We are going to be leaving the territory of-- some of you have had high school chemistry.

We're going to be starting to leave that territory and get into some of the things I've talked about before, that you may have never seen.

So for example, today we'll be talking about semiconductors and what happens when you take all these atomic orbitals and you put them into a solid.

And then, by the end of the week, we'll talk about metals.

And then next week, we're going to talk about crystals.

So that's what's coming.

OK, now there's one type of bonding-- Before we get into semiconductors and solids, there's one type of bonding we covered but I didn't put the stats up on the column.

So I just wanted to kind of make sure that we finish that, and that was hydrogen bonding.

And there's a little video-- you know, water, we're not going to go into all the amazing properties of water.

But if you're interested, this is a really cool website that talks about all the anomalies of water.

And many of water's properties are anomalous.

And much of the reason for that is the last type of bonding we talked about, intermolecular bonds we talked about last Friday, which is the hydrogen bond.

So remember our table, I just wanted to fill out the table because I didn't fill out the table, right.

That's the type of bond.

And the model was some kind of electronegative-- it was some kind of highly electronegative atom.

Like a was oxygen or nitrogen or fluorine.

So that you drew a lot of charge off of that hydrogen and you left it with a really strong delta plus so that it could then bond to some other atom.

It might be the same, by the way.

It was in the cases that we talked about.

Actually, I'll just use the same atom because we use HF and H2O, so then it would be the O or the F.

But it's to a lone pair on that atom.

That was the hydrogen bond.

And the range of attraction is in kilojoules per mole of bonds is something like, let's see, 10 to 40.

So it's pretty strong, you know, compared to some of the weaker things like London, the hydrogen bond can be pretty strong.

And one of the examples that is the classic example would be Hydrogen, hydrogen, there's the hydrogen bond right there, you see?

Hydrogen bond between another oxygen with its hydrogens as well.

Now, we talked about how water is so special because in the case of water, not only do I have exactly two hydrogens in each molecule that have caused this strong delta plus, but I've also got two lone pairs.

So the numbers work.

In HF, that's also a hydrogen bonded-- it can hydrogen bond to itself, but I've got three lone pairs on each F and only one hydrogen that I've taken that charge away from.

So you still get the hydrogen bonding, but this is the beautiful balance.

And there's a video showing water bonded to each other.

Oh, there it is.

Because there's nothing like just imagining what a hydrogen bond looks like.

And so of course, a hydrogen bond will-- OK, charges.

Oh, it's French.

[SPEAKING FRENCH] OK.

Oh, that's what it looks like.

Really, if you could get down there, it would be a ray of a laser beam light.

It's a liaison hydrogene.

And there it is.

And this is what's going on in the liquid.

They're kind of seeing each other.

And they're talking.

But they feel this strong, strong hydrogen bond.

And that's happening all over the place.

And in a liquid, they are making and breaking all the time.

Question?

STUDENT: Why do they separate?

Yeah.

Yeah, because-- well, it depends.

It depends.

It's a great question.

Why do they separate?

Well, they don't in ice.

So in ice, I mean, you know, they can.

But in ice, it's a solid, so they're sort of locked in.

And so in there, you're going to have pretty much each water molecule has four-- look at this.

It's you know from VESPR, right?

Four possible bonds.

You can make tetrahedral.

But there's actually lots of ways you can orient them in a solid.

There's many forms of ice, 12 different ways to make a solid of ice.

But see, here-- gesundheit.

So here in the liquid at room temperature, with the laser beams, they're coming in and they're bonding, but they have enough energy to break the bond, too.

So they stick and they unstick because there's enough thermal energy in the system to continuously break them and remake them.

That's when it's a liquid.

When is it a liquid versus a gas?

Right, the boiling point, remember, that is a proxy we talked about last Friday.

That is a proxy for the bonding strength between these molecules.

If it's a stronger bonding, then the boiling point will be higher and higher.

Oh, water should have boiled at a much lower temperature if it didn't have this very special hydrogen bond, which is not a dipole-dipole.

People try to write the physics of the hydrogen bond as 1 over r cubed.

That is not true.

That is not just as simple as, it's another dipole-dipole type interaction.

It's complicated.

Complicated.

Remember, we had covalent bonds were complicated too.

