Last week we talked about-- we went quantum.

We had Dr. Quantum show us the way.

And this week, we're going to go quantum for electrons, right?

That's what we're going to do this week.

These are my two favorite quantum mechanicists of all time.

I believe that this sign is meant to teach us something about the double slit experiment.

And just as you're leaving the lot, knowing that there is some chance that your one car could actually split into two, go through both exits, and recombine.

Just like an electron does through two slits.

And in fact, we know that that's possible now.

And if we wanted to, you could write down-- from de Broglie-- you could write down like the wavelength of a car, right?

So like in a car-- I don't know, your lambda of the car would be equal to h over mv.

And if you plus some numbers in, this is going to be something like oh, you know, maybe 10 to the minus 37th meters.

It's small, but it's a wave.

Your car is a wave, you know?

The baseball is a wave.

The electron-- we are all waves.

And I know, and that's what-- so the electron is a wave.

And I know-- hashtag, we are all waves.

Because it's a thing.

It's really a thing.

Because it's true.

I am a wave.

You are a wave.

We are all waves.

It's just-- our wavelengths are pretty small, but they're not zero.

And so the next time you get-- oh, we're on the same wavelength.

You really are.

It's true.

We are-- right?

That's deep.

We are waves.

All of us.

And so, the wavelength-- OK, the wavelength for an electron was something like 10 to the minus 10th meters.

Which made it a really nice way to interrogate images.

And that was my "why this matters" last week, right?

On Friday.

So you can take advantage of that wavelength to do things.

Like see electron microscopy, right?

But everyone's awake.

That's what quantum mechanics told us.

But what we want to do now is know how to represent this wave.

So everything's waving.

We've got that.

But how do I describe it?

Right?

And that's where we left off.

And we left off by me telling you that it all led to Schrodinger, who came up with the wave equation for quantum mechanics.

And the wave equation of Schrodinger is really pretty incredible, because-- so I'm not going to write out Schrodinger.

That would take me forever.

So I'm going to write S. Eq., right?

The Schrodinger equation, right?

And here-- so what he had is he had the sort of wave behavior that was captured.

And it was captured automatically.

And it gave us also allowed energies, you see?

So, allowed energies.

And this is also the quantization.

Ah-- I ran out of room.

Right.

But now, If you go back to classical-- so if you think about classical physics, what you would have is something like maybe F equals ma.

But what that gives you-- well, it doesn't give you wave behavior, but it gives you continuous energies.

Continuous energies.

So the powerful point here is that from that equation-- and we don't need to know this equation.

We're not going to derive it.

Right?

Today I'm going to tell you how it's solved, but you don't need to know how to solve it.

But you need to know what the solutions mean.

That's what we're talking about today.

But that equation-- what it does is it sets up the energies as a function of the wave.

Psi-- that function, psi-- is the wave function, OK?

And so you get waves by default, because that's what you're solving for-- these waves.

And then you have a minus second-- oh, there's an h bar.

That's an h over 2 pi.

Right?

But it's like the same Planck's constant.

OK.

But the wave behavior gives you-- because there's only certain energies that solve that equation.

So quantization falls out naturally.

You don't have to put it in anymore.

It is a direct result of solving this equation.

If I take a wave-- think about this, right.

If you take a wave and you put it in a box, then it can only have certain solutions.

Once you define the box, if you throw a wave in it, well, that's where those things are on the right are.

Maybe it's a standing wave.

And so it's going to look-- no-- it's going to be like this.

One solution.

And some energy.

Or maybe it's going to look-- I didn't leave myself any room-- like this.

And it's going to have some other energy, right?

1, 2.

That's it.

If you throw a wave in a box-- like, here's the first one, second one.

But there are only certain ones that are going to work.

All right, and you can get that sense.

Because otherwise it's not going to line up and then be able to wave and keep going.

It's not going to be a standing wave.

Right?

So you get this sense right away of quantization.

Only certain waves, and only certain energies, will solve this equation.

Now, here's the key, is this-- we're going to throw an electron into this equation, right?

But we're going to throw it into the right potential.

And that is the potential that we know it sits in, in an atom.

Which is simply a Coulomb potential.

So this is how we get there, right?

We take Schrodinger's equation of quantum mechanics.

And we solve it for an electron waving in the presence of the potential of the atom.

Which is a positive potential.

