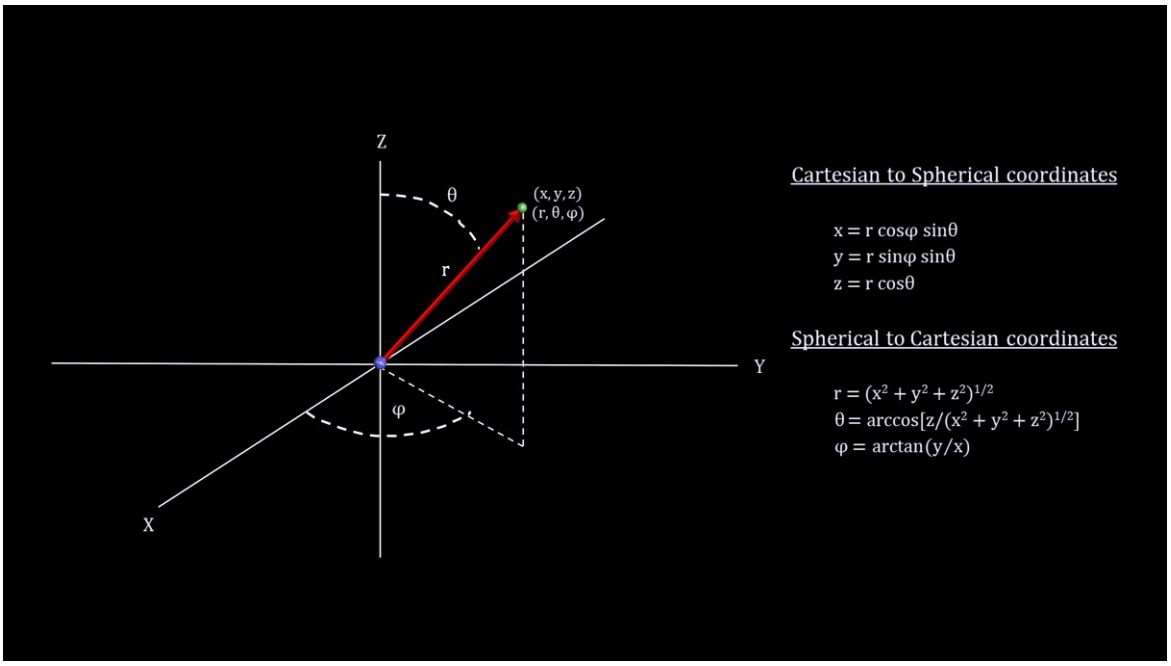
With a little stronger pull, we can even dislodge and move atoms. Here we see that the scientists at the Max Planck Institute moved the atoms one by one to spell their institute's initials MPI. The tag is just 6 nanometers wide.

Schrodinger's Atom

Before we get back to the atom, it's helpful to examine macroscopic orbital systems to see what varies with respect to energy, angular momentum, and orientation. In Newtonian Mechanics, we can calculate the angular momentum L for an elliptical orbit and its energy E. [The system depends on mass, the distances from the center mass, and the velocity of the orbiting objects. It turns out there are no limits on the number of combinations of distance and velocity that can produce the same value for energy.] We see that there are no limits to the number of angular momentum values L that can be associated with any particular system energy E. We also note that there are no restrictions on the orientation or azimuth angel for any angular momentum L.



Because we are dealing with a spherical system with the bulk of the mass at the center, it is common practice to use spherical coordinates. **r** is the vector specifying the position of the electron relative to the proton. Its length is the distance between the two and the direction is the orientation of the vector pointing from the proton to the electron. Theta is the polar angle most closely related to angular momentum. And Phi is the azimuthal angel associated with orientation.



