The Atom

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

We start with J.J. Thompson'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 it is that protons can hold together in the nucleus when there like charges should push them apart.}

Introduction

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.

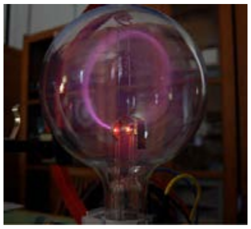
In our first segment, we used photons and electrons to see things down to the size of a carbon atom - 0.14 nanometers. That's small. Atoms are so small, that there are as many atoms in your DNA as there are stars in the Milky Way galaxy. In fact, there are more atoms in the breath of air I just took than there are stars in the visible Universe!

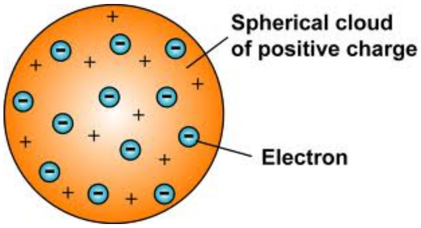
The structure of the atom is responsible for nearly all the properties of matter that have shaped the world around us and within us. But what do we actually know about atoms and how small are the particles that combine to create an atom. That's what this segment is all about. We'll start with early guesses 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.

> *[**Music:** Joseph Haydn – "*Cello *Concerto No. 2 in D" –This concerto was composed in 1783 for a cellist and Orchestra. The piece's authenticity was doubted for some time, but most experts now believe that the work is indeed authentic after Haydn's autograph score was discovered in 1951.]*

The Thompson 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 Thompson that first introduced the idea that an atom had parts.



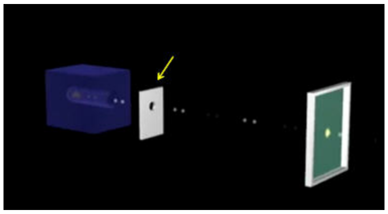


In 1898, with the electron being so light compared to the atom, Thompson suggested what is called the "plum-pudding" model of the atom – with a uniform mass of positively charged matter containing spots of electrons imbedded in it like plumbs in a pudding.

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.

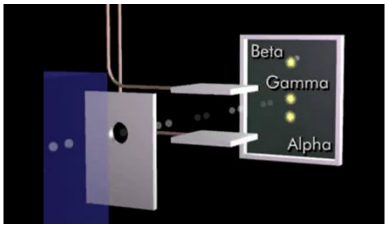
Radioactivity

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 effected 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.

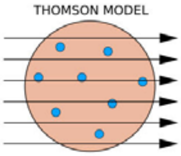
The Rutherford Atom

> ***[Music:** Mozart – "Violin Concerto No. 3" – This beautiful concerto was composed in 1775. Mozart was only 19 at the time.**]***

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.

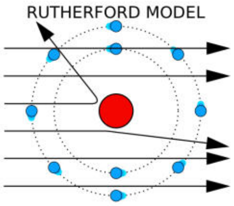
This beam is directed at a very thin gold foil. A movable zinc sulfide screen is placed on the other side of the foil. Zinc sulfide flashes when hit by an alpha particle. A microscope swivels to view all scattering angles.

If the Thompson 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.



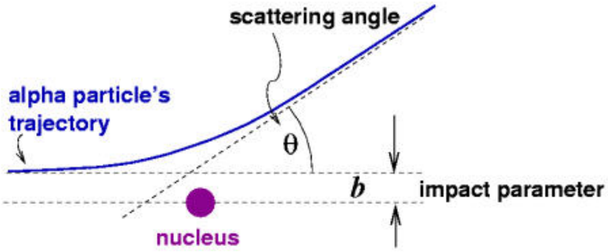
But after days of observations here's what they found.

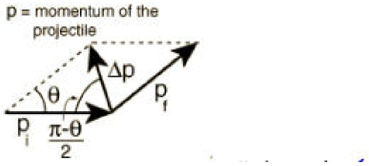
While most of the alpha particles do go right through with only minor deflections, some were scattered through very large angles. A few were even scattered in the backward direction! Rutherford described it as "almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came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





To analyze scattering, we use geometry similar to what we used in parallax analysis to measure distances to nearby stars. Rutherford also made careful measurements of the energy and momentum of the alpha particles before and after the collisions.