So we're not going to put a simple one over r something dependence for the hydrogen bond.

OK, good question.

OK now, all right, so that's an intermolecular-- now we're going to move on to solids.

So that was our last intermolecular interaction.

Oh, before I do, there's actually an example question I put there that you could now answer, which this is actually-- I just want to show you the type of-- so we covered all these intermolecular bonds.

Here's a question from a quiz last year.

List the intermolecular forces in each substance, methane, and then you've got that CH3OH.

And then you've got the CH3CL.

And so you can do this now because for methane all you've got is London.

It's not a polar molecule.

All you've got is London.

There's no hydrogen bond.

But for CH3OH, do you have hydrogen bonding?

Yeah, you do.

Because if you draw-- because you've got an H that's had the charge taken from it by oxygen, a highly electronegative atom, and then and then you've got lone pairs on that same oxygen.

So you can hydrogen bond and you've also got London.

All right, now is that it?

Is there anything else?

CH3OH.

What else?

Is there any other type of bond?

Well, the last question you need to ask is, is there a dipole moment in the molecule?

Because if that molecule had a permanent dipole, then there is another type of intermolecular interaction, like we talked about, dipole-dipole.

So you'd have three.

And for CH3CO, well there, you also-- you know that for CH3CO, you're going to have a dipole moment.

Remember, we compared dipole moments of molecules.

And if you have a diaper moment in a molecule, then that dipole moment can talk to another dipole moment and be interacting with it.

So that's a kind of bond.

And is that it?

London, London is always there.

It's always there.

We've also got London, don't forget.

So you'd have dipole-dipole plus London here.

Dipole-dipole London and hydrogen, just London.

So this is the kind of way I want-- and you might be able to think about, well, would a trend-- if you think about, like, OK well, if a molecule just has London, then what does London depend on?

All right, well, it depends on the polarizability of the charge.

So that can be something [INAUDIBLE]..

It depends on the size because that tells you how much-- if I say, well, what about SIH4 compared to CH4?

Which is going to be more strongly bonded?

You can answer that because you know the only force there is London.

And you know for London, it's how much contact area you have is going to be one of the factors.

The more contact area, the more chances these things can get close enough and polarize.

And then polarizability would be the other.

All right, now we're going on to solids.

Now, the big picture is pretty awesome.

If you look at the big picture, we have done the electronic structure of atoms and molecules.

So we've gone through and we've worked our way through the Bohr atom, the atomic model, quantum numbers, Aufbau, what happens when you go from one electron to many, and octet stability, which is Lewis and Vesper.

We have also now talked about how molecules come together in different ways, how the atoms come together in the molecule, and how the molecules come together in different ways.

Dipole-dipole, London dispersion, hydrogen bonding.

And then there's these forms of aggregation, gas, liquid, and solid.

And of course, we're excited about this.

So I put my little emoji there, because that's what we're going to do now.

We're actually going to take-- we talked a little bit about solids already, but now we're going to take the electronic structure that we've built up, these orbitals.

And we're going to make solids out of them.

So this is the week where these are the next two weeks, right, where we're going to talk a lot about solids.

And this week, this week we're going to do the electronic structure.

And next week, we're going to do atoms.

So that's our plan.

And next week Friday, we're going to break a bunch of stuff with parents.

So this week we're [INAUDIBLE].

What do electrons do in solids?

Next week, how do atoms arrange in solids?

And together, that gives us so much understanding of how chemistry effects properties in solids.

So that's the big picture.

Now, here's the thing, why does that matter?

Well, because so many properties can vary by so much depending on what I just said.

How does chemistry affect properties in solids?

Well, it's complicated, but look at this variation.

28 orders of magnitude of connectivity, electrical conductivity.

28 orders of magnitude, that's a lot of zeros.

[INAUDIBLE]

These are the electrical conductivities of different atoms, different materials.

And you can see that it's going to depend on what you have, but it's also going to depend on-- so that's the chemistry.

But it's also going to depend on the bonding.

It's going to depend on how the electronic structure comes together.

And it's going to depend on how the atoms come together.

That is what will explain 28 orders of magnitude.

That is a lot of variation.

So we start by thinking about how electrons see these solids.

That's how we're going to start.

Now, the thing is that these good old days of molecular orbital theory are going to end.