But when we move from matter systems to matter wave systems, we move from Newtonian equations for gravitationally bound systems to Schrodinger equations for negatively charged electrons bound by a positively charged nucleus. The relationships between energy, angular momentum and orientation are quite different. With Schrodinger's equation for the Hydrogen atom in spherical coordinates, we can separate the variables R, Theta and Phi.

$$-(\hbar/2m)[(\partial^2\psi/\partial r^2) + 2\partial\psi/r\partial r + (1/r^2\sin\Theta)(\sin\Theta\,\partial\psi/\partial\Theta) + (1/r^2\sin^2\Theta)(\partial^2\psi/\partial\varphi^2)] + U(r)\psi(r,\Theta,\varphi) = E\psi(r,\Theta,\varphi)$$

For a Coulomb force that does not change with polar or azimuthal angles we can separate the variables for the wave function as

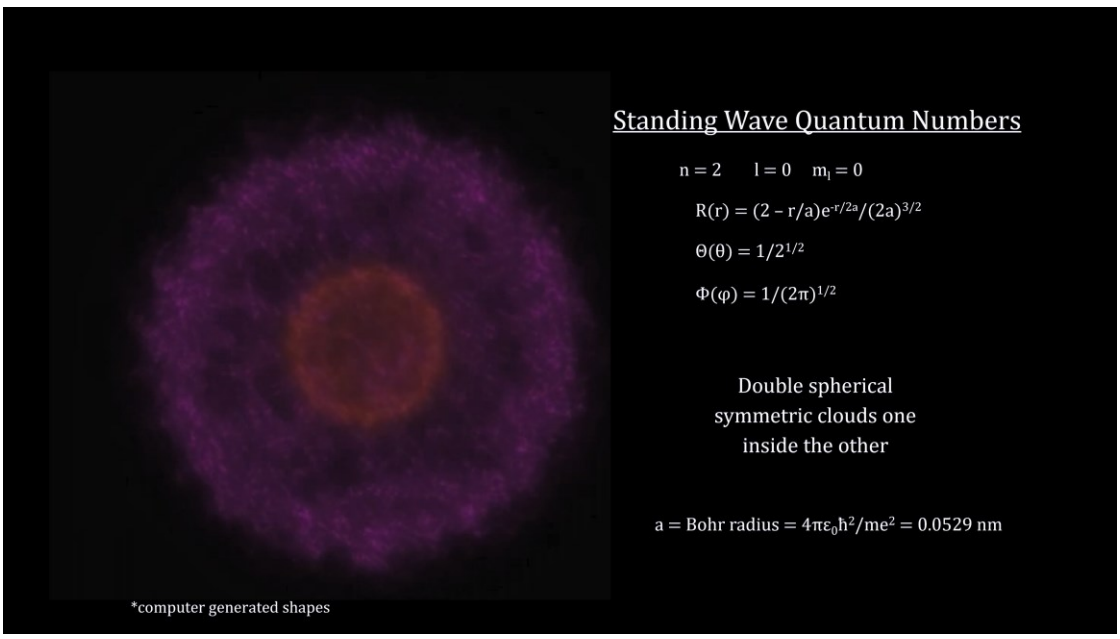
$$\psi(r,\theta,\varphi) = R(r)\Theta(\theta)\Phi(\varphi)$$

$|\psi(r,\theta,\varphi)|^2$ = probability density cloud

Solving Schrödinger's equation yields multiple wave functions as solutions. They define an electron's probability density cloud.

- Energy is quantized into electron shells designated by the letter *n*. It determines the distance the electron is from the nucleus. These energy levels match the ones proposed by Bohr.
- For each energy level *n*, the associated angular momentum is also quantized into electron sub-shells designated by the letter ℓ. It determines the shape of the orbital.
- And surprisingly, for each quantized angular momentum sub-shell, even the allowed orientations are quantized into orbitals and designated by the letters m_ℓ. It determines the orientation of the orbital.

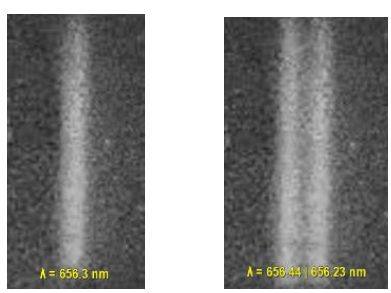
In chemistry, an atomic orbital is defined as the region within an atom that encloses where the electron is likely to be 90% of the time. It is these radii with their binding energies and interesting geometries that give atoms their chemical properties.



