[SQUEAKING]

[RUSTLING]

[CLICKING]

PROFESSOR: Today, we're moving on, and we're moving into what's really happening with those electrons. And I framed this on Monday by telling you that the Rutherford atom wasn't going to make it more than 10 to the minus 11 seconds-ish because classical E&M would tell you that the electron, if it's going around like this, is radiating energy, and so it spirals down into that positively charged nucleus.

So Rutherford got it right in the sense that you got all this positive charge concentrated in this one tiny little volume in the middle, and the electrons are way out there. That's the planetary model. But we needed some new theory to understand why the whole thing was stable, and that's where Bohr comes along, and that's what we're going to talk about today.

Now, we need to feel our oneness with waves. Now, I think probably most of about waves, but since we're going to be talking about them a whole lot today and as we go forward, I want to make sure we're all on the same page.

So here's a wave. Well, we know about water waves. They move a meter a second. Sound waves a little faster, right? And you've got these characteristics of a wave-- the frequency of the wave, the wavelength. And so I just want to make sure that we're all together on this because we're going to be talking about these variables a lot.

So if you have a wavelength, that might be in meters. How many meters do you have per wave? We're getting somewhere here. If you have the frequency-- and I'll use this notation here. You can use f. You might see f sometimes. I like using nu. And if you have the frequency, well, that's going to be the number of waves per second, for example. And so the speed is going to be the two of them multiplied together. And I know a lot of you have seen this, but again, I want to get it up here. So it's going to be meters per second if these are the units that I'm using, seconds and

meters.

Now, if you're light-- light as in light-- then you have a speed that is a constant in vacuum, and we know that. So the speed of light-- so for light, let's put that right underneath. For light, the speed is c, which equals 3 times 10 to the eighth meters per second. OK, just getting to some good stuff. So if I wanted to know, for example, how to go back and forth between frequency and wavelength for light, well, I can just use this relationship.

So let's put an example up there. So for example, if you're red light, so if you're red light, then you might have a frequency of around, let's say, 450. There's a range, but 450 terahertz would be a frequency. And so now I can get the wavelength because I have that relationship. So the wavelength for red light is c over nu, which is going to be equal to 3 times 10 to the eighth meters per second divided by 450 times 10 to the 12th per second, right? That's the measurement of hertz is 1 over seconds, right? That's frequency, 1 over seconds.

That gives me a number that seems like it's pretty small, 0.000000667 meters. And so we could use a different unit like nanometers, which is just 10 to the minus ninth. So we could call that the wavelength of red light, 667 nanometers.

Now, I'm doing this all slowly and carefully because we're going to be doing this a lot, and so I want to make sure that you're comfortable with this. And we're going to add something to this in just a minute.

If I was blue light-- where's blue light? Let's put blue light-- I'm going to leave red light here. And we'll put blue light right above it so that you can see them together. So blue might be something like, oh, a frequency of 650 terahertz, so higher frequency. And that would give you a wavelength of something like 460 nanometers. I'm putting these both here because I'll be comparing other aspects of these colors later.

Now comes back-- no. That's our oneness moment with waves and light, and now we come back to our story, our detective story of what's happening inside of atoms, and it really requires us to go back. I said Bohr had some postulates about quantization. There was work going on at the time already related to quantization.

So I want to share with you that work because it's what inspired Bohr, and then we'll do the Bohr model. But also just I want to get you thinking about what this all means.

So if you think about music, you can think about these two instruments. This might feel like-- the violin might feel like something where you can just play continuous notes, whereas here you can't. It's discrete. I can't play any notes in between. Something is discrete. It's quantized.

Well, that's pretty obvious. You look at those. You're like, OK, that makes sense. There's some things that aren't as obvious, like here. Oh, notice what's in the calendar. That's a screen.

Now a screen from far enough away might not look quantized, but then when you look at it closely, it is. It's pixels. There's nothing in between. There's nothing in between.

The thing is that at the time, people thought about the world in a certain way. There's some things where it's just obvious they're violins like light. And then they're like, wait a second. These things are acting like xylophones. What's going on?

So Planck's contribution was really-- he was one of the first kind of deep thinkers to talk about quantization, and it was quantization of light. It was quantization of light. And so what he said-- and this is together with Einstein, and I'll show you what Einstein did in a minute. So it's called the Planck-Einstein relation. And what he said is that the energy of a photon-- so Planck-- why did I just go all caps? I don't know. I don't know. There's a reason, Planck-Einstein.

