

MITOCW | MIT3_091F18_lec05_300k

So we're going to talk a little bit about Bohr.

And then we're going to talk about ionization energy, which is a theme that's going to come back a couple more times.

Very important.

And then we're going quantum.

We're going quantum.

And that's why this is a very exciting day for me.

And I'm hoping it will be for you, too.

It's why I have t-shirts.

It's also a goody bag day.

There's just all sorts of things happening today, so I'm very excited.

This is how we ended our lecture on Wednesday.

I was showing you this as part of my why this matters, which started with refrigerators and then we talked about a chemistry problem, CFCs.

And we talked about how those were destroying the ozone layer of this planet.

Now, why is that important?

And I had this picture up there to help explain it.

Just to make sure we all see what this is, I'm showing you here, again, this is the energy-- watts per meter squared-- that we get from the sun.

And you can measure that up on top of the atmosphere, so there is no atmosphere there.

And this is what it looks like.

That's the energy as a function of the wavelength or frequency, or energy of the photon.

You now know how to go back and forth between those.

But see, on the planet here, once you have the atmosphere, you get a different spectrum.

And the reason is-- and it's shown here-- that some of the molecules in the atmosphere are absorbing parts of that spectrum.

So what we see on the planet is different.

And we talked about this in terms of why this matters and how ozone-- so I mentioned, oh, if we can only see out here-- the visible is here.

That's the visible.

There's the frequencies of the visible.

But what if it were here?

The planet would be totally dark, because no light would get through the absorption that happens from those molecules.

On the other hand, over here, we've got ozone that's protecting us.

It is absorbing light in the higher frequency, which means lower wavelength.

That's the UV.

And that's why that was such an important problem is to not destroy those precious ozone molecules.

So I wanted to make sure that we-- and we understood this in terms of energy transition, because how do these things absorb light?

Well, they do just what we said, which is they take the energy from a kind of light and an electron then takes that energy, absorbs it to be excited, and it goes from one energy level to another.

And we did that in terms of the Bohr levels, so I wanted to do that.

And that is what you have right here.

You have a way to see that visible spectrum.

So we'll talk about that as well.

Let's talk about it right now, because have a slide on it.

So you have what is known as a spectroscope, or the most powerful way to see light ever designed.

And if you look through this-- OK, don't look at the UV please.

Don't look at the sun.

Please don't do that.

But we gave you light sources to look at.

Like that, right?

Or you can look at that.

Or you can look at things around you with it.

And what are you seeing?

You're seeing this.

You're seeing the spectrum in the visible.

You're seeing it in this part.

But from this, you are also-- what you are seeing are literally electron transitions.

You are literally seeing-- and that's where I want you to think about.

That's why I wanted you to have this power, this power, which means responsibility, as we know.

But now, you can not only see light.

You can see electron transitions.

I want you to think about that as you're seeing light.

And I want you to think about it in terms of what we learned on Wednesday, which is Bohr.

And so let's just put that here.

So we had Bohr.

And in the Bohr model, we said the energy is going to be quantized minus 13.6 electron volts times the atomic number squared divided by n squared.

And so for hydrogen, for example, this is equal to minus 13.6 eV, because z is 1 over n squared.

So that was for hydrogen. z equals 1.

And so that's it.

And that's it.

n is equal to 1, 2, 3, et cetera.

It's quantized.

This was the whole kind of discussion and theme of Wednesday's lecture.

And that got us to these levels, these Bohr levels, which are places where-- energetic places, energies of the electron.

So we'd have these levels that kept on going until you're at somehow-- so this is energy.

And this would be like n equals 1, n equals 2.

These are the energies right there.

I'm just plugging in.

And this would be n equals infinity, so the energy here would be 0.

And for the hydrogen atom, the energy here would be minus 13.6 electron volts.

That's hydrogen. So this is-- now, I'm leaving this here and I want to use this.

That's why I'm putting it back up.

Because look, now you can use this to start understanding light.

And so you use this, and you look at light sources, and you can start thinking about things in terms of energy transition.

So let's take a look.

So what if I just look at some light sources, all right?

OK, so that's what you see.

That's the sun, all right?

So that's daylight.

And now we're getting into the mode of seeing light like this, all right?

It's an intensity, salute versus wavelength, right?

And there you go.

And now if you look at sort of an incandescent light bulb, it will look something like this.

So it's warm.