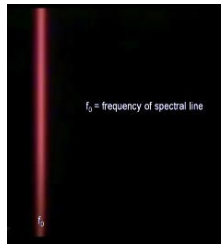
The Pauli Exclusion Principle

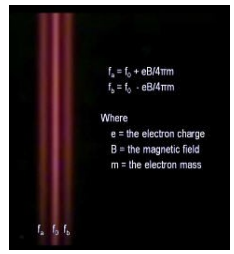
Although Schrodinger's equation went a lot further than Bohr and de Broglie, there were still a couple of things about the atom that were not completely explained.

1. When examined very closely, many spectral lines showed up as pairs instead of single lines as called for by Schrodinger's equation.



2. The splitting of spectral lines by magnetic fields was not accounted for. This is known as the Zeeman effect.

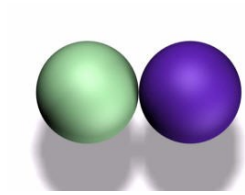




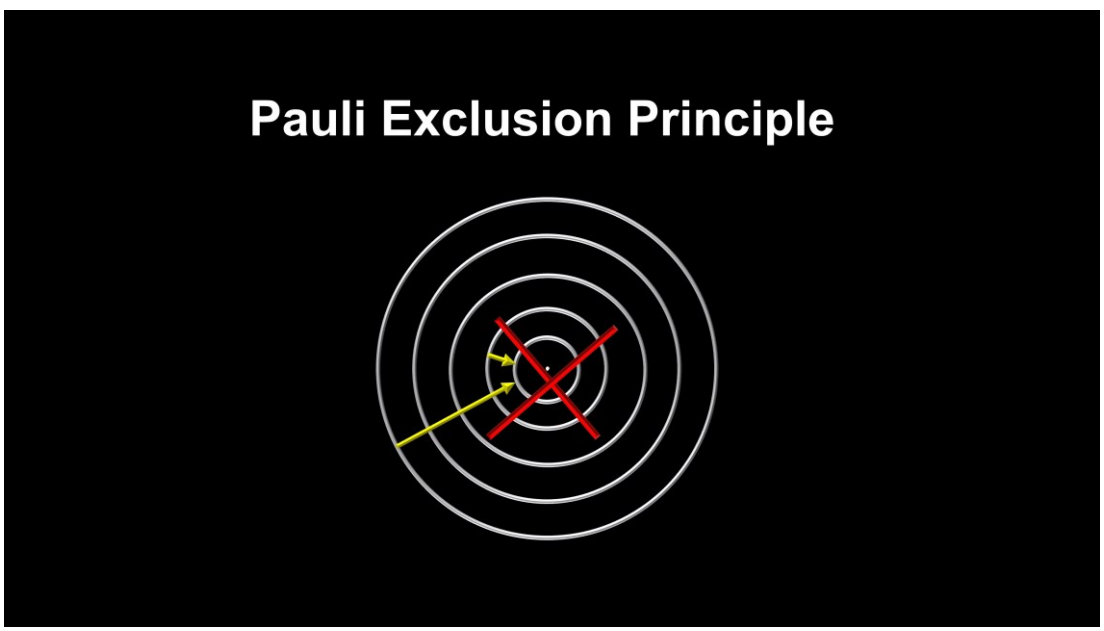
3. It was not understood why all the electrons didn't all move to the innermost lowest energy orbital.

In order to deal with these issues, Wolfgang Pauli proposed a fourth quantum number and his exclusion principle.

In classical physics, the exclusion principle states that no two objects can occupy the same space at the same time.



Pauli's exclusion principle stated that no two particles could occupy the same quantum state at the same time. But Pauli could find no physical explanation for the fourth quantum number.

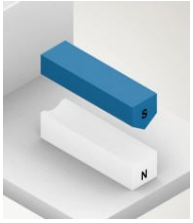


Electron Spin - Stern-Gerlach experiment

The physical explanation turned out to be electron spin. Electrons have an intrinsic property that is best observed with a modern version of the Stern-Gerlach experiment that used silver atoms.

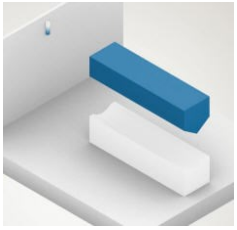
Here we use magnets and electrons directly.

The device has a north and south pole shaped to create a magnetic field that is stronger near the tip. This varies the forces on charged particles passing through.

