

Let's get started.

How is everyone?

[CLASS RESPONDS]

Whoo.

You know why you're so excited?

Because it's a goody bag day.

And I'm not going to deny you a weekend of goodie bag action, which is why you're getting it today.

Even though the topic in the goody bag is crystallography, we will not start talking about crystallography until Monday.

But I didn't want to wait.

I know you didn't want to wait to have these models in your hands because it's an enhancement of your weekend activities.

So the goody bag, again, just sort of like last time, I'm giving it out to you today, but the topic will start on Monday.

Today, instead, we're going to talk about metals and metallic bonding, and the properties of metals that are related to the kind of bond that you get in metals.

And you know, we kind of motivated band theory of solids with conductivity and we talked about conductivity.

There is a chart of conductivity and there's some elements in there, some elements and some materials that are more complicated, right, like glass isn't necessarily one element or silica.

There's diamond, silicon, germanium, copper.

And we talked about insulators, semiconductors.

And I've mentioned metals and that's what I want to focus on today.

We did this in the context of the bands, right?

So remember, here you've got your conduction band, and here you've got your valence band, and the valence band is filled.

And so if there's a large gap, that's the gap, if it's large like, you know, over 3 and 1/2 eV then it's an insulator.

And then the one in between we talked about a lot on Wednesday that, you know, if this is your gap, and it's say between 0 to 3.5 eV, then that's a semiconductor.

And so, you know, that's what we talked about.

That's in your goody bag.

We talked a lot about semiconductors.

And then we talked about how, well, if you have some kind of band that is filled in the middle there.

So it doesn't have to be filled in the middle.

It can be filled here.

It can be filled up here.

But the point is that the filling of electrons does not stop where there is a forbidden zone of states, right, where there is no states.

That's what this is.

That's what a gap is.

It's a gap of electron freedom, right.

Where there's metal, there is no such thing because you fill up to a point where then right above that infinitesimally tiny bit of energy away, there is a free state right?

There's a free state, so you're in the middle of a band and that's a metal.

Well, I mean, this is you know-- so we should say that the gap here has to be greater than zero, here the gap would be equal to zero.

There is no gap.

You can't define a gap in a metal.

And we talked about semi-conductor problems.

I looked up, just for fun, you know, I mean, why not?

Because what else do you do late at night?

You look up old exam problems and you play with them.

And you know, OK, so from last exam, all right, so this is for the semiconductors, we had-- oh, here's a good question.

2.8 centimeters cubed of pure germanium.

Now I want to highlight this because there's one more thing I want to say about semiconductors before I move on.

Now, it's doped.

And that's what we spent a lot of time talking about on Wednesday.

It's doped with 50 micrograms, OK, of selenium.

That was the problem.

And the first part of the problem is, how many carriers are produced?

And in the same question, it's like, are they n-type or are they p-type?

What type of carriers?

Those are things we can answer based on what we learned on Wednesday.

And so, you know, oh, then you have this.

And you say, aha, I know where to start because everything starts here, right?

And so pure germanium is doped.

Well, first of all, where am I?

This is my map, right?

Germanium is there.

Now, it's doped with selenium.

Is selenium to the right or to the left of the period?

It's to the right.

Right?

And so right away you know, if you put a selenium in there instead of germanium, so if you take a germanium atom and you put a selenium atom in, it's going to-- if it's a doping, it's going to be in the bonding network that germanium was, but add electrons to it, if it's to the right.

Right?

And so you know that it's going to be n-type material.

And if you do the math, the number of carriers is equal to 7.6 times 10 to the 20th per centimeter cubed.

Now, here's something important.

That is twice as many selenium atoms, twice as many because it's two over from germanium.

So it had two extra electrons for each dopant atom, right?

So the way you solve these problems, you figure out how many selenium atoms are there per centimeter cubed and how many electrons for each of those acted as a free carrier dope.

As a dopant, remember, dopant simply give you carriers either of their-- if it's n-type, or down here, if it's p-type.

OK?

And this is n-type.

N-type.

OK.

So that's nice-- now the reason I want to do this is because, again, it is-- you can never have enough fun doing problems like this and looking up old exams and stuff.

But I also want to go back to this picture that I drew for you on Wednesday, OK?

So this picture was of the number of carriers, and remember these are electrons that carry current, right?

And we sort of drew something that looked like this.