There's a little more red, all right?

You can't get up to the sun's temperature, but you can it pretty hot.

So you see this nice kind of looking spectrum.

And, oh, here's a halogen. All right, that looks kind of closer maybe.

Look at what happens when we go fluorescent.

Look at that.

Why does fluorescent light give us headaches?

Well, that might be other reasons.

But look at this spectrum.

That doesn't look like that.

That's why we all have so much trouble with fluor-- and here's LEDs.

And you can buy your cool white LEDs and your warm white LEDs.

But look, they're still not.

They're not that, right?

There's been a lot of work on this.

Can we get this to be that, all right?

And so now we are armed with great power.

And so you can look at things.

You can say, oh, I see that it has this large peak because I'm looking at a fluorescent light within it.

And I see that it's got this stronger peak in the blue.

Maybe I could just absorb a little of that.

What if I came up with a coating, all right, that I could put on that cool LED bulb to make it just a little bit more look like that?

Well, now we know how to we might be able to use a board to do this, all right?

Because what if-- so let's say, OK, the wavelength of that is around 450.

So if the wavelength is 450 nanometers, then I know-- remember we did this-- waves, light waves, we can go back and forth between frequency-- hello.

How's it going?

All right, welcome to class.

And energy, energy equals hc over λ .

Oh, Planck, thank you.

And this is roughly 2.75 electron volts for that wavelength 2.75 electron volts, right?

OK, well, but see now you say, but I want to absorb some of those photons.

I'll pose a problem.

What if I have an atom-- so question, right?

So I have an atom, an atom that has an $n = 4$ to $n = 5$ transition that absorbs.

So I'm telling you now this energy.

So I'm telling you now I've got some atom, and I know there's a transition in it.

It has an electron transition.

Because remember, we're going from electron transitions to light energy, back and forth, back and forth.

Look at this right here.

I'm just saying, look, OK, $n = 3$, $n = 4$, $n = 5$.

So what I just asked you is, I said.

Well look, some atom has a transition between $n = 4$ and $n = 5$.

That means that if an electron happens to be here, it can absorb energy to promote it up to here of exactly 2.75 eV.

Uh-huh, well, I can use Bohr.

I can use Bohr now, right?

Because so I can actually answer this problem, answer this question.

If I know that, then I can say, well, OK, if the initial state is 4, and the final state is 5, right-- I've gone from 4 to 5.

I've promoted an electron up in energy, lifting it up like a ball against gravity, all right?

Then if then I can use Bohr to say that 2.75 eV is equal to minus 13.6 z^2 times $\frac{1}{5^2} - \frac{1}{4^2}$, that would be the transition energy, this, dictated by Bohr.

And, oh, I have everything except z , which I can solve for.

$z = 3$.

$z = 3$, OK?

So now all I need to do is come up with a way to code all the cool LEDs with lithium, right?

Because this is lithium.

And we'll be fine, except there's a problem.

And this is part of the problem that we need to address and solve today before the weekend.

And that is that, see, Bohr only works for one electron.

Now, it's OK. z can be higher.

z is 3, and it works here.

But I assumed only one electron, which means this really would be like Li^{2+} .

Oh, OK, we're going to come back to that.

But Bohr only for one electron, all right?

So if I'm going to get something like z is 3, that's fine.

But it can only have one electron there.

It can only have one electron there or the model doesn't work.

It was one electron with some z of protons, all right?

So this could be Li^{2+} .

That sounds harder to put into a film, OK?

We will be coming back to this.

That is a serious limitation of the Bohr model.

And we can't do a lot if we can only do one electron, right?

You can do a lot of ions, but that's it.

OK, so I just said the word ion.

What happens now if you take an electron not from one level up in energy, but up, up, up, up, up, all

the way until it's free?

What happens if you do that?

Well, that is called ionization.

That's ionization.

And, in fact, when you look out into the beautiful world, the universe-- remember I said on Wednesday, 75% of this is hydrogen. A whole lot of it is ionized.

It's getting ionized from energy that's knocking electrons out.

And then electrons are coming back and cascading through these levels, emitting discrete frequencies of light.

And this is what you see.

It's beautiful, right?

And what you're looking at is ionized hydrogen getting electrons coming in and cascading through the Bohr levels, all right?

Ah, but what does ionized hydrogen mean, right?

OK, we can now understand it.

We can now understand it because now, instead of going from 4 to 5, I'm going from anything to infinity.