They're going to have to end today.

No, they're not ending for like quizzes and exams.

Don't forget about them.

But we've got to move on because this is just two atoms, two atoms.

But what happens when we have like a solid?

How many atoms do we have?

We've got a lot.

We've got a lot.

So let's go-- we'll go here.

OK, now, if you think about that picture, that was our-- I'm going to take 1s orbitals.

And let's just think about this for a second.

If we take 1s orbitals and you have two atoms that come together, well, they're going to do just what you see there where you've got your 1s and your 1s.

I'm not even going to fill these out, and they're sigma and sigma star.

And so that would be like two 1s orbitals that I have either-- OK, so up here, I took one.

It was a positive and I subtracted one.

And here, I took one and I added one.

And remember, this gave us this and this.

And this gave us the bonding orbital, right?

Bonding, anti-bonding.

It all came from LCAO.

LCAO is the linear combination of atomic orbitals.

That's the theory we use to build up our [INAUDIBLE] theory.

But now, we've got to go a lot further.

So I'm going to-- because I don't want to keep adding things here and here, it gets complicated.

I'm trying to find the MOs for more, and more, and more atoms because I've got to put a whole lot into my solid.

So I'm going to write them just on one side.

If I write the two here like that, and that would be sigma and sigma star, then you know what I'm talking about.

I'm taking the two and I'm combining them in two different ways.

But what if I had three?

Well see, if I had three-- so I've got three s orbitals now.

So if I've got three s orbitals, then if you think about it, I could also add them like that.

That would be a very nice bonding orbital.

If you add them, you see that they're going to have the same kind of really nice bonding.

But I can also subtract one in the middle.

And if I did that-- so I had a plus, minus, plus there.

So those are the signs, plus s, minus s, plus s.

Then you can see that in-between those, as I add these together, you have nodes.

So this is not going to look like a very happy bonding situation.

I'm going to have something in there and then something in there.

That's what that's going to look like as I bring them closer and try to add them.

I made two nodes here.

Well, there's one other way you could do this.

And we know that the number of molecular orbitals has to equal the number of atomic orbitals.

We know that.

So there's one more that we mean, because I started with three s's.

And you can kind of see that.

You can see that because if I had a plus and a plus, if I had done it that way or maybe the other way, maybe this was over there, but you see I got a little bonding going on in here between them.

And I got an anti-bonding there.

It's almost like it kind of cancels out.

So you can imagine that that combination of s orbitals, if I add and subtract them this way, it's going to be somewhere right in the middle.

Maybe not so bonding, maybe not so anti-bonding.

OK, good.

All right, well let's do four.

So if I had four, if I started with four-- this is like four hydrogen atoms, let's say, although I'm not filling it yet.

But so four 1s orbitals.

Well now you see, I could do this.

I can add them all up and get this super awesome bonding state.

I could also alternate them so that everything is anti-bonding.

Look at that.

These are going to cancel as you bring them together and you're going to get anti-bonding, anti-bonding.

But there's also two other options that are in between.

Right here, I can add 2 and then subtract 2.

And you can see, these are both negative.

And as they come together, they're going to overlap constructively, so will these.

But in the middle here, they're going to be kind of destructive interference.

And so you're going to have this one node in between.

But in general, you can imagine, this will be more bonding than anti-bonding, so we draw it down below the middle.

Whereas up here, you could also imagine something like this, where you start with a plus and you put two minuses in there and a plus over there.

I'm just adding and subtracting.

I'm just taking LCAO theory and moving it to more than two orbitals.

OK, so now you can actually understand this from a number of nodes, right?

This is like three nodes.

This is two nodes over there.

This is one node.

And zero nodes.

This is exactly the kind of thing we talked about already, but for only two orbitals.

Now we're just continuing it along.

OK, you can see why the good old days are over, because see, now we've got to jump this up a level.

And we've got to think, how many s orbitals do I have in a typical piece of matter?

How many do you think?

Take a guess.

Roughly?

Take a chunk of hydrogen this big.

How many atoms do I have?

A mole, exactly.

That's why we invented the mole, so that we could count big numbers by just saying 1, 1 mole.

But if I have a mole, then I've got like 10 to the 24th of these.

I just went to four.

I just went to four.