Now remember, this is, right, these are thermally activated and this is thermally activated, and these are from dopants.

Now, if I put dopants in, and then I can get you a constant number of carriers because it's just defined by how much of the impurity I've mixed in.

How much impurity and I can control that really carefully, which means I can control the number of carriers, and then eventually, thermally excited carriers will just start to dominate.

But this ability to dope with chemistry is what enabled the semiconductor revolution.

Now, you say, OK, what's going on over here?

What's going on in this part here?

All right.

Well, you know what's going on there because we drew this picture on Wednesday.

Because if I have an n-type, if that's my VBN and that's my CBN, the way this works is you introduce a donor level.

Right?

That's how doping works, right?

The electrons from selenium go into here and then you have enough thermal energy to get them up to there and that's where they're free carriers.

That's where they can carry electricity.

And so at a certain temperature below which, right, at a certain temperature, if you're lower than that, well there's not even enough thermal energy to get from here to here.

All right.

Not even enough to do that.

So it's got to go down eventually.

That's what that part is.

And then the one other thing I want to say about this plot is that we have names for this.

These are called intrinsic carriers, intrinsic, and these are called extrinsic.

Why?

I mean, that kind of makes sense.

That's actually a pretty good way to name them.

Because, you know, intrinsic means I didn't need to do anything.

It's pure germanium, pure germanium, intrinsic.

I didn't do anything to the material.

But it's still-- you can get carriers from thermal activation above the gap.

Whereas if it's extrinsic, it means I did something external.

I added dopants.

Those are extrinsically doped.

So those are just important distinguishing factors.

Now, I'm going to come back to this plot later.

I want to talk about metals.

Metals are all the way over here.

Those are semiconductors and doping them or thermal activation gets me carriers, which gets me this connectivity range, but metals are much, much, much higher.

Metals are much, much, much higher.

Why?

What is going on in metals that's so special?

OK.

Well, the way we understand everything is from the bonds.

So what is happening in a metal?

There's aluminum.

There's aluminum.

There's aluminum.

And there is aluminum.

For us, this is all we see now.

We don't see any of this anymore.

We only see atoms and electrons because those are the building blocks, right?

Those are the building blocks.

And the thing is that, if we write down all the orbitals of an aluminum atom, there's the core, there's a shell, there's another shell, there's these ones out here, way out here, right, in the case of aluminum, three of them that are really, really, kind of weakly bonded.

Right?

They're way out there.

They're loose.

They have the ability to come off easily.

We talked about this in the context of other properties, right?

You know, the polarizability, the electronegativity, right?

All of this feeds in.

Everything is determined by where these electrons are.

And then also, as I mentioned, by how the atom's arranged.

That comes next week.

But if they're really, really kind of almost free, then you can kind of picture it like this.

Let's dial it back to sodium.

And you can kind of picture a metal as truly free electrons, right.

They are truly free.

They're not in-- they're in the middle of that band and they've got all this freedom because the states are free all around them, and they're basically not really that attached to any one atom or another.

And this is really the picture of the metallic bond.

So let's use-- let's go over here.

They're so free when these atoms come together with those very, very loosely bound outer electrons, those electrons are so free we actually call it a sea.

We call it the electron sea.

It's a simplified way to think about it, but it's a good picture to start with.

So in metals, you've got this electron sea.

Electron sea, we'll call it a model because it's a way of thinking.

And in this model, you've got a positive charge, nuclei, so you've got positively charged.

Why?

Because they lost an electron to the sea.



They've lost it.

Those outer ones have gone.

Well, they're not gone because there are others around too, but they're all just kind of roaming free, and so you're left with these positive charges, right?

And these are the cations.

Positively charged nuclei are cations.

Right.

And they're in a sea of electrons and so this is because the valence electrons is not tightly bound.

This is what I've been saying, but this is why you get it.

Cations, so positive charge nuclei, these are cat islands in a sea of electrons.

So if you think about a metal in this way, you understand a lot of the properties of metals.

This is a very useful way to think about metals.

And let's go back to our pictures of bonding.

Let me use this board here.

I'll put it right under here.

And we're going to finish what we started.

We're going to finish what we started.

Because, remember, we talked about covalent.

We talked about ionic.

And then we did all these intermolecular forces and the bonding that comes about from those, but we didn't do metallic.