It's like escape velocity, right?

It's like, how much velocity do I need to give something?

So it just gets away from the earth.

And it's not going to feel the earth's gravitational pull.

Well, here, I don't want this electron to feel the potential of the atom anymore, the z , the positive charge.

I want it to just be free.

That's here.

That's infinity.

That's n equals infinity, all right?

So if I think about this, that's actually quite easy to put into Bohr, ionization energy.

So to save time, we'll write IE, right?

And this is equal to the energy required-- oh, I'm going to save time here.

Look at that-- req'd to remove the electron from the atom.

OK, good.

And if we think about this in terms of Bohr, it would be minus 13.6-- let's say for hydrogen, because I've got hydrogen up there.

So for hydrogen, meaning z equals 1, it would be minus 13.6 times 1 squared-- that's the z squared, all right-- times 1 over n_f squared minus 1 over n_i squared.

That's that Δe , which equals minus 13.6 times 1 over infinity minus 1 over n_i squared, all right?

And the last I checked with my colleagues in the math department, that's 0.

That's 0, all right?

So the ionization energy, you literally can just read off here.

If an electron is in the ground state, if there's a happy electron right here in this orbital, the amount of energy I need to input into that atom to knock the electron fully out, the ionization energy, is 13.6 electron volts, right?

It's just the energy of that orbital.

I can get a little fancier and say, well, OK, what about this question.

What if I had light with 100 nanometers?

So now I'm telling you my source, all right?

Before I didn't tell you.

Say something goes into it and knocks it all out?

No, now I know what's going in.

The wavelength, or the energy of the photons, is 100 nanometers.

And I'm shining that on hydrogen.

Well, now, OK, let's see here.

So a little example.

Oh, I'm asking a question.

What's the question?

What is the lowest energy level that the electron could be in such that a photon from this light source ionizes the atom?

OK, let's see, because I now know the pieces, and I know the definition, all right?

So now if I have the energy of 100 nanometer photon, well, that-- I'm not going to go through it-- is equal to 12.4 eV.

Uh-huh.

But see, if I now want to use this to ionize hydrogen, then I get something like this.

I get the ΔE equals minus 13.6 times $1/n_i^2$ squared-- OK, fine, we'll keep that for a sec-- minus $1/n_f^2$ squared, right?

This is how much.

This is how much ΔE I need to ionize from any given place, any given energy level n_i , all right?

OK?

But I have 12.

So what if I set this equal to 12, all right?

If I set this equal to 12.4, that's what I have.

Then what I get is n_i equals 1.05.

OK, good, that's my answer.

No.

Panic.

Everybody should be panicking.

No.

That can't be my answer.

I can't have-- I can't have something that's not an integer, because I said it over there that n is a quantum number.

1, 2, 3, nothing is allowed in between.

Mm-mm.

So what you can see is that I got so close.

I almost had enough energy to get n to be 1.

But I don't.

So I cannot use this light to ionize an electron in n equals 1.

I can't.

But I could use it to ionize an electron in n equals 2, right?

Because there, if n is 2, well, then I've got actually plenty of energy, all right?

So cannot ionize from n_i equals 1, but n_i equals 2 is OK, all right?

So that's the outcome of this.

Now, I could also use that same 12.4 eV photon to ionize from n equals 3, or 4, or any of these.

But I can't get to 1.

And there's one other concept that's important.

Because look, I gave it 12.4 eV.

What did I actually need?

Well, you can even see it here.

You know what these levels are from the Bohr model.

Here, this is 13.6.

This is minus 3.4, OK?

That's telling me this is definitely not drawn to scale.

But still, you get the point, 3.4.

And I just said, well, I could probably get it out of 3.4.

And I can because I've got 12.4.

And so if I did this-- so if n_i equals 2, then I would have excess energy.

Excess energy would be nine electron volts.

And that is kinetic, right?

So I've shot this thing out of the atom with light that's got way more energy than I needed to ionize it.

But that energy, it's still in the electron.

I transferred that photon energy to the electron energy.

And so that is going to be in the kinetic energy of the electron that's now free, the electron free.

I should have said free electron, right?

Excess kinetic energy comes after ionizing, right?

So these things are the things we can do with Bohr.

So I wanted to kind of take what we learned Wednesday and apply it to a few problems and go back and forth with this idea of photon energy to Bohr level changes.