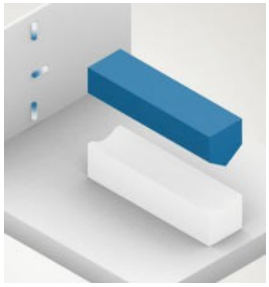


A magnet is sent through with the north pole up and the south pole down.

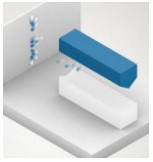
The magnetic field creates a force that deflects the magnet upward as it passes through the field.



As we change the orientation of the magnets being sent through, we see the change in the amount and direction of the deflections. The deflections depend on the orientation. This is as expected.

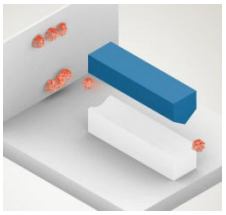


When we send large numbers of randomly oriented magnets through the field, they arrive anywhere vertically.



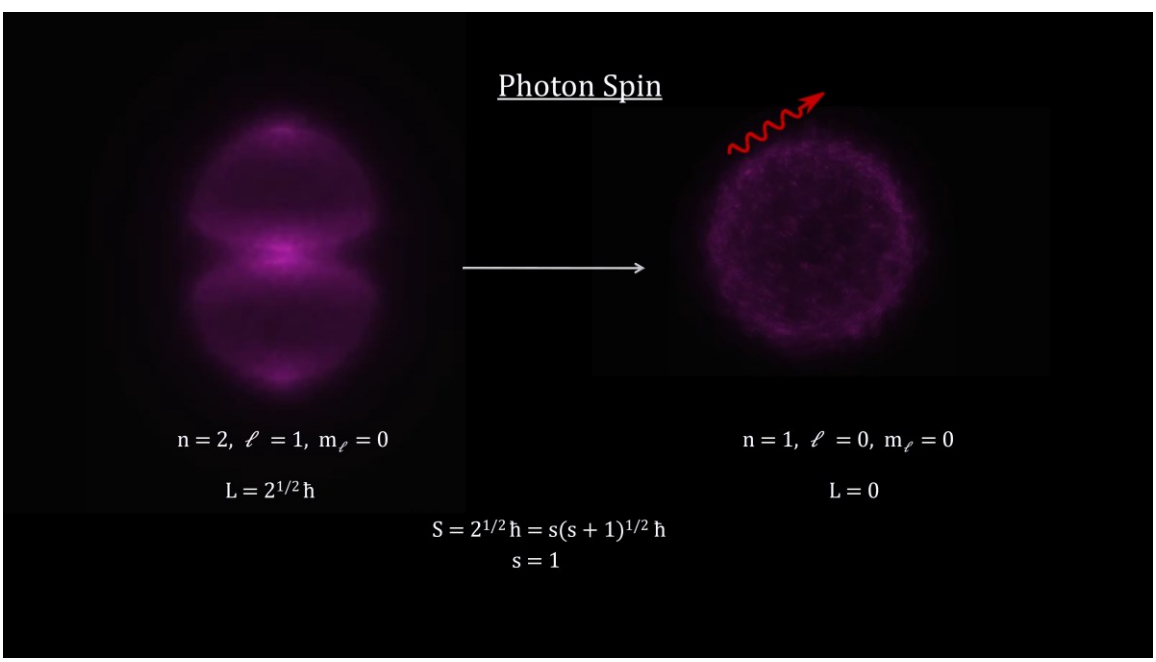
When electrons are sent through the field, they too are deflected.

But they always arrive at the screen deflected either up or down. Never in between like the magnets.



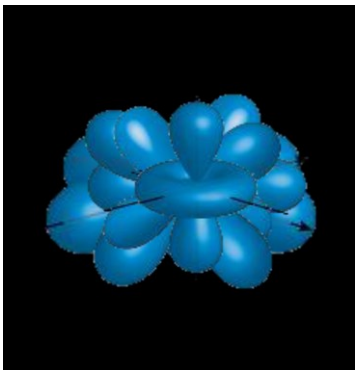
Each electron behaves as a magnet, but with only one of two possible orientations: up or down. This intrinsic property of an electron is called 'spin'.

It is interesting to note that, whenever an electron in an atom changes state, the atoms angular momentum changes. For example, here an electron moves from a higher energy orbital with angular momentum to a lower orbital with no angular momentum.