So if I do covalent over here, and I do ionic over here.

Well, OK, covalent we know, right.

You have a maybe HH.

That's a covalent bond.

That's a non-polar kind of covalent bond, caring is sharing.

Or we could have HF, right, where it's sort of like, you know, it's a little less sharing.

But you know, it's still a covalent bond.

It's a polar covalent bond, right?

I mean, now in the ionic sense, you have sodium chloride.

We did that example where you literally-- chlorine is like give me and sodium's like OK and that's it.

That's the relationship.

Right.

But see now here, let's take sodium and let's just put a bunch of them around.

Well, but see, if I've got an electron sea, then my electrons are simply wandering around in the sea.

So every ion has this kind of ocean of electrons around it.

They're not attached.

It would be like if you say, how are you going to parent?

Well, I'm going to parent them until, you know, you can only go into the backyard and play.

That's it, right?

You are confined to be in the backyard.

And now you say, well, OK, but now if I'm a proton and I'm parenting an electron, I say, you know what?

You can be in any yard.

It doesn't matter.

Don't even come home.

[LAUGHTER]

I don't care.

Another kid's going to come into my backyard at some point, so it's all good.

That is the relationship of a metal inside of a metal.

Maybe here it's like the two houses switch kids.

I don't know what's going on.

No.

One of them just took the kid.

[LAUGHTER]

But anyway.

All right.

I'm going too far.

But here's the thing.

I want to give you that sense that they're just roaming around in a sea.

They're roaming around.

They're roaming around.

Now, here's the point.

I've got a sea of electrons-- oh.

Oh, huh, we can finish our table.

I'm going to finish that table.

We started the table.

Let's finish the table.

Because now, if I have metallic, all right, and I have this kind of, you know, this kind of set up here, right.

That's the picture.

That's the model.

And here I have cation-- remember our table-- slash delocalize, delocalize electrons.

And the bonding strength for this table in kilojoules per mole is 75 to 1,000.

It's a big range.

We'll talk about that.

And there is an example, iron.

We finished our table.

That's it.

That's the last entry.

Now, remember, this part of the table was the basis of attraction.

Remember that?

And we filled this out for all the other types of bonds, the basis of attraction.

I want to point out that every single basis we have covered involves the same thing, understanding what electrons are doing.

What is the electronic structure?

That's the fundamental principle of this class, right?

That's the key component, the key ingredient, that's in all of our recipes.

What are those electrons doing?

That leads us to every single one of those kinds of bonds, including metallic.

All right.

So now, you can imagine how this controls the properties, right?

But first you say, well, if I had a group one metal, if I had a group one metal, then I've got two-- I've got one, right, like sodium.

But if I go over to magnesium, you can see that I've got a higher charge because both of those electrons are likely to leave and wander around in this sea of electrons.

So that would be what magnesium looks like in terms of this electron sea model.

But the thing is that this sea, this picture, is what helps us understand the properties of metals and that's what I want to talk about.

That's what I want to talk about.

And these are some of the properties, right?

You can go through metals and there's a whole range of properties, but these are some of the ones that are pretty consistent.

They are reflective.

They're shiny.

There are literally settings in like photo apps.

Metallic, make it look metallic.

Right?

Because metals are shiny.

They have luster.

Why?

They have high electrical conductivity.

That's the thing we've talked about a lot already.

They have a high thermal conductivity.

I'll mention that.

High heat capacity.

They're malleable and ductile.

All of these and many more properties of metals can be explained from this one simple picture of the electron sea.

This one simple picture.

OK.

So let's start.

So if I think about-- let's look at these top three, OK?

So I'll start, why is a metal shiny?

Why is a metal shiny?

Well, it goes back to the thing we've been talking about all along, you know, or certainly in a lot of parts, right?

The bohr model, and atoms, and electrons getting excited.

The semiconductors and LEDs, right?

It has to do with how photons interact with the material.

But see, now I've got these electrons that are sitting out there at any energy level ready for the photon.

There's no gap.

Right?

So I've got electrons sitting around in the sea and they're ready for anything, any kind of energy that hits them, well, they've got room to be excited.

Right?

They've got room to be excited.

And it's that, combined with other things, but it's that, that is a big part of why metals have their shininess.

Those electrons can be excited in almost any way, right, because there's an infinite number of states right there.

There's no gap.

Right?