From the geometry we have:

$$\Delta p = 2mv\ \sin(\theta/2)$$

Where:
p = alpha particle momentum
m = mass of the alpha particle
v = velocity of the alpha particle
θ = scattering angle

$$r_0 = 2Ze^2 / 4\pi\varepsilon_0 T$$

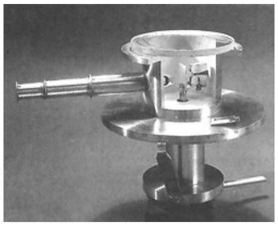
Where:
r_0 = max radius of the nucleus
Z = number of protons in nucleus
= 79 for gold
e = charge of the electron
ε_0 = vacuum permittivity
T = energy of the alpha particles
= 7.7 Mev

$$r_0(Au) = 0.00003\ nm$$

At this level we use electron volts instead of joules. 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. The energy of the naturally occurring alpha particles used by Rutherford where 7.7 million electron volts.

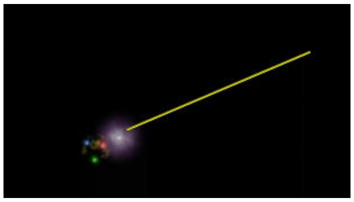
Rutherford's calculations showed that the diameter of a gold atom nucleus was 0.00003 nanometers. That was ten thousand times smaller than the size of a gold atom.

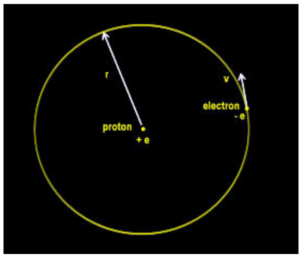
Here's a picture of the test apparatus Hans Geiger (of Geiger counter fame) and Ernest Marsden built to look for scattered alpha particles from every angle. The microscope could be swiveled all the way around the gold foil.



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 find the Higgs Boson.

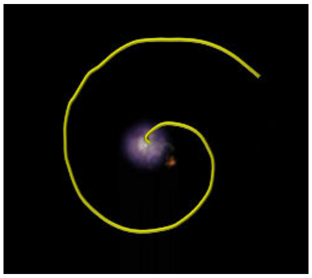
The Rutherford model of the atom left one outstanding problem. In the Thompson model, the electrons were stationary in the positively charged pudding. But what keeps a negatively charged electron from falling into the positively charged nucleus - given that opposite charges attract each other.





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.



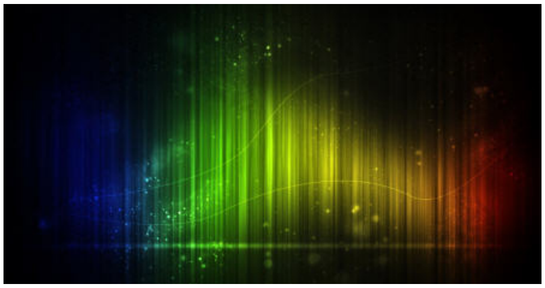
This would create a stable atom if the electron didn't have a charge. But classical electromagnetic theory points out that an accelerating charge radiates energy. Theoretically, the electron should collapse into the nucleus in less than a trillionth of a second.

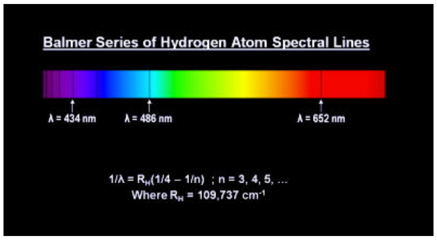
And yet, we see that it does not collapse.

The Bohr Atom

> *[**Music:** Tomaso Albinoni – "Adagio in G minor" – This Adagio for violin, strings and organ, is a neo-Baroque composition popularly attributed to the 18th-century Venetian master Tomaso Albinoni, but actually composed by the 20th-century musicologist and Albinoni biographer Remo Giazotto, purportedly based on the discovery of a manuscript fragment of Albinoni. I fell in love with it watching the movie Gallipoli.]*

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.