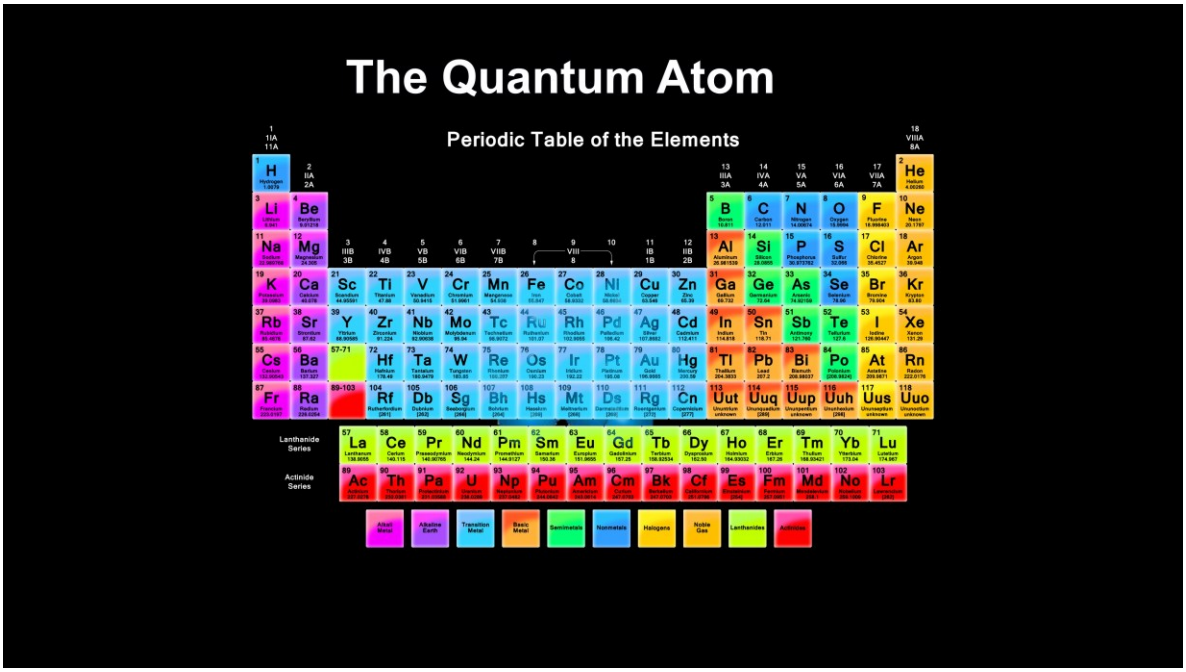


We see that the emitted photon carries away both the energy and the angular momentum, giving it a spin = 1. This has been measured to be true for all electron quantum leaps.

With the Pauli Exclusion Principle and spin as the fourth quantum number, the full set of spectral lines, orbitals, their geometries, and interactions with each other fell into place.

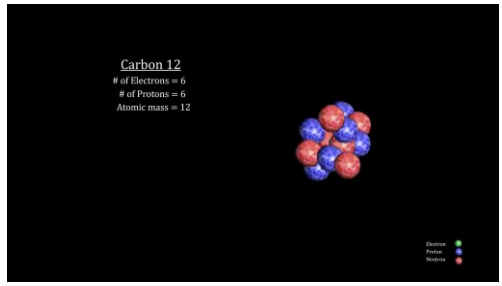


In fact, when we add this forth quantum number to Schrodinger's equations, we can generate the entire periodic table of the elements.