They said that the energy for light-- the energy is related to the frequency through a constant. And so in particular, E equals h times the frequency.

Now, this was revolutionary. It might just look like a simple relationship. Oh, and by the way, this is a constant named Planck's constant. I'm going lowercase-- Planck's constant. And, like other constants, it has a value that doesn't change. So that value for Planck's constant is in-- I'll give it to you in two different units so we have it up here, 6.26-- no-- 6.62-- sorry-- 6.626 times 10 to the minus 34 joules-second. Or it's also equal to 4.136 times 10 to the minus 15th electronvolt-seconds. Notice the units

are in energy times time, Planck's constant.

And you can see that it must be because this is energy and this is 1 over time, the frequency. So it's got to have those units for that to work.

But this-- did I make a mistake? I'm worried. The dates are still hitting me hard. I'm still thinking about that, having nightmares. 100 years off, 90 years off. I'm good.

So what they saw is that light had a frequency, but it also had somehow this unit of energy, a unit, a discrete unit of energy that you can now calculate using Planck's constant. And that was revolutionary. That was not what people thought.

And it really all came home in what Einstein did, which was the experiments that he won the Nobel Prize for-- his deeper thinking was in relativity, but he won the Nobel Prize for this because it was absolutely essential to further this idea of quantization and show it with an experiment.

And because it was so important, I want to share that experiment because when Einstein did was really quite amazing. He shined light on a metal, and what he saw-- so this is called the photoelectric effect-- photoelectric effect. And this is where you take a piece of metal- and they knew there were electrons in there, right? That discovery had happened. So they knew there were electrons in-- and they also knew if you hit this metal with energy, the electrons could come off. I mean, after all, that's what the cathode ray tube was. It was electrons coming off of a piece of metal.

And in that experiment, you hook up a voltage, but here he shined light. And so if you take, for example, different frequencies of light and you shine them on this same piece of metal-- so let's go red and then let's go green and then let's go blue. Now, if you shine red light on the metal, nothing happened. Nothing happened. But OK, we know if you just crank up the intensity of the light, something's got to come off. You're putting all this energy-- no. Why? Because the particles, the packets, the quantized amounts of energy that that kind of light can have is determined here.

So what Einstein says is, well, OK, then these must be particles. They must have some energy associated with a particle that follows this relationship equals h nu. And notice that doesn't depend on intensity. It just depends on frequency.

So I might be hitting this metal with more and more of them, but each one only has the same amount of energy limited by that relationship. So if I want to hit those electrons with more energy, I've got to change the frequency. And so what was found in this experiment is that, well, maybe here you get nothing. So here you get nothing. How am I going to write that? OK, nothing. I just won't write it.

Here, maybe you get electrons. And here, maybe you get faster electrons because these fundamental particles of energy of light have different energies, right? So now we can calculate it. We can calculate the energy of that red light, and we calculate the energy of the blue light because I have now E equals h nu. And if I plug it in, which I won't do, you get something like 1.8 electronvolts. And if you do the same thing for blue, from the Planck-Einstein relation you get the energy is 2.7 electronvolts.

So what Einstein did in this experiment which was so crucial is he measured. He changed the frequency of light that you shine on this metal, and what he saw is that there was nothing, nothing, nothing, nothing. And then all of a sudden you literally had, if you plot the kinetic energy of the-- I'll write it out here-- ejected electron-- so that's what's being plotted there. You plot that versus the-- I said it wasn't going to use f-- versus the frequency of the light, nothing, nothing, nothing. And then all of a sudden there's a linear relationship between them. That explains it because these things were made of these particles that were called photons. So photons.

Now, while this was blowing their minds and winning prizes and all that, Bohr came along. Now remember, this is what I mean by dates. Apart from the dates, we know now there was a lot of time between those experiments, especially Dalton and Thompson. And that's where we were on Monday.

So Bohr was looking at all this. Remember, he wrote the paper where he said it is with great interest the Rutherford atom, meaning it's wrong. It needs fixing. Meanwhile, Bohr's looking over at Einstein and Planck and he's like, wait a second. Let me get some of that quantization stuff because if it works for light, what if it also works for electrons? What if it also works for electrons?

And so he's looking at this stuff going on with Planck and Einstein, and he applied it to the electron in Rutherford's atom. That's what he did. That's what Bohr did.