Balmer Series of Hydrogen Atom Spectral Lines

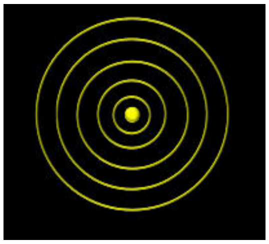
$1/\lambda = R_H(1/4 - 1/n)$; $n = 3, 4, 5, \ldots$

Where $R_H = 109{,}737\ \text{cm}^{-1}$

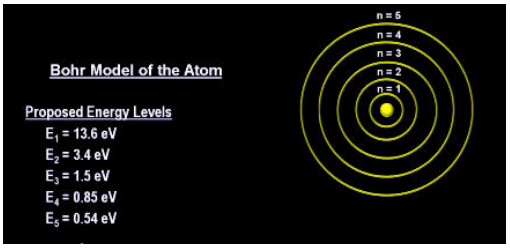
In 1885, Johann Balmer broke out a subset of these lines for Hydrogen and developed some mathematical interrelationships between them.

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.



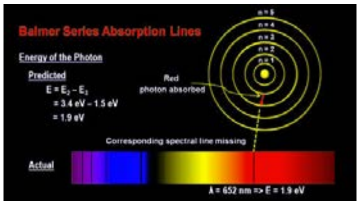
Each shell had its own unique energy level n were n was a positive integer = 1, 2, 3, etc. These were called the atom's quantum numbers.



Bohr Model of the Atom

Proposed Energy Levels

$E_1 = 13.6$ eV
$E_2 = 3.4$ eV
$E_3 = 1.5$ eV
$E_4 = 0.85$ eV
$E_5 = 0.54$ eV



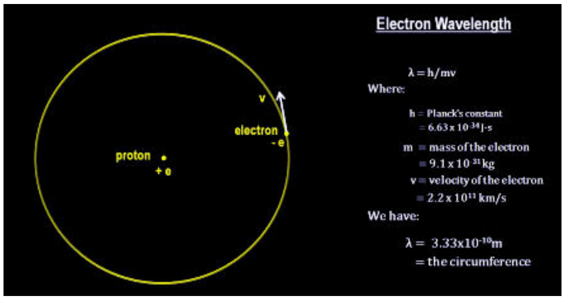
Electrons radiated energy when they changed energy levels and the emitted or absorbed light radiation had the energy difference between the levels. This energy was equal to Planck's Constant times the frequency of the emitted light.

Although Bohr's model explained the Balmer series for Hydrogen spectra, there was no explanation as to why the rapidly orbiting electrons didn't radiate, and there was no explanation as to why the distances from the nucleus were as described.

The de Broglie Atom

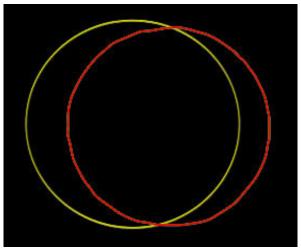
In 1925, Louis de Broglie came up with the answer these two objections. Leveraging the recently developed light particle-wave duality demonstrations, he proposed that electrons may have the same wave-particle properties.

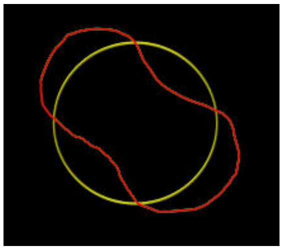
Earlier, we calculated the velocity of the electron, so like we did for electron microscopes in the previous segment, we can now calculate its wavelength. We get exactly the length of the electron orbit's circumference as enumerated by Bohr! In other words, the wavelength of the electron is exactly the length of one revolution. This would create a standing wave!

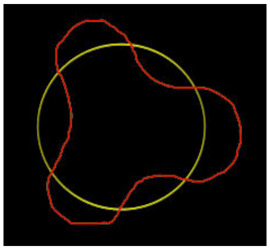


A standing wave is a wave constrained to vibrate in a distance that's an exact multiple of its wavelength. Anything more or less would create destructive interference and the wave would collapse.