But it just keeps on going.

So [INAUDIBLE].

So here is like, OK, 1, orbital of atom.

Boy, I labeled that well.

2, bonding, anti-bonding orbitals of molecules.

And then you see, OK, there's 3.

There's 4, 5.

Then you keep on going.

OK, they got a little tired and then they wrote 30, 30 orbitals.

30 is still way, way far away.

But if you had 30s orbitals, you might still be able to kind of see these lines discretely.

Remember, this is energy space.

There might actually-- remember, like, these energies are really important, the lines between the deltas.

But if you have a mole, if you have a mole, it looks like one continuous set of lines.

These don't really look discrete anymore.

These don't look discrete.

And so if you have a mole of atoms, so if you have a solid, we call that collection of orbitals that are all stuck in here a band.

We call that a band.

OK, so if I take-- Now, I'm not going to draw 10 to the 24th.

10 to the 24th s orbitals.

But pretend, pretend that that's what I have here.

Then what happens?

What I do is I form a continuum of states, a continuum of levels for those electrons.

And this, they're so packed together, we don't even draw them as lines.

We can't.

So this is called the 1s band.

Oh, band.

It's a band.

It's a band of states.

It's a band of states.

It's the 1s band.

Now, there are some important rules that are really just the same but we follow them here too.

And so I want to emphasize them.

We've already learned them for molecules.

They holds for bands and solids.

All right, so now, for example, let's think about this.

Each of these orbitals could occupy how many electrons?

Two.

So I have two electrons for each s orbital, s orb.

Well, you know, if you think about that, then for n atoms, and here what did I say, 10 to the 24.

Because there is one s orbital that comes from each atom.

So for n atoms, then I would have 2n electrons can go into the one s band.

So if I have-- that makes sense, right?

Because I had two for each orbital, and I had that many.

So if I know how many atoms I have that make up this solid, then that's how many electrons that band can have in it.

All we did is go from the atomic orbital picture to the solid to the mole.

OK, why am I saying that?

Because filling is incredibly important And it happens in the same way-- just like it was so important for molecular orbitals, it's also critical for bands.

So filling is like MOs.

It's just like in MOs, start from lowest energy.

Then fill up.

So you know, if you think about this, if I had-- now I'm going to actually say, well, OK, I made the one s band by combining these orbitals.

Now I'm going to see what I have, which tells me how to fill it.

Now I have hydrogen. So if this is the one s band and I have hydrogen, hydrogen has, for each hydrogen, I've got one electron.

But I just said that in the one band, for each atom I could have put two electrons.

So you know if this is made of hydrogen, it's half filled, which is exactly what that picture shows.

That's the filling.

You see it's filled half way.

The band is filled halfway.

It's like the MO picture, but it's not because I've made a continuous band out of it.

So instead of putting lines here, lines there, and then occupying them with arrows, I fill them up with 10 to the 24-ish electrons.

So we fill bands up like that.

And we know that it's half filled because it's hydrogen. Another thing that we know is that the energy-- remember this?

This is also important.

That energy is related to interaction and overlap.

This is something we talked about, interaction and overlap of the orbitals.

How much does that band go up and down in energy?

This is something we've talked about.

How much that band goes up and in energy, it's just like in the H2 molecule.

One s, one s, sigma, sigma, star.

How far apart are sigma star and energy?

Well, it depends on the interaction strength and the overlap between those orbitals.

The same is true in solids.

Same is true in solids.

So what is the width of that band?

How spread out over energy is it?

OK, so those are some ground rules.

Now, what else can we say?

OK, so there's the absolute least bonding state, anti-bonding, all the way up there, the most bonding.

And then you've got almost infinite number of possibilities in between.

But the thing is that filling determines the properties.

You can't just know the bands or the band structure, as it's called.

But you have to know the filling.

And when you have a situation like this-- so let's write this down.

When you have partial filling, partially filled band is a metal.

And I will talk about metals on Friday.

So we will come back to that concept and the electronic structure of metals on Friday.

But today, I want to talk about something else that happens, which is when you fill not in the middle of a band, not in the middle, but maybe all the way up to the top.

And there's a gap.

That's a different kind of material.

So if filling stops somewhere partially in the band, in the middle of a band, if you're filling, you're counting, and you didn't go all the way up, then you have a metal.