He said, you guys got your light thing. I'm going after the atom. And what I'm going to do is I'm going to make just actually a fairly simple postulate. Why is that up there? Let's go. I'm going to say that for me, the quantization is going to happen with the angular momentum.

So Bohr came along, and he said that the angular momentum of the electron-- so L equals mvr, right? Remember, we listed the Bohr postulates, and one of them was that Newtonian mechanics would hold. Angular momentum, mvr.

But there's something added to it. He said that that angular momentum could only have discrete values, quantization. So he said that this is quantized, and he was specific. He said it could only be some number, some integer times Planck's constant divided by 2 pi. Why not? When in doubt, divide by 2 pi. As you all know, that has to do with just making math easier.

Now, that's the angular momentum, and he said it's quantization. And n-- n is the key here because n is this thing that says that it must be some unit. This is like the unit. It's like that photon energy. It comes in units, but it comes in units of integers. So you could have 1 or you could have 2 or 3 but not 1.5. That's quantization.

So n had to be 1, 2, 3. It had to be some integer. And oh, here we go. Watch this. We're going to call it a quantum number. I should have written that in all caps.

[LAUGHTER]

Thank you for that.

It's our first quantum number, and we've only known each other for a week-- a week. Here we are. We're quantization things. Quantum number because it's a number that counts quantumness, quantization. This is how you get the quantization from this. So it's a quantum number.

Now here's the thing. I will not derive Bohr's model for you. I will tell you what it is. And if you want, you can look up the derivation. It's actually really fairly straightforward. It basically has to do with this one assumption, quantization of the angular momentum of that electron, combined with F equals ma and the electron is stable. It's not accelerating in.

And when you do that, what you get-- let's put this-- no. Let's put that back up. So what you get with Bohr's model is the following. You get three extremely important outcomes.

First, you get that the radius of that electron is quantized. Then you get that it's energy. The energy that it's allowed to have is also quantized. And finally, you get that its transitions are-- you got it. Thank you-- quantized.

So if you plug stuff in and you do the Bohr derivation for the hydrogen atom, what you get is that the radius of that electron is equal to n squared times a constant divided by the atomic number. And this constant is called the Bohr radius. It's got a value of 0.0529 nanometers, and it's also called the Bohr radius if you want. But that's its value.

And then if do the same-- now if you look at energy-- so energy, the equation for that comes out to be minus 13.6 times the atomic number squared divided by n squared in units of eV. Those are the two key results that come from Bohr's postulates and from Bohr doing this-- doing that. And we're going to talk about each one of these.

All right, so if I-- OK. So let's go one at a time. No. So let's talk about the radius for a minute. Let's talk about the radius.

So if I look at the radius and I'm a nucleus here-- so I've got like some positive charge in the middle and some nucleons and stuff, what I'm saying-- I'm sorry, some nuclei neutrons-- is that the electron can only be at certain values. This distance r can only be at these values, and that's dictated by the quantization number, the quantum number. n equals 1 would be the closest. You can see that from the formula. And n equals 2, and so on. And this is not necessarily drawn to scale.

You can see that for n equals 1 and z equals 1, the electron is simply at the distance of the constant, 0.5-ish angstroms, 0.05 nanometers, right? At n equals 2, it's 4 times that. Oh, yeah?

STUDENT: What's the z?

PROFESSOR: The z is the atomic number. And that comes in-- let's put that here. And that comes

in because it tells us how many positive charges there are, how many protons there are in the nucleus. And that comes in when you do F equals ma. That's how it comes into the Bohr equations.

And r is a distance, and n is the quantum number, and a0 is a constant. z is the atomic number in both cases.

This is mind-blowing because it tells you that that's where the electrons are in the atom orbit, in the atom, but they're only those distances. Remember, quantization means nothing else. Think about that. I'm saying that that electron can't be anywhere else. It must be this distance, or it must be this distance. So this was a big deal. People were like, what? It didn't compute. It didn't make sense. But this is what they were starting to see, that this was the only way to explain what they saw, which we'll come to when we talk about transitions.

So this would be the radius. And when you look at the energy, it's just as-- I'm going to use-- yeah, I'll put it here. So you take that equation for the energy of an electron in an atom, and you take z equals 1. If z equals 1, then the energy is quantized, and it's minus 13.6 divided by n squared electronvolts.

And if I just draw this as energy going up-- so if I just draw this, then what that means is that I've got one line down here showing kind of the energy level. So this would be minus 13.6 eV. And then I've got another line here, and this would be minus 3.4 eV. And then I've got another one here, minus-- let's see-- 1.5-ish eV. And I'm going to start to get crowded. As you can see, they're going to get closer and closer until you get to some level here, which would be like n equals infinity.