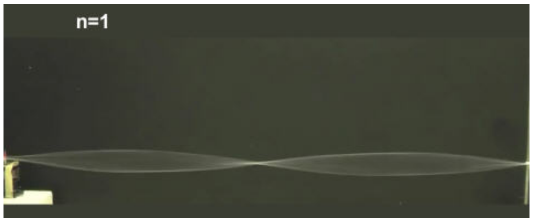
So the first energy shell would have to have the radius that creates the circumference that exactly fits one wave. The second shell would have to have the radius that creates the circumference that exactly fits two wavelengths. The third shell would have to have the radius that creates the circumference that exactly fits three wavelengths, and so on.

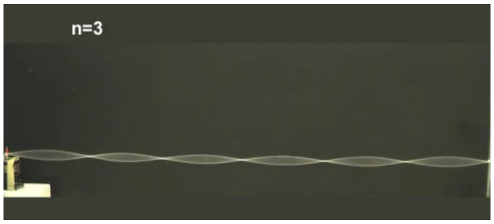




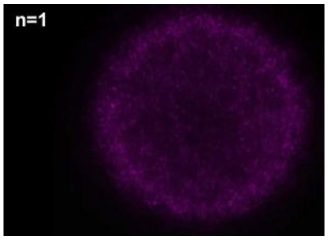


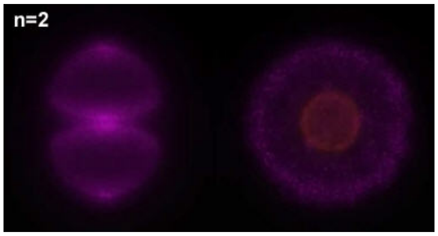
Here are a couple of standing waves on a string.

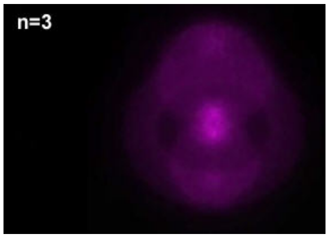




Here's what three dimensional standing waves look like around a hydrogen nucleus. It is these interesting geometries that give atoms their chemical properties.







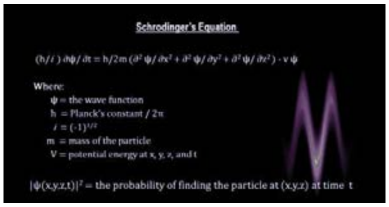
So the proposed answer to the question "How can an electron sit way outside the nucleus without orbiting?" is that the electrons exist as standing waves that envelop the nucleus. No orbital motion is required and therefore, no radiation is emitted.

Remember that: electrons in an atom do not 'orbit' the nucleus like planets around the sun – they exist as stationary standing waves.

[This 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, it follows that a generalized wave equation was needed to mathematically describe them. Erwin Schrodinger did just that.



Schrodinger's Equation

$(\hbar / i)\, \partial\psi / \partial t = \hbar/2m\, (\partial^2 \psi / \partial x^2 + \partial^2 \psi / \partial y^2 + \partial^2 \psi / \partial z^2) - V\psi$

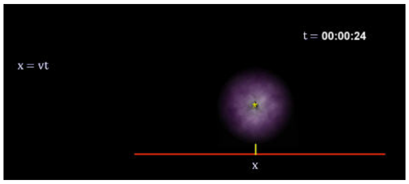
Where:

ψ = the wave function
$\hbar$ = Planck's constant / 2π
$i = (-1)^{1/2}$
m = mass of the particle
V = potential energy at x, y, z, and t

$|\psi(x,y,z,t)|^2$ = the probability of finding the particle at (x,y,z) at time t

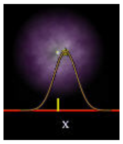
The rub with Schrodinger's equation is that the function that describes matter waves does not represent an observable physical quantity. For example, there aren't any electric or magnetic fields involved that can be measured. The value of the wave function associated with a moving object at a particular point and time is related to the likelihood of finding the body at that time. To be more precise, 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.

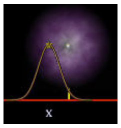


But, because the particle moves as a matter wave, we need to use Schrodinger's equations. So, when you touch the wave at time t, it collapses into a particle. Where classical physics says it is here, quantum mechanics says that 'here' is the most likely place.