And that's what I want you to do in this week's goody bag as well.

OK, now ionization is an important concept.

Like I said, we'll be coming back to it.

And you can actually plot-- and many people have measured, and this is very important-- the ionization energy of elements, like as a function of atomic number.

And the first ionization energy is important, right?

The first line ionization energy is the energy required to remove the outermost electron, right?

All right, so if I did have a whole bunch of electrons, so if I had a whole bunch of electrons.

the first ionization, the first ionization energy is the energy required to remove the outermost one.

So that's an important number to know.

If you want to take an electron off of an atom-- and will be doing that because that's how we're going to make our first bond next week-- if you want to take an electron off an atom, this is one of the first things you ask.

How much energy does it take, all right?

How much energy does it take?

So that's an important concept that we'll be coming back to.

Now, OK, I alluded to the fact that we needed a new theory.

I already alluded to that.

And I said that's where we're going today.

The Bohr model works for one electron, as I said.

All right, that means it's good for hydrogen. Oh, we just ionized hydrogen, and that was a lot of fun.

And it's good for He plus, or Li2 plus, or osmium 75 plus.

Yeah, that's real.

But what about two electrons?

What about two, just two?

Forget about 10, 20, 30.

2, the Bohr model can't do it.

It doesn't.

It doesn't work for anything more than one electron.

It was derived for one electron, and that's where it stays.

So we're at this point now.

This is the picture I started to draw this week on Monday.

And we're at a point where we need something deeper.

And some of the work that went into this was really critical.

And I'm not going to go into too much detail, but there's one thing I want you to know.

So remember, Planck and Einstein, they were playing with quantization of light.

And Einstein's photoelectric effect found the particle nature of light, with Planck's-- you know, with together, their energy, equals $h\nu$, that quantized energy of the photon.

And they saw it was a particle, but it also was a wave, all right?

There's a lot of discussion, and interest, and attempt to understand what it was.

Was it a wave or a particle?

And Compton did some beautiful experiments to show, again, evidence that it was a particle because he saw how light interacted with electrons.

And it was like this collision that particles do.

And then De Broglie, or as [INAUDIBLE]..

De Bro-- I don't know-- would say he went further.

De Broglie went further.

He said, everything is both.

Light, you guys don't know if light is a particle or a wave.

Well, I'm saying everything is a wave, everything.

And everything has this duality.

So he wrote something very important, which is the relationship between anything's momentum and its wavelength.

So we'll come back to this.

So I want to write it down.

So De Brog-- he said that the wavelength of anything can be written as, oh, that same quantization number, constant from Planck divided by its momentum, literally, just its momentum.

Anything has a wave nature, anything.

And light has a particle nature.

What's going on?

And it led Einstein to say at the time-- he said, quote, "It seems as though we must use sometimes the one theory and sometimes the other while at times we may use either.

We are faced with a new kind of difficulty.

We have two contradictory pictures of reality.

Separately, neither of them fully explains the phenomena of light.

But together, they do.

This was the beginning of quantum mechanics.

This is where they were.

And later, once they really got the quantum mechanics kind of going, he said this.

The more success the quantum mechanics has, the sillier it looks.

Why?

Well, there's only one way to really learn about quantum, especially if it's the first time.

And that's from Dr. Quantum himself.

And so I do have a video of Dr. Quantum.

And I'm hoping you will enjoy this with me.

I enjoy it every time I see it, most Friday nights.

So let's watch Dr. Quantum explain.

Because there was one really important experiment that happened that he's going to explain right now called the double slit experiment.

And it changed everything.

And I don't care if you see it in another class.

We're seeing it here because this sets up what happens next in chemistry.

Let's watch Dr. Quantum.

--are.

The granddaddy of all quantum weirdness, the infamous double slit experiment.

To understand this experiment, we first need to see how particles, or little balls of matter act.

If we randomly shoot a small object, say a small marble at the screen, we see a pattern on the back wall where they went through the slit and hit.

Now, if we add a second slit, we would expect to see a second band duplicated to the right.

Now, let's look at waves.

The waves hit the slit and radiate out, striking the back wall with the most intensity directly in line with the slit.

The line of brightness on the back screen shows that intensity.

This is similar to the line the marbles make.

But when we add the second slit, something different happens.

If the top of one wave meets the bottom of another wave, they cancel each other out.

So now there is an interference pattern on the back wall.