And then above that, zero. Above that is free. The electron is free. When its n is infinity, when it's gone through all the quantum numbers up to infinity, it's free.

And that is called ionization, taking it out of the atom and making it not part of an atom anymore but just out there free able to do what it wants. You've ionized it. You've taken it out. We'll come back to this concept multiple times.

These are the energy levels for an electron, and they are also quantized. They could only be this value or only that value or only that value. They cannot be anything else in between. That was revolutionary at the time.

So this would be like n equals 1. Sorry. That's not E. That's eV. n equals 1, n equals 2, and so on.

Now the third point that I made is that it also tells us about transitions. This was also absolutely essential, and this was the key piece to understanding what they saw. And so we'll get to that in a few minutes.

What do I mean by a transition? I mean that an electron is in an atom at some level, and it moves to another level. But now the point is that if it can only be in these discrete quantized levels, then when it moves from one to another, those changes in energy are also discrete values. There are only certain changes in energy that that electron can experience.

And there's even more which is that what can cause those changes or what happens as a result of those changes is the interaction with light. And we're all the way back to here, photons.

Think about it this way. If I take a ball and I have it up here, it's got some mgh in it, right? And I drop it, and now it's got some KE. More energy was in gravitation, was stored in mgh, gravitational potential energy, and then I let it go.

Well, for an electron, this is its ground state. This is the kind of happy place. So this would be like a ground state, and these would be all excited states. It would be like excited states. So that would be like lifting the electron up and letting go.

But when you let it go, it's not kinetic energy that that transfers into. It's photon energy. Photon energy gets emitted from electrons changing their levels in an atom. Light is an absolutely intricate part of this system. And not only that, but everything is quantized.

I almost need a moment, but I'm OK. I'm going to keep going.

This was our big deal, a really big deal because now I can tell you-- OK, there we go-- because now I could tell you, for example, that if that's the second orbit in hydrogen and that's the first orbit, I can tell you what frequency light is emitted if you go from n equals 2 to n equals 1 because the change in energy from n equals 2 to n equals 1 is nothing more than minus 13.6 eV times-- well, it's the final state minus the initial

state, right? The final state is 1 over 1 squared, and the initial state is 1 over 2 squared. All I've done is pull out the 13.6, but these are just-- these are just the Bohr-model energies, and I'm just taking the difference. This is equal to Ef minus Ei. It's the change in energy.

So in this case, it's minus 10.2 electronvolts and going from the second orbit, quantized orbit, down to the first. And I can now tell you because of what we did in the beginning that that would emit a photon of frequency 2.5 times 10 to the 15th hertz. That's the energy difference, and that energy goes into producing a photon-- goes into producing a photon, and we can tell you exactly what that frequency is. So these are the kinds of transitions that the Bohr model explains, and it explains energy changing from electron transitions to photons.

So what are the implications of this? Oh, there's the question that I just answered. If this is hydrogen, z equals 1.

The implications of this are groundbreaking, and it explained what people were seeing because you could be talking about angstroms already. Angstrom looked up, and he's like, well, you guys think it's a rainbow, but I'm not seeing that. I'm seeing discrete lines all over the place. And that's what he saw when he looked at the sun. Don't look at the sun. It's not a good idea, but Angstrom did.

And now he said, well, but there are all these lines, and the wavelengths that he tabulated happened to come on this sort of thousands of angstrom kind of scale. And so he liked that, and chemists love it, and that's why the angstrom is named what it is, 10 to the minus 10th meters. Chemists love it because a bond is a few angstroms. Spectroscopists like nanometers. They're like, thousands are too big. We don't like thousands. We want to work with hundreds. So spectroscopists use nanometers. Other people, chemists, whose angstroms. I say can't we all just really get along? And it's fine. Don't make spectroscopists angry, really. I did it once.

Now on Friday, you are going to get the most accurate spectroscope that has been made by humans in your goodie bag, and you'll be able to look at the fact that the world is not continuous. These frequencies, wavelengths that are coming out are not continuous. If you take a tube of H2 gas and you apply an enormous voltage to it, you will excite electrons in the H2 molecules. As they go back down, they will emit

photons, and what will you see? Only certain values.

Well, this is what people were doing. They were looking out into the universe--which, by the way, is 75% hydrogen. And they're like, really? The hydrogen economy.