And you can see, this is a-- OK, V is a potential.

And that's an attractive potential, right?

That's an attractive potential.

So what happens when we do that?

So that's what we're going to do.

I'm going to leave this for-- so.

What happens?

So if we take a hydrogen atom-- or as I've written it here, a "Hatom," then you would have-- Let's see.

What do you get from this?

Well, you get that the potential is going to be minus 1 over r.

Yeah.

It goes as minus 1-- an e and an over-permittivity of space and stuff like that.

But it goes as minus 1 over r.

That's the Coulomb potential.

Right?

A plus charge is sitting outside of a minus charge.

Ah!

That was a close call.

The electron is sitting out here, and there's the nucleus.

And the potential feels is a Coulomb potential.

OK, so that's point 1.

Point 2 is that we're going to apply the-- oh, I'm saving myself so much time.

I love it, when you're efficient like that-- S. Eq.

Schrodinger equation.

The Schrodinger equation.

But the way that we're going to do this is in spherical coordinates.

Spherical coords.

And all that means is that I'm going to write psi as a function of r, theta, and phi.

And the last thing that I'm going to do is I'm going to separate the variables.

And that's just a standard way to solve mathematical equations.

So I'm going to separate the variables.

So separate.

And so what that means is that psi can be written as one function-- let's call it r-- of the radius, times another function, p, of the theta, times another function f of phi.

So that's my separation of variables, OK?

Again, I'm not all about having you know how to solve this equation.

OK?

That you can learn in quantum mechanics.

But what I'm all about is what that quantum mechanics solution means for us here in chemistry.

And this is deep.

This is deep, because-- why?

Because this whole thing is quantized, right?

So just like before, there's only certain allowed values for the wave function.

There's only certain allowed values for these variables.

Each one of them is quantized.

Each one of them.

OK?

And that's what we need to learn about today, is what that means for chemistry.

All right.

So now we're going to go through the quantum numbers and see what they mean.

And we're going to start with this first one, the radial function.

So radial, and then two angular functions.

OK.

Now the radial-- so what do we get?

Well, the radial function-- I'm using a word here, I got to tell you why.

Orbital of hydrogen, right?

So I've solved this now.

I've solved this equation, and I'm plotting this orbital for first quantization-- n equals 1, n equals 2, n equals 3.

We've got to talk about this, right?

But I'm just plotting what that function looks like.

Well, let's compare.

Because if I'm Bohr, then what I have is this.

Remember this?

I've got like this-- OK, there's 1, and it's going around, and it's very happy in an orbit.

But see, we don't call these orbits anymore.

And it's going to be clear as we talk about this picture, because these aren't distinct orbits, right?

So now in quantum what you have-- I'm just going to take that n equals 1 solution.

That's the first solution to the radial part, right?

That's not a fixed distance.

That's not a fixed distance, right?

So now I'm just looking at the radial distribution of the probability.

The probability of finding the electron.

That's how we speak when we speak quantum.

We don't say it's here, right?

We say it could be here.

I don't know.

But there's a probability that it's there.

So now it sort of has a probability cloud.

Now it looks like it gets a little less-- why am I drawing spirals?

It looks like it gets a little less probable as you go further out, right?

So maybe the density of probability-- because look, right?

It's got a peak there.

But even if I just plot that first orbital, even as I go all the way out to here, it is non-0, right?

So this is probability.

And this would be n equals 1.

So it's a totally different story.

That electron isn't in one place, but it is.

But it's not.

It could be anywhere.

It could be-- if I'm a hydrogen atom, my electron could literally be across the room.

There is a finite probability that that's true.

And we know that since we are all waves, and we can all be defined by the equations of quantum mechanics-- and in fact, that is the way to define the world-- that we also have probabilities.

They're small, you know.

But I could be over there right now.

There's a chance.

My weight function is not exactly 0 all the way over there.

It's close.

But it's not exact, right?

That is a very different picture of an electron in an orbit-- in fact, it's not an orbit.

It's an orbital.

And that's where we say orbital.

That's why we stay it.

OK, but-- we're listing quantum numbers.

These are the quantum numbers.

Oh, oh- I'm going all caps.

Didn't even know I was going to do that.

So we have n, which is called the principle-- principle quantum number.

And this is going to be related to the main energy, kind of e, 1, 1, right?