But the thing is, here we've only talked about one kind of band.

This is the s band.

But every single one of these-- look at this, I can take a mole of orbitals and make an s band, like that.

And then, oh, that was 1s.

Here's 2s.

And look at this, I can take the p electrons and make a p band.

2p band.

I can do this with every single-- 2s, 1s, right?

So depending on a bunch of factors, these bands could look different.

What does it depend on?

So this is-- the way these bands are organized, right, we know now that these orbitals will come together like I've drawn there and there and form bands.

What is the difference between these?

Did they overlap, maybe?

Are they very far in energy?

These things depend on atom-- depends on the atom and the bonding and the structure.

How do those bands look?

Well, it depends.

And we'll be talking about some examples.

But for now, I want to focus on that filling thing again.

What happens if I-- if I fill this 1s up to half way, it's a metal.

If I keep filling, keep adding electrons, and I fill this up halfway, it's a metal.

But if I stop filling, you know, if I fill this and then I fill this, and then there's some space in energy and I didn't fill any more, that's not a metal.

That's not a metal.

So if there's a gap-- so there can be gaps between bands, which I just showed you.

That's between an s and a p band.

And then the filling determines the properties.

And they're so important.

They're so important.

So let's talk about that.

I want to talk about these.

What we're going to do is we're talking about semiconductors today and Wednesday, and metals on Friday.

And we're going to see it through the vantage point of these electronic levels.

And then next week, we'll see it through the vantage point of atoms.

So, if you have a gap between-- well, you always have gaps, the question, again, comes to filling.

If the filling stops-- I want to define this very important-- so if the filling stops, then you have this highest occupied level.

So let's suppose I have filling-- this is some band.

And this is some band.

And this is unoccupied.

So unoccupied by electrons.

This is fully occupied by electrons.

So what's going to happen?

Well, first of all, let's just get some terms.

The last band that I filled, the last band that I filled, is called the valence band.

Valence band.

And we use vb for short.

And the first band that I don't fill is called the conduction band.

And as we talk more about these materials, you will see why they're called that.

Conduction band.

But for now, just trust me in the labeling.

So the conduction band and the valence band depend on the filling.

Because again, it's, you know, it's where I filled to.

Now, the top level, the very, very top level out of these 10 to the 24 level, that very top one has a special name.

It's the vb maximum.

That makes sense, right?

It's the top valence band, or the vbn.

And the one down here also has a special name, it's the cb minimum, or cbm.

So the valence band maximum and the conduction band minimum are, you know, it tells you, if you think about it, it tells you if you subtracted those energies between those parts of the band, you'd get the distance and energy between the two bands.

So the conduction band minimum is this energy level here and the valence band maximum energy is-- Now remember, valence electro-- it's called the valence band.

It makes sense because valence electrons, remember, are those ones, as we talked about with Louis, where you've got kind of the bonding and the action and the chemistry happening.

And so this is that top most state of the top most valence band.

This is the valence band top most state.

And there's one more definition we got to know about.

This is why we're setting this up, which is that that energy difference here, oh, that is really important.

So the energy of the conduction band minimum, minus the energy of the valence band maximum, is equal to something called the bandgap, energy gap.

It's the band gap.

This is an extremely important property of materials, the electronic band gap.

What is the energy difference between the highest occupied electrons and the lowest unoccupied electrons?

That's what we're talking about.

So if that filling, if that distance-- OK, so here's an element, carbon, in the diamond structure.

And just-- purple is filled and not purple is not filled.

So you know that that's the connection band.

And that's the valence band.

And the band gap, which we've just defined here, is the energy difference between those.

And for diamond, it's really large.

For diamond, it's really large.

So you can see that, in fact, the gap is over-- if we converted this to electron-volts, is more than five electron-volts.

Now, I mentioned thermal energy in answer to the question about water.

And room temperature-- so room temperature, the energy of room temperature, if you think about temperature as an energy of like vibrational motion, and you can do that.

You use something called the Boltzmann constant and you can get what the room temperature energy is.

And it's something like 0.025 electron-volts.

Seems pretty small.

So the point is, if that's the energy, the thermal energy I have at room temperature, then if my bandgap is so much larger than that, is so much larger than that, it means those electrons are always going to be stuck.