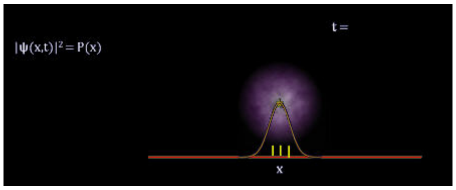
But there is a smaller probability that it is here.



Or an even smaller probability that it is here.



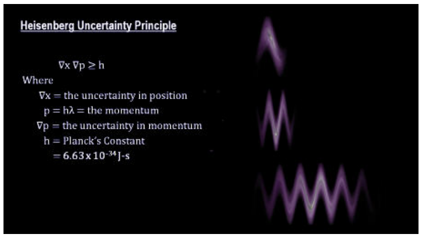
In fact, there is a chance that it may be anywhere along this probability curve, with the probabilities dropping rapidly as we move away from the most probable point.



Heisenberg Uncertainty Principle

> ***[Music:** Dave Porter – "Breaking Bad" – This is the opening theme for the Breaking Bad television series. The star character, Mr. White, chose Heisenberg for his cover name. This is the physicist who came up with the Uncertainty Principle.**]***

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.



Heisenberg Uncertainty Principle

$\nabla x \, \nabla p \geq h$

Where

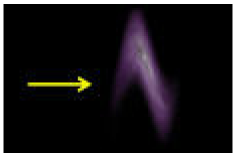
∇x = the uncertainty in position

$p = h\lambda$ = the momentum

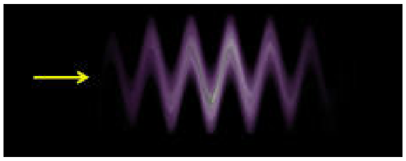
∇p = the uncertainty in momentum

h = Planck's Constant

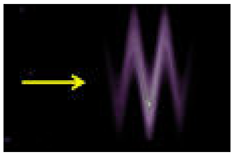
$= 6.63 \times 10^{-34}$ J-s



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.



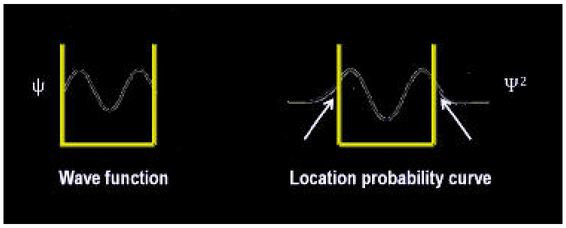
The bottom wave packet contains plenty of wavelength information, but it is quite spread out and its location is more uncertain.



The wave packet in the middle has enough wavelength information 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.

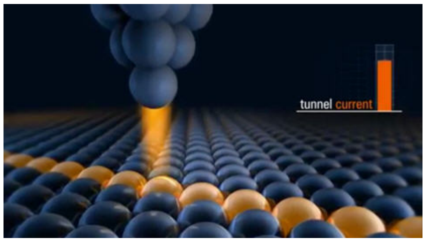
A good way to illustrate this is to look at an electron in an energy well too deep for it to get out. But remembering that the electron has a wave function that gives the probability of finding it at any given point, and some of these points (admittedly with very very low probability) can be found outside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



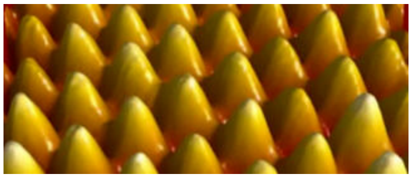
Scanning Tunneling Electron Microscope (STM)

> *[**Music:** Georges Bizet – "L'Arlésienne" – This is the Overture for Alphonse Daudet's play L'Arlésienne (usually translated as 'The Girl from Arles'). was composed for the first performance of the play on 1 October 1872.]*

This tunneling effect is the basis for Scanning Tunneling Microscopes or STM for short. 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.

[With a little stronger pull, we can even dislodge and move atoms. Here we see that the folks 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.]

Hydrogen Atom

> *[**Music:** Gustav Mahler – "Symphony No. 5" – Composed in 1901 and 1902, its most distinctive feature is the frequently performed Adagietto. The musical canvas and emotional scope of the work are huge.]*

Now let's get back to the atom.