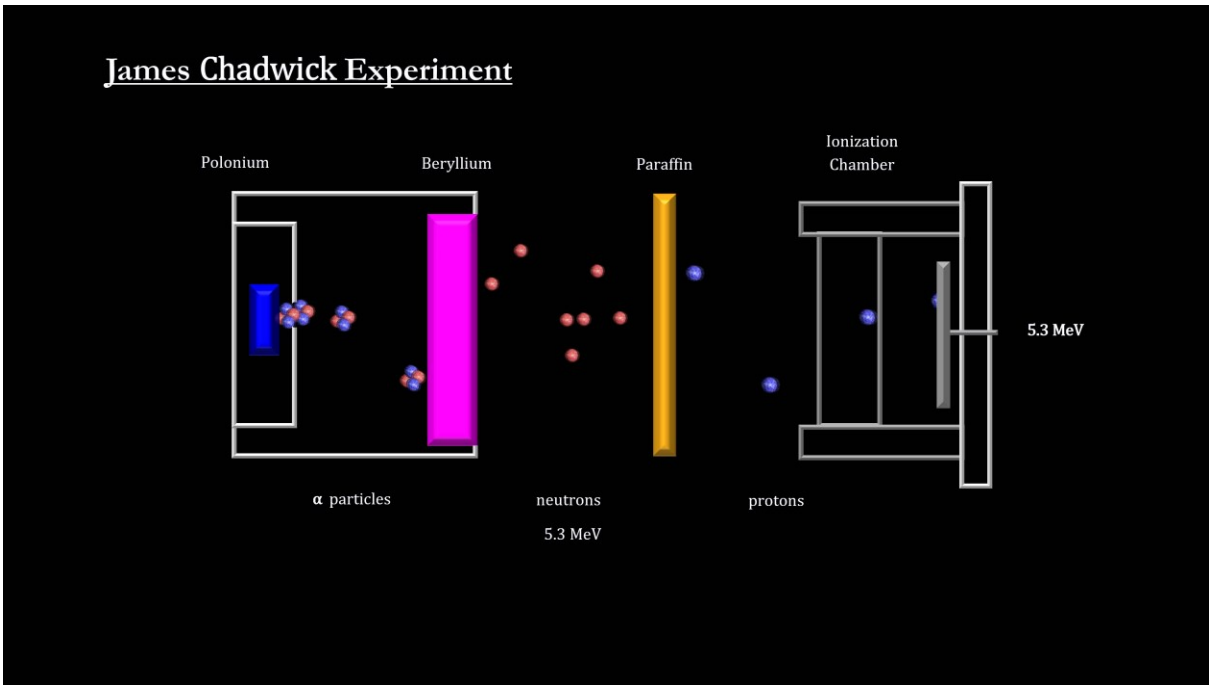
The Nucleus

Now that we have a handle on the electrons around the atom, let's take a quick look at the nucleus. For atoms to be neutral, the number of protons with a positive charge must equal the number of electrons with their negative charge.



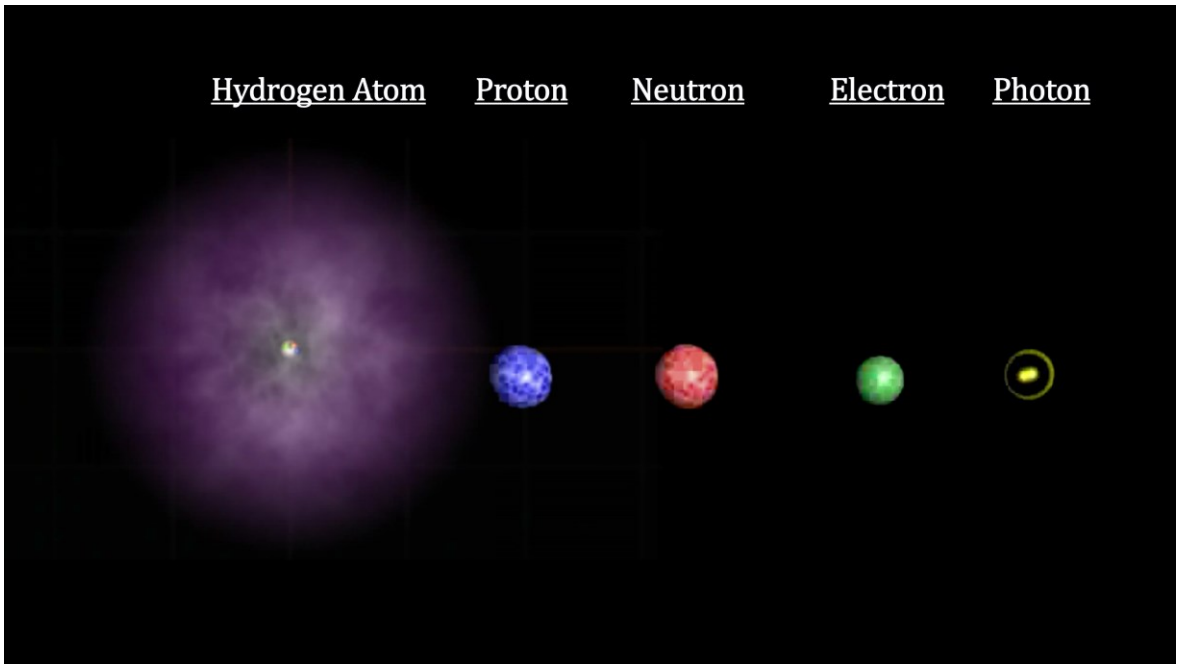
But mass spectrometers showed that atoms have more mass than the number of protons alone could account for. For example, carbon has 6 protons and 6 electrons, but its mass is just a tad more than the mass of 12 protons.