So the way they can respond to energy from photons is extremely diverse.

And you can imagine, if I'm a photon and I excite an electron, it doesn't matter what the wavelength is.

That electron can fall back down and re-emit the same energy photon.

OK.

So that's luster either way.

But now you know and you really should do this.

And look at.

I just said, you know, you have this.

I've got electrons.

I shine photons on these.

They might come up and go back down.

There's all this room here for electrons to be excited.

There's no gap.

Whereas in diamond, in diamond I've got this huge gap, right?

Remember, diamond is an insulator.

Diamond has a gap of 5.5 electron volts, roughly.

Now you know why diamonds are transparent.

It's of course, the electrons.

Of course.

Because if I shine photons on this, and those photons are indivisible, you know that even if I go all the way up to purple photons, they're not that high.

So any time I try to-- so I'm shining photons on my diamond and those electrons are like, can I go here?

No.

Can I go here?

No.

Can I go here?

No.

You have to shine a very high energy photon to excite an electron in diamond.

That's why it's transparent, right?

That's the reason.

When you go to the store to buy diamonds, you should ask them about the electrons.

That's what's making it all work.

That's what explains the properties.

But back to metals.



Now, OK, so that's luster.

What about electrical conductivity?

Well, it's the same thing.

It comes from the sea.

It comes from the sea.

If I take a wire of copper and I look at this as a sea of electrons, so I'm not drawing the ions, they're in there.

I'm just drawing the fact that I've got this sort of ocean of electrons that are fairly free, right?

Well, what is happening when I try to move them, it is with some kind of electromagnetic force, right?

There's some kind of force.

And what that means is that I have a charge that this doesn't like.

So like it's a repulsive charge over here.

It's going to push one of these into the other one.

But then that's going to feel-- that's going to feel-- a force is going to move and then that's going to push this one, and that's going to push that one, and so forth.

That's how electricity works, right?

So the actual movement of the electrons could be slow in the material, but the current, right, you can push a lot of current through because these ones over here feel pretty quickly the forces you applied over there.

But they only feel it, they only feel it if you can get this freedom.

If those electrons are not free, they're going to feel this repulsion there and be like, you know what?

I'm stuck in my state.

I'm stuck in this full valence band.

I got nowhere to go.

Sorry.

You know, put your energy into the atom or something else.

I don't know, but I'm not moving.

Ah, but if it's a sea, then you can push them and they can push each other.

Right?

That's the same thing.

It's a conceptual picture, but it's the same thing as almost an infinite number of free states.

That's what that means, right?

That's what that means.

Now, if you've got copper, we see copper now as something like argon with those d electrons filled, 3d 10, and an s electron here.

And so if it's a copper wire, those d electrons it's all filled up in there and they act as a way of shielding, it's a beautiful shielding, so that this is even more free and less interested in staying around.

So copper's such a great conductor and that whole column is such a great conductor and this is, by the way-- oh, such a good thing to do on a Friday night.

I'm so glad it's Friday.

Look at this, ptable.com.

Ha.

You could spend hours on this.

I do spend hours on this.

[LAUGHTER]

Property.

Electrical conductivity And then it shows you dark green is high-- dark green is high-- and then no coloring is low.

And then you can just click around.

You just click around.

I could do it now, but I would never leave.

But you go there and you can click on different properties and see the trends in the periodic table, and it's a beautiful thing.

And you now can understand them because of the ways in which electrons behave.

So you can see-- look at this.

There's copper, right?

So there's copper, and below that, we've got silver and gold.

And this is a really high set of elements.

The electrical conductivity in these three elements is super high, and this is why.

Because in each case, you're going down but you're filling the d's.

Oh, and then you're even filling f's.

Screening.

Further away.

Really free s-electrons.

High conductivity.

Electron sea, right?

Very easy to push them around.

Now.

OK.

This also goes back to this picture, because, remember, I snuck it in because it was on the graph, and I didn't take it out.

But there was a picture of silicon, where I showed silicon increasing its intrinsic carrier concentration with temperature, and on top of that was a metal.

I think was tungsten.

Now if you draw the metal, it would be way up.

It would be way up right here because this is higher, but it went this way.

But now you know why.

It's totally different.

The reason is totally different than why this changes, but you can understand it for the same root, which is the behavior of the electrons.