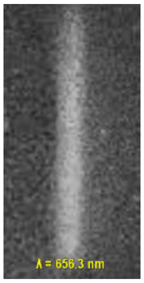
The Schrodinger equation solution for a negatively charged electron trapped by a positively charged nucleus identifies three quantum numbers for three quantities that are quantized in an atom: energy, angular momentum and magnetic moment:

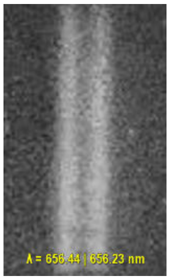
- For Energy, we get the principal quantum number *n*. It describes the electron shell.
- For angular quantum we get the orbital angular momentum quantum number ℓ. It describes the sub-shells within each shell called out by n.
- For magnetic moment, we get the magnetic quantum number m_ℓ. It describes the specific orbital within each sub-shell.

Spin and the Pauli Exclusion Principle

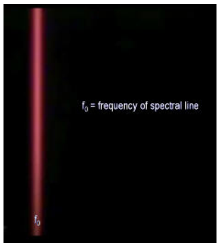
Although Schrodinger's equation went a lot further than Bohr and de Broglie, there were still a couple of things about the atom that were not completely explained.

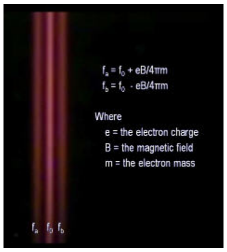
1. When examined very closely, many spectral lines showed up as pairs instead of single lines as called for by Schrodinger's equation.





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

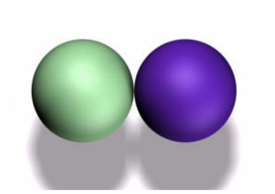




3. It was not understood why all the electrons didn't all move to the innermost lowest energy orbital.

In 1925, 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.



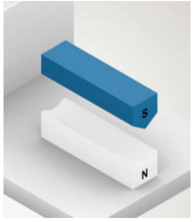
Pauli's exclusion principle stated that no two particles could occupy the same quantum state at the same time. But Pauli could find no physical explanation for the fourth quantum number.

Electron Spin

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.

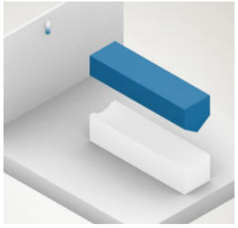
Here we use magnets and electrons directly.

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.

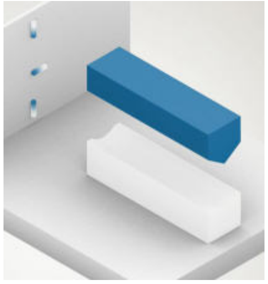


A magnet is sent through with the north pole up and the south pole down.

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



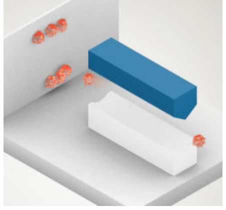
As we change the orientation of the magnets being sent through, we see the change in the amount and direction of the deflections. The deflections depend on the orientation. This is as expected.

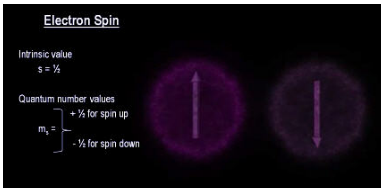


When we send large numbers of randomly oriented magnets through the field, they arrive anywhere vertically.

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

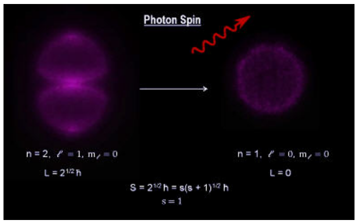
But they always arrive at the screen deflected either up or down. Never in between like the magnets.





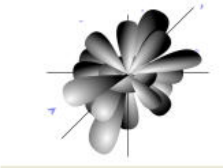
Each electron behaves as a magnet, but with only one of two possible orientations: up or down. This intrinsic property of an electron is called ‘spin’.

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.