In the 1920s it was assumed that electron-proton pairs existed in the nucleus to account for the increase in mass without an increase in charge. But with the advances in quantum mechanics, it became clear that an electron couldn't exist in a volume as small as the nucleus. Ernest Rutherford and James Chadwick proposed that a new particle (the neutron) must exist in the nucleus to account for the data. In 1932, Chadwick and others performed a series of experiments verifying his suggestion. They began by beaming alpha particles into Beryllium. This produced a radiation that was not affected by applied electric fields. In other words, it was electrically neutral. At first, this was thought to be gamma rays. But when this new radiation was used to bombard a hydrogen rich substance like Paraffin, a Proton radiation was produced. The energy acquired by these protons was measured and found to be more than a gamma ray could possibly impart to a proton. In fact, the protons ejected from the paraffin on the right was equal to the energy of the radiation coming out of the Beryllium on the left. The conclusion was that the particles hitting the paraffin were of the same mass and energy as the protons but without any charge. At this point it was generally accepted that the neutron had indeed been discovered.



How Small Is It

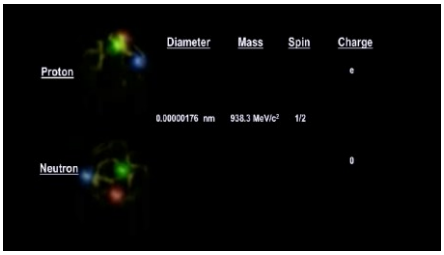
In this segment, we developed the basic quantum mechanics for electrons around the atom and measured the size of atomic components. At the end of our previous segment, we used a Scanning Electron Microscope to see carbon atoms 14 hundredths of a nanometer in diameter.



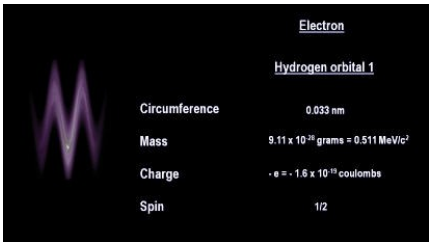
Using Rutherford scattering techniques covered in this segment, we measured the size of a proton at 1.76 millionths of a nanometer. That's 20,000 times smaller than the atom. At this scale, we find that the neutron is about the same size and mass as the proton.

Also, in this segment we added spin as an intrinsic property of particles to go along with mass and electric charge. Protons and Neutrons both display the same spin properties as electrons when they traverse the Stern-Gerlach apparatus, so their spin is ½.

The notable difference between these particles is that the Proton has a positive charge with the same magnitude as the electron's negative charge, but the Neutron is neutral with no charge at all.

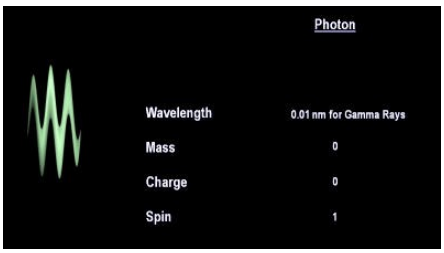


For the electron, it's hard to talk about its size because their wave packet is different for every circumstance from standing waves in thin atom shells, to scattered waves in an electron microscope. What we did in this segment was to calculate its length around the Hydrogen nucleus at .0033 nm.



For photons, we see that they have no mass at all, no charge, and a spin equal to 1.