[LAUGHTER]

We-- we-- they were looking now and seeing discrete lines all over the place. How to explain it? The Bohr model is what explained it. You will be able to see these discrete lines yourselves with the goodie bag on Friday.

But if you do this, this is what you might see from hydrogen. And there were a lot of people trying to understand this at the time, Balmer and Rydberg. And I'm not going to go through-- they came up with sort of partial formula. They were empirically based.

But Bohr-- it was only the Bohr model that could explain it all. It was only the Bohr model. And the reason is that all the transitions were out there, and you need the complete picture of these changes in energy between these levels. You need to know about that to understand all those transitions.

And at the time, it was a real thing. If you were like, I want to get something named after myself in science, you'd be like let me go and find a new set of transitions, which are just basically lines, right? It's that you're not seeing a continuous spectrum in some part of space. You're instead seeing discrete lines.

And look at this, right? So like Lyman-- you could go down to n equals 1 like we did there, 2 to 1, so from here to here. But you could also go from 3 down to 1. You could also go from 4, and that's a series. Then you could also transition not from to 1 but to 2 or to 3, and each one of those is a series. And then I think finally they're like, OK, no more names. After Pfund, we're done with naming. But all those series are still there. Those are the allowed transitions.

So you can also imagine that it's not just emission. So we've been talking about emission. You get these lines when an electron goes down in energy. Remember, like the ball, the mgh goes down. It gives off photons in these quantized energies,

but it can be the opposite. You could shine light. So I could get a continuous spectrum here and shine it on an atom, and that atom absorbs light literally by exciting electrons.

An electron is in some state, maybe the ground state, and some energy comes along that exactly takes it up to one of these states. But it can't sort of take you up there. It's got to take you there-- quantization. That's why you see these lines here. These are the places where that atom can absorb a photon because it exactly corresponds to a transition from one discrete level to another. So it works both ways, absorbing photon energy, promoting electrons up.

And when we do X-rays, we just ionize everything. But right now, we're staying within the potential of the atom. So we're staying within these levels, absorbing or emitting energy.

So I could answer questions like this. If you look at a question like this-- I forget. I think this might be from a couple of years ago quiz, so you might get something like this tomorrow. No, you won't. I just looked at a few faces, and no because we're talking about it today. So you wouldn't get it tomorrow because, remember, you only get quizzed on things that we did ending on Monday. It wouldn't be fair otherwise. But next week, OK.

What's the lowest-energy photon shown? Which one is it? These are photons now. We think of these as photons. Red. High wavelength, lower energy. OK, good. That was over here.

Now what about what transition give us the line at 639 nanometers? Hello, [? Humana. ?] 639 nanometers. So I want to know what electron transition did that. Well, to solve that problem-- to solve that problem, I need to do some erasing. Maybe I'll just draw it in here because what you have is minus 13.6 eV times-- I'm assuming z equals 1 because did I say it was hydrogen? I did, hydrogen. z equals 1. So z goes away, 1 squared. So minus 13.6 eV times 1 over the final state squared minus 1 over the initial quantum state squared quantum number squared.

So that's going to equal 639 nanometers, which we now know we can convert, thanks to Planck and Einstein, into an energy. We now know that, a photon energy. It's that photon, that unit, that discrete unit that does it, and this is 1.94 eV because

it comes from 639 nanometers. Back and forth, back and forth. We've got to get into that mode and into that mood. Back and forth, energy, frequency, wavelength.

And so I won't do the math here, but the answer comes into something like it's the 4 to 3. 4 to 3 is the transition that gives you that. We're not going to go-- some day. We're not going to go kind of all-- this is a fairly simple one to see. We're not going to go like n equals 528 to n equals 7,000, but some simple kind of back and forth. Just get into that mode of looking at these transitions as changes in energy, which means the emission or absorption of a photon. That's what I want you to get out of this.

And you can go further than that and answer questions like power questions. So here's a question. If your favorite radio station broadcasts at a frequency of 107.9-- I'm not saying that's mine, but you got to groove out-- and a power output of 50 kilowatts, how many photons are emitted by the transmitter each second? I just want to give you this as another example of thinking about this. Light, energy-- light to energy, photons, particles.

And here, you can get that by simply knowing what power is. And if you know what power is, it's watts per second. Sorry, watts is joules per second. And so the energy here is h nu. And I won't go through all the math, but I'll leave the answer here. You get 7.12 times 10 to the minus 26 joules per photon.