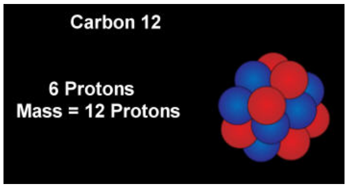
With the Pauli Exclusion Principle and spin as the fourth quantum number, the full set of spectral lines, orbitals, their geometries, and interactions with each other fell into place.



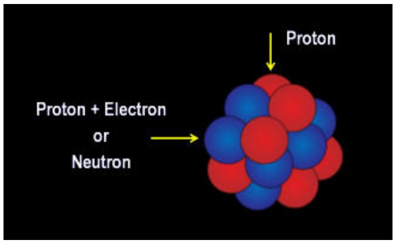
The Nucleus

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.

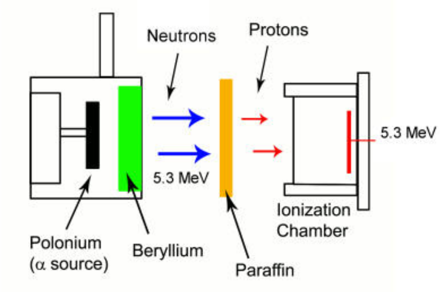
But mass spectrometers showed that atoms have more mass than the number of protons alone could account for. For example, carbon has 6 protons and 6 electrons, but its mass is just a tad more than the mass of 12 protons.



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.



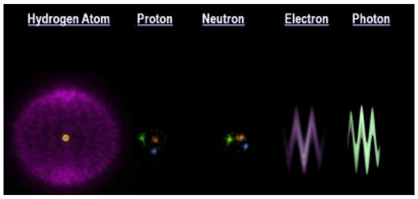
In 1932, Chadwick performed an experiment verifying his suggestion. He ran alpha particles into Beryllium that produced a radiation he guided into paraffin and then collected the resulting protons.



The key was working energy considerations back from the protons ejected from the paraffin on the right to the energy required in 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

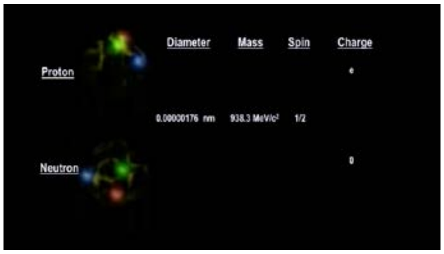
In this segment, we developed the basic quantum mechanics for electrons around the atom and measured the size of atomic components. At the end of our previous segment, we used a Scanning Electron Microscope to see carbon atoms 14 hundredths of a nanometer in diameter.



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.

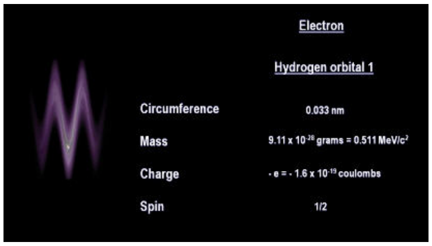
Also in this segment we added spin as an intrinsic property of particles to go along with mass and electric charge. Protons and Neutrons both display the same spin properties as electrons when they traverse the Stern-Gerlach apparatus, so their spin is ½.

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.



	Diameter	Mass	Spin	Charge
Proton				e
	0.00000176 nm	938.3 MeV/c²	1/2	
Neutron				0

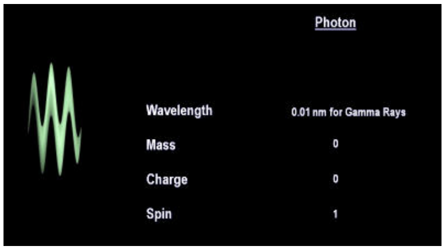
For the electron, it's hard to talk about its size because their wave packet is different for every circumstance from standing waves in thin atom shells, to scattered waves in an electron microscope. What we did in this segment was to calculate its length around the Hydrogen nucleus at .0033 nm.



	Electron Hydrogen orbital 1
Circumference	0.033 nm
Mass	9.11×10^{-28} grams = 0.511 MeV/c^2
Charge	$- e = - 1.6 \times 10^{-19}$ coulombs
Spin	1/2

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.



	Photon
Wavelength	0.01 nm for Gamma Rays
Mass	0
Charge	0
Spin	1

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.