Places where the two tops meet are the highest intensity, the bright lines.

And where they cancel, there is nothing.

So when we throw things, that is matter, through two slits, we get this, two bands of hits.

And with waves, we get an interference pattern of many bands.

Good so far.

Now, let's go quantum.

It's my favorite line.

I love that line.

An electron is a tiny, tiny bit of matter, like a tiny marble.

Let's fire a stream through one slit.

It behaves just like the marble, a single band.

So if we shoot these tiny bits through two slits, we should get, like the marbles, two bands.

What?

An interference pattern.

We fired electrons, tiny bits of matter through.

But we get a pattern like waves, not like little marbles.

How?

How could pieces of matter create an interference pattern like a wave?

It doesn't make sense.

But physicists are clever.

They thought, maybe those little balls are bouncing off each other and creating that pattern.

So they decide to shoot electrons through one at a time.

There is no way they could interfere with each other.

But after an hour of this, the same interference pattern seemed to emerge.

The conclusion is inescapable.

The single electron leaves as a particle, becomes a wave of potentials, goes through both slits, and interferes with itself to hit the wall like a particle.

But mathematically, it's even stranger.

It goes through both slits, and it goes through either, and it goes through just once, and it goes through just the other.

All of these possibilities are in superposition with each other.

But physicists were completely baffled by this.

So they decided to peek and see which slit it actually goes through.

They put a measuring device by one slit to see which one it went through, and let it fly.

But the quantum world is far more mysterious than they could have imagined.

When they observed, the electron went back to behaving like a little marble.

It produced a pattern of two bands, not an interference pattern of many.

The very act of measuring, or observing which slit it went through meant it only went through one, not both.

The electron decided to act differently as though it was aware it was being watched.

And it was here that physicists stepped forever into the strange neverworld of quantum events.

What is matter, marbles or waves?

And waves of what?

And what does an observer have to do with any of this?

The observer collapsed the wave function simply by observing.

You can see why I show this.

We do movie night at home a lot.

It's Friday night, guys.

Seriously, you know that you might have a little free time.

You get together with some friends.

There's no better way.

But look, this is why-- I hope that's as mindblowing to you as it is to me.

It was mindblowing to them at the time.

This led Bohr to famously say, anyone who is not shocked by quantum theory has not understood it.

And here's the thing.

There are some things in life where you can ask questions, and you can understand them.

You can answer them.

I can ask the question why do-- or somebody might say, why do I wear a t-shirt under a sportcoat?

That's weird, stylistically.

I mean, stylistically, that's-- and I say, well, OK, here's the answer.

Because I put comfort over stylistic norms.

Whatever, that's an answer.

I could ask a harder question.

I could ask a harder question.

Why did Justin and Selena break up?

Because Jalina was totally hashtag couple goals.

Just saying.

OK, anyway, now I'm just-- why?

Oh, but still that's a harder question.

Why?

Because they were so good.

But there's probably an answer to that too.

That's a harder question.

This was mindblowing.

How can something be both?

How can something that they know is a part of-- How can everything be both?

This was really difficult to answer.

And, in fact, the real answer came with a mathematical solution, which we'll learn next, but not quite understanding.

We still don't really understand quantum mechanics, even after 100 years.

We don't.

And so this is just, I think, a beautiful part of our story, of our detective story.

And just to show, people are still working on this, right?

How much quantum can we see?

How much quantum can we see?

Well, look, here's somebody who, about 15 years ago-- and it's published in this paper-- and a little more, almost 20 now.

They actually took-- forget about electrons.

Those are teeny, tiny things.

This is a whole group of 60 carbon atoms.

That's called the fullerene.

And they shot this through-- they did the double slit experiment with this.

And they showed that with the grating-- wait, without, you get the signal peak, just like we saw.

And with you get interference.

You get interference.

So they were able to show all the way up to a large molecule, 60 carbon atoms, that these things are still-- they have wavelength.

They have these incredibly strange properties that you just saw in the double slit experiment.

All, everything has this duality, this duality between waves and particles.

Now, OK, so we needed a new model, right?

Bohr can only do one electron.

And meanwhile, there's all this stuff going on at the time showing that you need a new theory.

How do we explain waves?

And to kind of put a little bit more of the nail in the coffin for Bohr, you had Heisenberg.

So I'm just going to write this on the board before I go into why this matters.