They're never going to, even in some accident, have enough thermal energy to come up and be excited.

We'll talk about this later.

So they're really stuck.

And if an electron is in the valence band, it can't move very well.

It can't move-- it's kind of stuck in those bonds.

It's kind of stuck in those bonds.

And so here, if there is electrons all stuck here, and they never have a chance to go up here where they can travel freely, this is really going to be an insulator.

This is not going to carry electrons through the material.

So that's an insulator.

When you have a very large bandgap like that, we call that an insulator.

OK, but see if we compare-- and this is interesting, you're just going down in the periodic table if you compare diamond to silicon to germanium.

Well, see, what happens is the bandgap gets smaller.

I'm still filling them all the way up to the top of a band.

So I'm not a metal.

If I'm a metal, it's because my filling stopped somewhere in the middle of the band.

I'm not a metal.

I'm a semi conductor and not an insulator.

And the reason I'm a semi conductor is that my bandgap is kind of small enough so that electrons can kind of jump from here up to there, sometimes.

Sometimes.

That's why it's called a semi conductor.

They kind of conduct electrons.

So the electron can get up to the conduction band in a semiconductor more easily than an insulator.

And that's why it can conduct electricity much better than an insulator, many, many, many orders of magnitude.

But the way that we have talked about getting electrons excited, getting electrons excited, is with light.

And we can do that for solids, too.

We can do the same thing for solids.

And so if I plot the absorption, if I take a solid and I shine a light on it, there's something called the absorption coefficient, which is really just a measure of how far down into the solid did that light have to go before some electron got excited by it?

How well does it-- well, OK, different solids can absorb electrons differently.

But the main point is, if I plot that versus energy for something like silicon, the bandgap of silicon is 1.1 ev, so that's what it would look like for silicon because you know-- this should be wavelength.

Because you know-- because as you get to lower-- so you know that above-- it's exactly like for atoms.

If I shine light, like remember the Bohr model, if I shine light on an atom that has enough energy to promote it up, then it can go.

But the difference here is that for silicon, for silicon, that's right there, for silicon there's this big empty space with no bands.

It's just like the atom.

When there's no states, I cannot have an electron there.

I'm stuck.

So the electron can be here or it can be here.

And so the light, the energy of the light, has to be at least the bandgap to get it up.

At least the bandgap to get it up.

So what that means is that if I had-- so, why does this go to zero with increasing wavelength?

What it means is that if my valence band looks like this and my conduction band looks like this, and now my light is coming and I want to take an electron and promote it up.

Well, it cannot be-- the electron cannot exist inside the band gap.

There is no states there.

It's just like we talked about with atoms.

But now we've got almost infinite states here, no states there.

And then again we start with infinite places where it can be.

What that means is that if I shine light on it, the only way silicon, the only way it can absorb this light is if the light has energy greater than the band gap, or wavelength less than the bandgap.

That's why I drew it that way.

So it goes to zero-- so it's going to absorb in wavelength up to a certain point.

And then if I go higher in wavelength, those are energies that are too small.

It's going to try to absorb here but there's nowhere to absorb to, so it can't.

But if it has enough energy, oh, look at this.

This is also different.

I could absorb here or here or here.

I can take all those higher energies and absorb light kind of continuously, unlike in the atom where it could only be the level to hop you from n equals 2 to n equals 4.

But here I have an almost continuous set of states, except in the band gap, where I have no states.

So that's what the absorption would look like for silicon.

And if you look at the bandgap, this very, very important quantity, there is carbon.

There's diamond.

There's silicon.

There's germanium.

And you can see that as you go to heavier and heavier elements, you change the gap.

And you can understand this from what you know about the electron levels in atoms.

Because in an atom, remember, if I get heavier and heavier, those electrons all the way on the outer edges of the atom, they're more weakly bonded to the atom.

They're weaker and weaker and weaker.

So the energy is higher and higher.

Remember, so instead of being these deeper energy levels that have more bonding, more energy in them, lower in energy, they're less and less energy.

So you can imagine, now, that those are what you're starting with.

Those are what you're starting with, these.

Right these are the atom levels that you're starting with.