Now, how did I get that? I got that because I know the frequency. Why am I using joules? Remember, I put them both up there-- joules, eV. Ah, because I know my power, which is 50 kilowatts, which is 50,000 joules per second. And so I've got joules in the power. I just decided to stick with joules. Joules, eV, joules, eV, back and forth. They're just units. If you're not familiar with changing back and forth, you should do a little practice. And this gives you something like 7 times 10 to the 29th photons per second.

So you can think about power, photons, light, energy. If I shine that many-- if I shine that many on Einstein's medal, will the electrons come off? That seems like a lot. We don't know because we need to know how much energy it takes to liberate the electron, and then that ties into the frequency. Radio waves are much lower energy, much, much lower energy.

Now, this gets me to why this matters, and that's how I want to end today's lecture in the last five minutes because we talked about emission, electrons transitioning from their discrete orbitals as a way of emitting or absorbing. And so you if you have different atoms, there's hydrogen. These are the absorption lines-- absorption lines. Carbon has absorption lines. Oxygen, nitrogen, sulfur, so lots of things have absorption lines.

Why does this matter? Well, of course, because of the refrigerator. I love this plot. This is a plot of the refrigerator. This is the size of the refrigerator as a function of year. The date only goes to 2002, but it's roughly stable now. Does anybody know what's limiting the size of the refrigerator? It was going up and up and up a lot, and then it stopped. Does anybody know what limits that?

STUDENT: The average height [INAUDIBLE].

PROFESSOR: It's the width of the doorway. It's the width of the doorway. That's why we don't-- that's why people get a second refrigerator.

And there's energy use. Energy Star program, big deal. Big deal. But I'm talking about refrigerators for a different reason. I'm talking about refrigerators because until the 1980s, the way we did cooling was with a chemical called chlorofluorocarbon. This is a chemistry class, and we're going to be talking about-- when we talk about why this matters, we connect to the chemistry.

Well, the chemistry there was critical. There was a molecule there called chlorofluorocarbon or CFC. And the thing about that molecule is that its absorption, the chlorofluorocarbon-- so let's write it down here, CCl_3F. There's a gas phase, so we put the little lower script g.

That molecule has its own absorption. It has its own absorption, and it has its own reaction, and so do other things. But it has a really important reaction because when it absorbs UV rays, then it doesn't just like then like electrons get pumped up. No, the whole thing reacts. And so this thing degrades into CCl_2F plus chlorine gas, and that's once it's up maybe in the atmosphere, which it was getting released all the time because of the refrigerators.

And the thing is that the chlorine gas would then react with ozone, and this would go

to ClO plus O2, and here's the problem. I mean, we're already at a problem, but now it's a big problem because-- you can already tell. But I've got ClO, and that's really reactive. So what happens is ClO, which was a gas-- I'm just being careful here with my subjects-- would react with atomic oxygen to give you O2 and the chlorine atom back.

And here's the thing. Now I've got my chlorine atom again. These are very reactive atoms in the atmosphere. They love ozone. And so this cycle for one CFC molecule would happen 100,000 times, roughly, before that chlorine kind of goes away and does something else-- 100,000 ozone molecules per CFC.

Why does that matter? Well, it matters because-- I love this plot. This is the energy of light from the sun hitting our planet. So this is the energy that you get, and look at this. This is the sun at the top of the atmosphere. Here it is on Earth, the red. Look at these things here. If you could only see in this spectrum, the world would be dark because there's no light on the planet. Why? Because it's all absorbed in the upper atmosphere by water. Absorption-- it's the same principle, Bohr.

Bohr's idea of electrons getting promoted when they absorb light is happening, and it's protecting us because right here, see, that's the UV. Low wavelength, high energy. UV, O3. Look at that. Look at what it's taking out. That little sliver of yellow is critical. And ozone was doing the job, and it still is.

But if we tear all that ozone out with these CFCs, we have a crisis. And that led to what is one of the greatest policy decisions which was made in 1987-- it's called the Montreal Protocol-- to essentially globally ban-- the US led this. This is a great example of how policy is so critical. The US led this effort to ban CFCs, and we're almost back to normal ozone levels, and we avoid this.

And by the way, if this had happened, the prediction at the time was roughly 280 million additional cases of skin cancer over this generation. This is a very big deal, and it all comes back to absorption of electrons by atoms and molecules.

See you guys on Friday.