E, 2, 2.

So this is going to be related to the main energy level.

Main energy level, right?

And it's also related to the radial distance.

Radial distance.

And sometimes we call this the shell.

So sometimes the principal quantum number, this is the shell, right?

It's the first shell, or the second shell.

So we can call this the shell.

Come back to that later.

And you can see right there.

OK, we'll talk about n equals 2 and 3 in a second.

But you know-- oh, let's talk about it now.

You can see right there-- it I go from n equals 1 to n equals 2, then I'm further out.

Yeah, I still got non-0.

Doesn't quite go to 0 like I just said there.

But it's further out, it's shifted out, right?

So with increasing n-- so another thing we can say about n is that with increasing n, you get higher energies and larger distances.

larger r, a [? hole. ?]

But you get something else.

You get something else, and you can see it there.

And I've been avoiding it, but I got to address it right now.

Because those higher end functions go to 0.

So they have nodes.

With increasing n, you get more nodes.

What's a node?

A node is when it goes to 0.

And there's more of them.

Now that's also really important.

So what that means is in Bohr, my orbit was r, e.

Heisenberg said no way.

Can't have them both.

But I had a r, which was one value.

And remember-- it could only then be the r of the next value, the next n.

Now I'm basically saying the opposite.

I'm saying, well actually, r can be anything.

There's a probability cloud, so this electron could be anywhere with certain probability, except not here.

In this one place.

This node.

It's 0.

It's literally 0.

The probability of finding the electron there is 0.

It cannot exist there.

But how bizarre is that?

The electron can exist with probability out here, and probability in here, but it can't exist in between.

It's the same electron.

Try explaining how it goes back and forth.

It skips over a zero probability part, r.

Right?

But these nodes are very important.

And just to give you a sense of what this means, I found a very fancy animation, because I want to make sure you guys understand.

So here's the electron.

It's doing this, it's doing its waving thing.

But look at what happens with the nodes.

Uh uh.

The probability there is always 0.

You see that?

Can't go-- no.

Not gonna exist there.

Right?

That is super deep.

Because there's no classical analog to these things.

There's no classical analog.

But this is what electrons do in an atom.

Now, this is what the clouds look like-- a little better drawn than mine-- for n equals 1, n equals 2, and n equals 3.

So again, you can see-- this is one electron.

It's colored differently, just to contrast it, but this is one electron.

Right?

And These are plus and minus of this, right?

So it'd be like a wavy wave, and you're just showing plus and minus with the yellow and blue.

But that note in between is where it can never be.

But it's one electron.

It's not many, it's one.

And this is its probability distribution.

Now I want to make one really important point, which is that-- look at what I'm plotting here.

Maybe something-- that's size squared, right?

So it's only the square of the wave function that we can relate any physical meaning to.

The physical meaning I just related to you is the probability of being somewhere.

That's what size squared gives us-- the probability of being somewhere.

But ask your teacher-- not me, your quantum teacher.

Ask what psi means.

And they're going to have a-- it's going to be an interesting class.

Force the issue.

Don't let them go until they answer.

We still don't know.

We still don't know what the actual wave function means physically.

We cannot attribute a physical meaning to it, only to its square.

So we go about our business for the last 100 years squaring it and saying, OK, well it describes nature.

That's great.

But no one, still, knows what it means without the square.

I find that incredible.

And it led Bohr to say all sorts of really cool things.

He said, "We must be clear that when it comes to atoms, language can be used only as in poetry.

The poet, too, is not nearly so concerned with describing facts as with creating images and establishing mental connections."
That's Bohr like deep in it, trying to figure out what psi means.

He's like, I don't know.

It's poetry.

OK, that's cool.

That's cool.

V squared means something, v doesn't.

Really?

No.

Classically, we've got nothing.

Right?

Quantum mechanics.

Oh, it's a beautiful thing.

Why does this matter?

I'm going to give you a brief "why this matters." And we've been talking about the quantum nature of things, right?

Quantum is this.

Quantum are these differences.

It's this probability.

It's quantization of energies.

What does that mean, right?

Well, last time I told you that by electrons being quantum, and having wave-like character-- we're able to use them to see things, and it started the nanotechnology revolution about 30 years ago.

And today I want to give you an example of that in my "why this matters." And it has to do with quantum domination, which should be the title of a movie or a book.