So if you're going to make your atom have electrons that are more weakly bound with energies that are higher, that aren't as deep and don't have as much distance between them in the energy landscape, then the gaps will get smaller.

Now, that is a very simple picture of why the gaps get smaller.

It happens to work for this column.

So if you look at this-- and some people would even call this a metal because that's such a small gap, but I like to call it a semiconductor.

But if your gap is zero, if you're band gap is zero, it means there is no gap.

So by definition, you're in the middle of a band.

You're somewhere in the middle of a band.

And if you're in the middle of a band, you're a metal.

You didn't fill that band all the way up.

But if you filled it up and then there's a little gap between you and the next level, then you're a semiconductor.

Or if it's much greater, like greater than three electron volts, 3 and 1/2 electron volts, then you're an insulator.

Now, so you can imagine now why-- so as the interactions get stronger, you can imagine those distances can get larger.

And so you could have larger and larger gaps, which can explain this trend.

But as I said before, where did I put it, it depends, not just on the atom and the original energy levels, but on how those bond together.

And we'll talk about that on Wednesday.

And it also depends on the structure, which we'll talk about next week.

So it's not just as simple as that but you can understand basic trends this way for some of the columns.

OK, we'll get to that on Friday.

Now, the goodie bag gives you a way to probe this.

So what's in the goodie bag?

The goodie bag has a way of measuring conductivity.

Now, I just told you that you cannot get connectivity unless you have electrons that are able to make this jump, to make this jump.

They've got to be able to go up from the valence band to the conduction band.

If they don't do that, they're not free.

They're stuck in these bonds.

You've got to get them free so they can roam around and not be stuck in this bond or that bond.

This measures that.

This measures that.

This measures whether electrons are in the conduction band.

You've also got some LEDs and some LEDs.

And this is a critical piece here.

Why do you have two kinds of LEDs?

Because on the one hand, what you're doing, I'm giving you a way to excite electrons.

So I'm giving you a way to actually promote electrons.

So you've got 10 to the 24 electrons in this valence band.

This would be the vb and this would be the cb.

And what you're doing is you're shining light with different frequencies, which means different energies.

That's what the right hand side gives you.

It gives you different energies of the photons.

So you have a way, now, of shining energy on a semiconductor.

And that's the second piece.

The semiconductor is what an LED is.

It's a semiconductor.

If LEDs were insulators, they wouldn't work.

And if they were metals, they wouldn't work.

Because the way an LED works is actually just the opposite.

Here, I'm shining light so that an electron gets promoted and I can measure that with my voltmeter.

But I could also put electrons in with a current.

That gets them up here and then they drop down and emit a photon.

That's what an LED is.

That's what an LED is, it's just the reverse.

And so what I've given you on the left are different semiconductors.

Because you see, just like a different-- you know, if I'm running current through like I am on the right, you see, those are different colors.

That means that I must have different gaps.

You know that.

You know that already from the work we did with Bohr and the color or frequency of light that you can get out of different transitions.

But it's the same here.

Did an electron-- you know, did I put an electron up here and then it fell down and gave off some lambda?

That lambda has to equal the difference between the-- it just equals the band gap.

It equals the band gap because the electron was put up to here in the cbm and then it drops down to the vbm and it emits a photon.

That's what's happening on the right hand side.

But see, you've also got, on the left hand side, you've got sensors because it's just LEDs on the left.

But if I hook those up, if I hook those up to the volt meter, well, I'm hooking up a semiconductor to a voltmeter.

So I'm hooking this up and I'm reading, is there any current.

And the answer in a dark room is no, unless you make it really, really hot.

Don't, because you have to go to very high temperatures.

Well, what if I shine light on this now?

Well, if I shine light that has the right energy to put electrons up in the cbm from the vbm, then I'll read something on the voltmeter.

So by giving you different LEDs on the left, I'm giving you different band gaps.

You have in your hands different semiconductor band gaps.

And by reading whether there are electrons in here or not, you can tell what frequency of light I shined on them.

So they're both emitters of photons and detectors of photons.

And it all comes together because you have a semiconductor, which happens because you took 10 to the 24 states, you put them together, you had gaps between them, and you filled all the way up to one of those gaps, and that gap was around 1 or 2 electron volts.

We will pick this up on Friday and go further-- I mean on Wednesday, on Wednesday.