And what Heisenberg said is that you cannot measure both position and momentum exactly.

You can't get them both, OK?

And that's called the Heisenberg uncertainty principle.

But that was it, because-- look, you look at that and you say, well, he-- saying this.

But in Bohr, you got them both.

You've got them both exactly.

Remember, I put energy down there.

But we also put r on Wednesday, exactly.

Gesundheit.

You can't get them both.

And if you can't get them both, that's another problem with Bohr because Bohr says you can get them exactly, right?

So we had these pointers, double slit experiment, Heisenberg's theories, all right, saying we need something more.

We need something more.

And I'll get to that.

I want to do my why this matters.

And then I'll get to the more that came, which we'll then expand on Monday.

Before I do that, let's do why this matters.

So why is the fact that an electron is a wave, why is that so important?

Well, first of all, because it sets up this detective story of the electron set up this new theory that's coming, quantum mechanics, right?

But also, it immediately, just like we started painting with them, right-- remember that was one of my why this matters-- we also realized that we could see with them.

We could see with electrons.

Because if electrons are waves, then I can shine them just like I shine light and see what it shows me, right?

It can illuminate matter.

So if you look at the frequency here of light, this is an electromagnetic spectrum-- radio, microwave, infrared, visible, UV.

Now, here's the thing.

If you want to see something, some feature size, you're limited by the wavelength of the light.

It can't be bigger than the features you're looking at, roughlyish, OK?

That's what you'll-- so if you're trying to see something tiny, but the wavelength of light is really big, you won't see it.

So we need-- let's say we want to see atoms.

Let's say we want to see atoms, or even more, nuclei.

Look at how tiny.

Those are 10^{-10} to the minus 10^{-12} meters.

Those are x-rays or gamma rays.

But the problem is, if we shine x-rays on things-- and we will do that when we look at crystals-- but if we shine x-rays, it's very hard to then collect them and make a photographic image, OK, at least one that gets you that resolution.

And gamma rays are even harder to catch, all right?

But see, here's the thing.

The electrons give you exactly what you need.

Because if we do this math for an electron-- so bring this one back down-- if we do this math for an electron, well, I'm going to use-- oh, I thought I had the middle one.

So if I have an electron-- let's suppose I have an electron that is-- electron, I'm going to say I accelerate it over 100 volts.

I'm going to take an electron, and I'm going to put it over 100 volts.

I'm going to give it some kinetic energy, right?

So its kinetic energy is then going to equal 100 eV, right?

OK, so now I've got an electron moving with a kinetic energy that's 100 eV.

Now you can convert this to joules, and you can set this equal to $\frac{1}{2} mv^2$, right, mass of the electron times its velocity squared.

And then once you have the velocity, so you get the V .

And then once you have that, you get the momentum, the p , right?

And then once you have that, you get the wavelength, right?

So I can go now from something that's easy to do.

100 volts is a lot, but in a lab safe, not in your dorm.

You could apply 100 volts to an electron, get it going at this speed.

And once you know the speed, you know the momentum.

And if you know mv , then you know its wavelength.

In that case from this relationship, you would get that it's about 0.12 nanometers.

But look at this.

The wavelength of a simply accelerated electron is right where I need it.

It's right where I need it.

It's an Angstrom, right?

So now, if I take advantage of the wavelengths of the wave nature of electrons and I shine them on materials, then I can see materials that way.

And I can see them at that scale.

And we do that all the time, all the time in many, many different areas of technology and research today.

We use electrons to image.

In fact, the best images you can get are made with electrons, all right?

Here's an example of using what's called a scanning electron microscope, all right?

So you see a butterfly, but you want to really see a butterfly, or we can go even further.

And instead of just drawing pictures of this these beautiful materials made of carbon-- those are called nanotubes.

This is called graphene.

Gesundheit.

It's a single sheet of carbon atoms, one atom thick material.

Notice with these materials every single atom is a surface atom.

That's pretty cool.

They also have lots of other cool properties.

And I'll give you examples throughout other lectures of how these kinds of materials can be used.

But for now, I'm talking about seeing them.

And this is what happens when you actually look at them with an electron.

That's a nanotube.

And here is a picture of graphene.

The only reason we can see graphene is because we have electrons.

And we take advantage of the wave nature of those electrons.

Well, you say, well, OK.

But why does that matter?