And here's the example-- take a piece of bulk material.

Take a piece of silicon, OK?

And if you look at the silicon, it's kind of boring to look at optically.

It's not very interesting.

You know, don't tell that to your iPhone, which has a lot of it in there, but it's boring.

But now, take out your little nano ice cream scooper, which you all have.

And you take a little tiny piece of it.

We call that a quantum dot.

Why?

We call it a quantum dot because it kind of looks like a dot-- its smallish.

But we call it that because quantum mechanics takes over the properties.

Here's an example-- shine light on this thing, or have it glow, and all of a sudden instead of being boring it can be any color you want.

Why?

Because if I shine light-- and here is a very sophisticated picture of a laser, which you can tell I had trouble making transparent.

And so you shine a laser on this piece-- we know already, light excites charge in an atom.

It does the same in solids.

So here's my piece of silicon, and what that does is it sends an electron up in energy levels.

But the electron is up in an energy level, and it left behind a hole.

Which is something we'll talk about later when we talk about semiconductors.

A hole is just a positive charge.

And it turns out that that electron and that hole are attracted to each other, but not-- they can't get too close.

But they want to kind of hang out at a certain distance.

Where?

How far?

Yeah-- quantum mechanics.

N 2, 3, 4 tells us.

It tells us how far, right?

They want to hang out-- and now all of a sudden, I've nano-scooped them out, and I've made a quantum dot.

But if I try to do the same thing there, that's how far they wanted to be.

And they can't be.

They literally ran out of real estate.

They ran out of atoms.

They cannot be out here, because there's no stuff out here.

So they get squeezed.

I am squeezing these quantum mechanical objects, this electron and this hole.

And by squeezing them, I'm changing the way they interact with light.

I'm changing the color, right?

Oh.

It's all stuff we've learned that explains something revolutionary.

By the way, if you buy a QLED TV, is it as good as OLED?

No.

But is it better than LED?

Yeah, a little bit.

That's what they-- QLED-- quantum is quantum dots.

They have coated the LED emitters with quantum dots to give you better color.

Now what this means, why this matters, is nothing less than the periodic table itself.

Because I think I showed you this-- we're living in these different ages, which I love.

We have the periodic table that has-- the ability to work with it is what's brought us to the age of materials design.

But see, now I've literally just told you that by changing the size-- nothing more than the size-- I can tune a property because of quantum mechanics.

Because I'm changing its quantum mechanical interactions.

It would be like saying, I can take a piece of this table-- I can break a piece of this table, and it's now going to be red.

That's exactly what we're doing, but at a very tiny size.

That is as if we are taking this periodic table and giving it a whole other dimension.

Every element can do more things than we thought possible because we can tune these properties related to quantum mechanics.

So that's a big deal.

That's a very big deal.

OK.

Now back to our solution to the Schrodinger equation.

For an electron in an atom.

Because now we move on and we solve for the next one, which is-- OK, so we did r.

Now we're going to do theta.

And what happens with theta is, you get another quantization of that variable.

And that's the second quantum number.

So we're going to put it here.

N-- I'll number them so we don't lose track.

Who knows how many they'll be?

OK.

Not three.

OK, so the second one is l.

And it's the angular quantum number.

And it has also quantization 0, 1, 2, 3.

OK?

So l can be-- Oh, I need to put that up there.

I meant to put that right here.

This is just like what we learned for Bohr, right?

And l can be 0, 1, 2, all the way up to l minus 1.

OK?

I'm going too fast, that's n minus 1.

L goes from 0 to n minus 1.

OK, so what that means right away-- what you can see-- if it only goes to n minus 1, then for example for n equals 1 then l equals 0.

That's it.

There's no other options.

It stops, right?

And we have a name for this.

We call this s, the s orbital.

And for n equals 2, l can be 0, or l can be 1.

And this one is called-- this one is the p orbital.

This is the p orbital, OK?

Those have names.

And so on and so on.

d, f-- This is old notation that we cannot get rid of.

This is called spectroscopic notation, because it comes from spectroscopists.

And remember, I told you-- never make a spectroscopist angry.

Don't do it.

Don't do it.

And so we use their notation still.

s was sharp.

Spherically symmetric.

p was principle, because it sort of dominated in some experiments. d was more diffuse.

f fundamental, because it looked kind of like the hydrogen atom sometimes.