So, in a metal, in a semiconductor, this temperature dependence-- here, I've taken the temperature dependence away because I've doped it.

But here, that's the thermal activation into a place of freedom, into a place where they can be free in the conduction band.

Here, they're already totally free.

So what's happening?

Well, when I increase the temperature, I'm increasing the vibrations.

You can imagine if these electrons are trying to do their thing, and they're in a sea of electrons, the last thing I want to do if I'm moving along is to see some big ion moving around.

And if it moves more, then it's going to make me more upset and get in my way.

That's temperature.

That's why this goes down and not up.

For metals, the more thermal energy you put in, the more those ions-- those positive charges-- they're positive-- are going to interfere with the electrons being free.

Oh, we are understanding things.

We understood why these have-- now, you might say, why are these different?

That's complicated.

Relativity.

Not part of this class.

But if you want to understand why these are different from a-- they're all really high-- why are they different, it goes beyond the scope of this class, but it's also a very interesting story that has to do with, on the one hand, relativistic effects.

Now there's another thing, which is that if I go back to that table and I did this-- I really did this a lot-- I clicked back and forth between electrical-- now look up here-- thermal, electrical, thermal, electrical.

And you can click-- you can go, mm, mm, mm, mm, and you can watch the table, and do that for like five minutes or so.

And what you'll find is almost nothing changes.

You can't do that with very many properties.

But for metals, and those two properties, they're almost identical.

Again, it goes back to the sea.

It goes back to the sea, because if you have electrons that are free in a material, those can carry not just electricity, but those can be the dominant carriers of thermal energy as well.

And I'm not going to go into those details.

I just want to mention it.

So this is actually called the Wiedemann-Franz law, and it basically says for metals, that the thermal-- let's see.

OK, I'll just write it here.

Thermal divided by electrical conductivity for both-- conductivities-- OK, I'm putting it in parentheses.

It should be thermal conductivity divided by electrical conductivity goes as some constant times the temperature.

So if you look at that-- so you can look at it by toggling back and forth and see how similar these are.

You can also just plot stuff or find plots of stuff.

Here's one.

Electroconductivity-- this is at some fixed temperature-- orders of magnitude variation, orders of magnitude variation, tons and tons of different materials.

This is the Wiedemann-Franz law.

Now you know that law is simply a result of the fact that these free electrons can carry both electrical-- their charge-- as well as thermal energy, and that the way they carry those are dependent on the material, dependent on the number of electrons, and the way they bond and all that, but if they carry it in one way for one, it carries it the same way for the other.

And look at all these things.

Now notice here, this word that's used a lot-- alloy.

Alloy.

Alloy.

These are all-- alloy.

And I mentioned the word alloy on Wednesday, because I said if you wanted to make a good blue LED, you had gallium nitride but it was a little too high in its band gap, so they alloyed it with other-- then that meant mixing in.

Well, that is done for metals.

That is done for metals in a very wide range.

Why?

Because the metal properties that you have, you may like some of, but you may not like others.

So like for copper, it's got this really high conductivity, but maybe you don't want it to be so malleable.

What does that mean?

We'll see in 10 minutes.

Malleable, ductile.

So what do you do?

Well, you alloy it.

And so here's an example.

Brass.

Well, we know brass, but now you know brass.

Now you really know brass.

You didn't know brass.

You don't know anyone or anything until you see it as an atom and electrons.

That is how you know someone.

Copper?

With zinc substituted in a two-to-one ratio-- copper to zinc-- the connectivity drops.

Those electrons are not as free, but they're still pretty free.

Maybe they're a little more tied up in the bonding, and maybe what you've done by alloying has changed its mechanical properties, because I kind of want something copper-like but I wanted to have different mechanical properties, so I alloyed it.

And this is really how we think about materials design.

I need something to have, let's say, a certain set of properties.

Well, I don't have them all here, and I don't have them all here, but how can I use a combination?

How can I use a combination to get what I want?

Bronze is copper with tin-- four-to-one ratio of copper and tin.

The conductivity drop even more.

It's not so simple.

It's not so simple as, oh, if it's four-to-one, then maybe it's like 20%.

No, because it has to do with what happens to the freedom of these electrons.

That's what it has to do with, and that's more complicated.

But maybe I've changed the properties in another way that I really like.

Well, those are substitutional.

Substitutional alloys means I took one of the coppers out and I replaced it with something else.