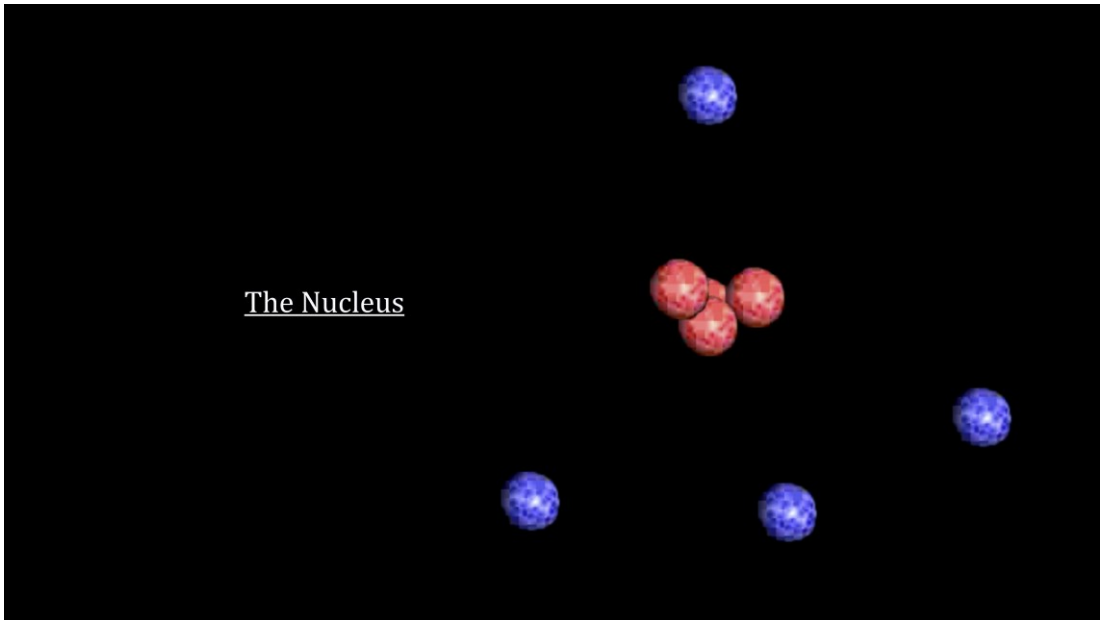
And, like the electron, it doesn't have a volume per say because it's a wave. But we can measure its wavelength. For the gamma rays used by Rutherford, the wavelength is one one hundredth of a nanometer. That's 51,000 times smaller than the wavelength of green light.



Looking at the atom's nucleus, we see one main question:

- How do positively charged protons pack together in the nucleus when their repulsive positive charges would have them flying apart?

We'll go into how we answer this question in our next segment on elementary particles.



Music

@00:00 Albinoni - Adagio in G Minor - Berlin Chamber Orchestra; from the album "50 Must-Have Adagio Masterpieces" 2013

@05:02 Haydn - Cello Concerto No 2 II Adagio - Franz Liszt Chamber Orchestra, Miklos Perenyi; from the album "50 Must-Have Adagio Masterpieces", 2013

@09:12 Mozart - Violin Concerto No 3 - Christian Altenburger, German Bach Soloists; from the album "50 Must-Have Adagio Masterpieces" 2013

@20:01 Dave Porter - Breaking Bad Theme - Music from the Original Series Breaking Bad

@22:16 Bizet - L'Arlesienne Suite No 1, Op 23 - Budapest Philharmonic Orchestra; from the album "50 Must-Have Adagio Masterpieces" 2013

@26:57Mahler - Symphony No 5 III IV - Budapest Festival Orchestra; from the album "50 Must-Have Adagio Masterpieces" 2013

Greek letters:
- α β γ δ ε ζ η θ ι κ λ μ ν ξ ο π ρ σ τ υ φ χ ψ ω
- Α Β Γ Δ Ε Ζ Η Θ Ι Κ Λ Μ Ν Ξ Ο Π Ρ Σ Τ Υ Φ Χ Ψ Ω

⇒ → ± ⊙ ∞ ↛ ∃ ∄ ∈ ∉ ∬ ∫ ≅ ≥ ≤ ≈ ≠ ≡ √ ∛ ~ ∝ ħ ÷ ∂