These shouldn't be the names, but they are because they still stay with us.

Spectroscopic notation-- s, p, d, f.

What they really do is just correspond-- that looks like "porbital." Let's put a little more space there.

P orbital.

Right?

Because what they really mean is nothing more than a quantum number for the second quantum number, which is the angular quantum number.

These are also sometimes called subshells, right?

Because they're underneath the shell.

And as you can see from this picture, it describes the angular shape.

The angular shape of the orbital.

The shape, right?

So it gives us a shape.

Now-- if we think about it, and we think about that p orbital, that has a different shape than the [INAUDIBLE]..

Then you think, well, this could be aligned in different directions, right?

So if I think about that p orbital-- well, I could align it in three different ways.

It could go along this axis, this axis, or that one, right?

And so those are p orbitals in hydrogen that have three possible orientations, and that automatically gets us to another quantum number.

In fact, it's the last one, and it makes sense just physically-- that's phi, right?

So the last one-- or is it the last one-- is m.

And we call it m sub l, which is the magnetic quantum number.

And this one can have quantization equal to minus l, right?

All the way up to plus l.

So the number of orientations that the orbital can have-- if I'm a-- n equals 2, l equals 1, then m sub l equals minus 1, 0, 1.

Three orientations.

Right?

It's got three orientations-- three possible magnetic quantum numbers.

Why is it called the magnetic quantum number?

Because those are the experiments.

We're done using magnetization-- magnetic fields-- to discover these orientations, to sort through it.

To see it.

OK.

So if I had l equals-- all right, so if I had, for example l equals 2.

Then m sub l would have five values.

Five possible values, right?

I should put a plus there, just to be sure we keep track of minus and plus.

These are quantum numbers.

That means they're integers, right?

But they can have these ranges and these rules, and again-- you may think, why did this come out like this?

Well, it comes out from solving this equation-- separation of variables-- with that potential.

That's it, right?

That's it.

You get these allowed values of these three variables, quantum numbers.

OK?

All right.

Now-- and there's the d.

These are called d-- So if n equals 2-- OK.

So did I mess up here?

No, I didn't, because n would have to be 3 here, right?

For l to be 2.

Because l can't be equal to n, because I wrote it there as the list of allowed quantum numbers that l can have.

And there it is for the d orbitals.

And so there's five d orbitals.

Five d orbitals.

And you can do f on your own, and it's an extremely fun thing to do at night.

Late at night.

Now, this gives you the orbitals.

You get them all.

These are the orbitals.

These are the beautiful, beautiful orbitals that we have from quantum mechanics describing an electron in an atom.

Which tells you everything about where that electron can be, how it behaves, what its energies are, and therefore literally why atoms are what they are.

That's what it tells us, right?

Because the next thing we need to do is fill up our atoms.

And know how electrons fill atoms using these quantum numbers.

And then relate that to chemistry, and to the differences between one atom and another, and explaining why atoms are different.

Remember, that was our goal.

We needed to go quantum to answer that.

So we're getting closer, but we're still not there.

And one of the reasons is there is a fourth quantum number.

And this came out of experiments-- I thought that I would just show you a very quick video, because it's easier to see this in action.

And so these are experiments that were done by Stern and Gerlach that showed something very peculiar about atoms and their electrons.

OK, so here I'm just going to play it.

It's got a really nice-- oh, that's loud.

Yeah?

I don't mind that.

That's good.

So you can read the text here.

OK, so they've got a magnet.

A very strong magnet, right?

Magnetic fields.

And they're going to shoot a classical magnet.

There's a classical magnet.

And you see it's going to react to the magnetic field, and be pushed up.

That magnetic could've been oriented differently.

Like sideways, right?

In which case it doesn't feel this.

And you could throw a bunch of them randomly, with random orientations, and that's what you would see.

[JAZZY MUSIC]

It's good music, too.

But now-- watch this.

I love the animation here.

Look at this.

And the bass, right?

Those are what happens to atoms with electrons.

You'd shoot them through a high field, and you don't get anything.

You get two things.

Only two.

[JAZZY MUSIC]

I'm just playing it because I really want to groove out and I'm debating-- OK.

All done.

I was debating how embarrassing it would be if I just started grooving out.

So this was Sterin and Gerlach, who discovered, essentially, a fourth quantum number that can have only two values.