Well, that matters tremendously, because one of the first experiments that really did what Feynman, what Richard Feynman wanted-- Richard Feynman predicted the field of nanotechnology 50 years ago.

He gave a famous speech at Caltech called There's Plenty of Room at the Bottom.

He's also an amazing teacher, and he taught actually the double slit experiment.

I highly recommend you Googling that lecture.

And he said that someday, you can put the atom where you want, all right?

And the first time that was done was in 1989 by IBM.

They had 35 xenon atoms.

They moved them around.

But the point is, you couldn't realize nanotechnology.

You couldn't realize the ability to actually move atoms if you can't see them, right?

And this, the ability to see what you were doing changed everything.

It changed everything.

And nature had been doing this.

And I love these examples.

So nature has been doing nanotech for a long time, all right?

So you have the inner ear of the frog.

It's a cantilever that is sensitive to a few nanometers of movement.

The frog can actually hear that.

You've got features in the ant's eye.

I love the silk moth.

The male silk moth has a single molecule detection system onboard that can sense a single molecule pheromone.

It can detect a female silk moth two miles away, two miles away.

We have nothing like that.

We have no technologies like that.

I can't even tell if someone's in the next room.

I have to look at my phone or something.

This is because of nanotechnology, that kind of detection system, all right?

But it was the ability to see atoms and molecules with electrons that kind of blew open this entire field, all right?

And it's made it so that we can now try at least to rival nature.

Here is one example.

You're not just seeing graphene, but check this out.

You're seeing a single atom of graphene, and you're seeing what happens under a certain kind of irradiation.

And you're seeing this hole.

And you're seeing the hole grow.

And that's really important, because something I care a lot about are membranes, right?

And another one of these matters, I'll tell you about membranes.

But here, I'm actually making the thinnest possible membrane that you could make because it's only one atom thick.

And I'm controlling how I make that.

But I would never know what my controls do if I couldn't see it in real time, all right?

So this is my why this matters.

It's seeing things at this scale.

OK, so back to-- so there's Heisenberg, right?

So in every moment, the electron is only inaccurate position and an accurate velocity.

This was, again, saying Bohr, you got to think harder.

And we had all of these experiments showing the particle wave duality of nature and of atoms.

And so the question then became-- all these pointers from all these different places, the question became, how do we describe this wave nature of matter?

How do we describe an electron as a wave?

And I like this, because not only do I think if Schrodinger had a swimming pool this is what he would do with it, but also I thought maybe this would be cool for an MIT kind of swim test thing in the future.

Because you can see what happens.

This, I want to do this experiment.

But this looks really fun.

They're making waves.

They're literally making waves, right?

And so the question now-- look at that.

See, but the question is, how do you describe it?

Well, there are wave equations.

There are wave equations that we know about that relate the position of a wave to its time change in velocity, all right?

And the thing is, this is it.

This is how we need to describe an electron.

Now, think about it for a second.

An electron-- by the way, it just occurred to me that you should take this on that same day.

Because if they don't have real candles, and they only have these, you should have this handy because you might need to pull it out and see what kind of LED they gave you.

What does it look like?

Anyway, OK, that was a side point.

Bohr says electrons can only be in certain places.

Well, that's mindblowing because-- by the way, think about that.

An electron can only be here or here.

Oh, and by the way, it can transition between them, but it never be between them.

Did anyone catch that?

What's up with that?

It can't be between them.

I am telling you that.

But it can transition.

Where did it go?

Oh, OK, that's a weekend kind of exercise.

And now we're saying they're not even in one place.

They're waves.

They're like the water.

That's cra-- that really was at the time-- they thought how is this possible, right?

But Schrodinger came along.

And he said, I'm going to write down the equation anyway.

And that was the big contribution, the really critical contribution that Schrodinger did, is he said, I'm going to write down kind of like a Newton's equation, like $mv^2 = ma$, but for quantum mechanics and for waves, like electrons.

And I'm going to describe them.

So he developed the equation that we saw, that on Monday we'll solve, to describe the wave nature of quantum mechanics.

And just to end-- so on Monday we'll pick up with Schrodinger's equation for the atom.

The question is, what does this all mean for chemistry?

What does this mean?

We're not-- I'm not teaching you quantum.

I'm not teaching you quantum just so you know quantum in isolation.

I'm telling you about it because it completely revolutionized chemistry.

And that is where we will answer that question on Monday.

Have a really good weekend.

[SIDE CONVERSATIONS]