So they showed that this thing that the atoms were doing was because of the electrons.

And that the only way you could explain what these electrons were doing is if there was this thing called spin.

And if you really need to think classically-- and it's dangerous to do so, because then you'd forget that I could be all the way over there, right?

Right now.

But if you really need to think classically, you can think of a charge spinning.

It could be spinning this way or that way.

In a way, it helps our classical needs.

But it's so not what's happening.

But it does help our classical desires, right?

And that would give you the two fields.

Right?

But I would advise you just to think about spin as simply a fourth quantum number.

And so we call that one number 4.

Not from that equation.

Separate from those experiments.

And that is-- we call that the spin M sub s is the spin quantum number.

And it's the fourth one in our series.

The fourth quantum number.

And it has only two values-- up or down.

Or if you like, you could write it as plus 1/2 or minus 1/2.

Or you could write it as up or down arrows.

So many ways to distinguish two things that are opposite.

Spin up, spin down.

Two values that this can have.

So we're not ready yet.

We're still not ready.

We're so close, but we're still not there, right?

Because we still don't know, well-- can't I just--?

How do I--?

What are the rules here going to be?

For populating atoms.

And this is where Pauli said something very important, and made this incredibly important statement.

Which is that-- no particles in the same system-- read that as, no electrons in the same atom-- can have the same quantum numbers.

No electrons in the same atom can have the same quantum numbers.

And that immediately tells you-- again, this has to do with the nature of quantum mechanics-- what that immediately tells you is that two electrons-- ho-ho, this is big!

Two electrons-- I'm gonna write, two electrons, and watch this-- I'm going all caps-- only two electrons can occupy an orbital.

An orbital.

Or- bi- tal.

That called the Pauli exclusion principle.

OK.

So we're going to use all of this information to fill stuff up, but I don't want to do that yet.

On Wednesday, we're filling up the table, and we're making our first solid.

I figure it's our three week anniversary.

I know you've all been looking forward to it.

I've been thinking about it a lot.

And I think-- that makes me want to throw t-shirts.

[STUDENTS CHEERING]

I got to do this.

All right, all right.

I got to go-- I know.

I got to go backwards here.

OK, OK, hold on, hold on.

I'm missing there and there, I know.

I got to work on my arm.

Yeah.

Show you love over there.

OK, all right.

I'll have more on Wednesday, and the reason is-- we're making our first solid.

I almost wanted to leave it a surprise, but I couldn't help myself.

It's extremely exciting.

This is, after all, solid-state chemistry, and Wednesday we're making a solid.

Literally.

We are not only making a bond, but we're making a solid, and talking about solids.

So that's going to be very exciting.

And we're going to start by filling elements.

But before we do that, I want to explain something very important, which is-- what's happening with energy?

Right?

We did just see something dramatic.

As you go from Bohr to quantum, something very dramatic happens with where, right?

And we know something else very dramatic happens, which is right here.

You know, that's Bohr over there- n equals 1, [INAUDIBLE]..

Bohr is 1, 2, 3, right?

Now it's like-- oh no, wait, it's 1 and with 1s.

2-- I can put two electrons in there, thanks to Pauli.

I know I can put two, because I have all the quantum numbers in here the same except spin, which could be up or down, right?

N equals 1, l equals 0, m sub l equals 0, spin is up or down.

Two electrons.

But look, I'm now here at n equals 2, going from the Bohr model to quantum.

I've got this and I've got three p orbitals.

So now all of a sudden I've got eight electrons that can go into that quantum number.

That's very different than Bohr, right?

And that's called degeneracy, because you're taking something and you're saying, there are multiple variations that have the same n.

But there's something else that I want to explain right now that is can be understood very physically and intuitively.

And that is that the motivation-- one of the big motivations of this was that we needed more than one electron.

Our atoms are not all hydrogen and ionized atoms.

They are just neutral atoms going up above hydrogen.

So what happens when we go multi electron?

And there's two effects that I want to explain.

Here's what happens to this picture.

Those are the one electron levels, right?

There's Bohr over there.

This is what you really get in terms of energy.

Right?

This is energy along the vertical axis.

And there's two reasons, two very important reasons, why these are so different.

One is called shielding, and the other is called orbital penetration.

And I want to talk about both of them in the last seven minutes of class.

I'm going over here.

All right, so let's talk about-- let's do shielding first.

Shielding is something that you can kind of get a very good sense for it if you just look at what's happening in these atoms.

So in shielding-- let's take an example.

I'm going to give you an atom with 15 protons and 15 electrons.

What atom is it?

Anybody know?

Phosphorus.

OK, so I'm going to take an example of the phosphorus atom, OK?

And I've got 15 positive charges in the nucleus.

15.

Right?

Now, 15 protons.

Now I've got my levels, and it gets very difficult to draw levels as probability distributions.

So I'm going to draw them as rings, even though we know they're not, right?

These are my orbitals.

But I've got like an electron here, OK, an electron there.

All right.

So I've got two electrons in the n equals 1-- n equals 1 orbital.

So that's the 1s orbital, right there on the bottom.

But if I go out-- OK, so I'm going to go out now.

To n equals 2.

Well, now I've got 1, 2, 3, 4, and-- oh boy-- put one here.

5, 6, 7, 8.

Because here-- remember, I've got l equals 0 and l equals 1, and as we just said, because n sub s can be up or down, I've got eight electrons here.

So here, I've got n equals 1, here I've got n equals 2.

That's eight electrons.

And then out here-- oh, now it gets really interesting-- because for n equals 3, how many electrons do I have left?

1, 2, 3, 4, 5.

And you can already see what's happening.

So remember, Rutherford told us this.

All this charge is in a tiny little point here.

You wouldn't even see it if we were drawing this to scale.

But what are these electrons seeing?

What are the electrons seeing?

Well, these inner two electrons are just seeing this huge amount of charge.

They are happy.

Low energy.

Right?

The lower your energy, the happier.

You're in a lower energy, ground state.

Remember-- more lower energy, more happiness.

More bound.

Right?

So those electrons down there are enjoying.

But these out here-- you see?

These electrons.

Those electrons in here-- these shield the outer electrons, these shield outer electrons by two electrons.

Right?

Because they're kind of in the way.

It's like you've got this negative charge that's in between me and the positive charge, right?

And so if I'm out here at n equals 2, I don't see 15 positive charges there anymore.

I see 13-- ish.

That's shielding.

But now this one-- this has the 2s and the 2p electrons.

It's got eight electrons there, and those are shielding another eight.

Those are shielding another eight protons.

To things outside of them, right?

Shielding doesn't work going in.

If I'm this electron and I look out, all I see are repulsive-- I see repulsiveness.

In a good way-- I'm sure they're friends.

But two negative charges, right?

So I'm only looking in.

Looking in, shielded.

Looking in, shielded.

So these ones have sort of an effective-- z effective.

You know, that could be something like five.

This is not the z of the atom, but it's sort of like those very outer electrons don't see the whole proton cloud.

They get shielded.

And that's a very important part of what happens to orbitals when you put a lot of electrons in them.

And there's one more thing that happens, and it goes back to this plot that I showed you before, which is that the orbitals themselves are very different than just like a straight line from Bohr.

They have these shapes to them, right?

So the second thing that happens is that if I have that same s-- this is s, so this would now be the 1s.

And now I'm going to put-- remember that 2s looked like this.

And it had the node there.

Remember that?

Ho-ho.

That was 2s.

But now we're going to put the 2p there.

May not be drawn to scale.

OK, here's the 2p.

Something very, very interesting is happening here.

And again, it has to do with what electrons are seeing.

Because the 2s electron has a probability-- remember, this is the probability normalized, so we put in r-squared there, right?

That probability-- all that means is if I integrate over everything, it's 1.

It's got to be somewhere.

But look, the 2s electron can be really deep inside the shells.

It can be in there.

Literally-- look, it's even inside the 1s.

So it gets some of that 15-proton action, right?

Orbital penetration leads to massive changes.

Orbital penetration is what this is called, and it allows some of the-- but the p doesn't have that.

It goes in a little, but the p doesn't have that peak, all the way in there, where it can see all that positive charge.

And that is why these two things are different.

You see?

Because the 2s now, can be closer sometimes to those 15 protons, right?

And so it lowers the energy, making it happier.

It finds a lower happy place.

OK?

So we're going to use this, and on Wednesday, we're going to take these things and we're going to fill them up, and we're going to make our first solid.

Have a very good rest of your day.