Here's brass with zinc in there.

You can also have interstitial alloys, and interstitial alloys-- so if we-- we'll be coming back to this when we talk about crystals next week-- but alloys-- actually when we talk about defects.

So you can have substitutional-- substitutional-- and you can have interstitial.

This is just a preview.

Interstitial.

This is just a preview.

We're not going to be testing on this or thinking about this much until we get to actually putting these things in to a crystal later as defects.

But right now, I just want to make sure you see them for the first time.

Substitution, interstitial.



Why would you want one or the other?

Well, it depends on what properties you're trying to change.

In this case, I really like iron.

It's really cheap.

It's a cool material, but I need it to be stronger.

And that's called steel.

And if you just put a little bit of carbon-- really, not much, like a percent by weight-- it can dramatically change the mechanical properties of iron.

You can make it a much stronger material.

There is a massive, massive range of properties.

And part of why metals are so tunable is because of the electron sea.

I just came across this article actually, and I thought I'd show you.

This was just literally yesterday.

Was that yesterday?

I thought it was an interesting example.

Car Wars: Steel Getting Stronger, Lighter to Curb Aluminum Rise.

Well, see the thing is that as you want cars to be lighter-- think fuel efficiency, or range, if it's electric-- you've got to use lighter metals.

And aluminum is lighter, but aluminum is more than twice as much a steel.

It's a lot lighter though.

So there's a lot of interest in making steel-- that's just iron with carbon interstitial-- and making steel stronger.

And you can see, "Newest alloy." That's what caught my attention.

Newest alloy.

They've come up with something to add to iron to make its mechanical properties really, really good at a very, very thin layer because they need steel to be even stronger in a much thinner layer.

Otherwise, it's too heavy.

And you've got to make it really thin and really strong.

Those are going to be alloys.

OK.

Back to our property list.

Now, malleability and ductility.

Why are metals malleable and ductile?

Well, I thought this was a good place for a video because I really like it.

I think this is a really simple explanation of malleability and ductility as it relates to the electron sea.

So let's watch this.

This looks like a minute long.

[VIDEO PLAYBACK]

- [INAUDIBLE] shatters.

And this is because the applied force pushes light ions close together.

They violently repel each other, breaking the crystal apart.

That's an ionic crystal.

- In contrast, if you hit a metal with a hammer, it doesn't break.

It just dents.

Metals are able to deform form in response to an applied force.

The mobile sea of electrons shields the cations from each other, preventing violent repulsion and allowing the metal to change shape.

The most malleable metal is gold.

A similar property to the malleability of metals is their ability to be pulled into long thin wires.

We call this ductility.

Ionic compounds are not ductile for the same reason they are not malleable in general.

If an ionic compound is forced into a long cylinder, it breaks apart because of the repulsion of like ions.

In contrast, a metal can be pulled into a long cylindrical shape because the cations can line up, shielded from each other, as the fluid-like like sea of electrons flows around them.

That's my favorite part.

- The most ductile metal-- [END PLAYBACK]

Fluid-like sea.

Oh.

The most ductile metal.

What is it?

We may never know.

We may-- it's platinum.

But see, now, OK.

OK.

Now that's a-- I just gave it away.

OK.

Ductile, malleable, sea of electrons.

Sea of electrons.

Because did you see them flowing around?

You're moving the atom-- that was like a dance thing.

That can be like a new dance move that you try out tonight, the sea of electrons dance move.

I'm just saying.

I'm not going to do it again.

But the sea of electrons is able to move around the ions as they're distorting because the metal bond is not so directional.

The metal bond is not so directional.

Because think about it.

What is holding the metal together?

It's these interactions between the house and the kids that are roaming all over the place.

Is the kid over there or over there?

It doesn't matter.

Just give me a little bit of negative charge so I can feel an attraction.

That's the metal-- that's why it is so able to deform in the way it does.

Because those atoms are moving around, and the electrons are just kind of there as this happy sea, and the bonding remains.

It doesn't break.

That's the key.

Otherwise, this thing would just crack.

So wire-pulling-- check this one out.

You can take a 10-centimeter by one-centimeter platinum rod, and you can spin that into a wire, and this is done-- this is a very thin case for effect.

But you can make a 0.0006-millimeter diameter wire, and you get 2,777 kilometers of wire.

Now that's because of ductility.

The only reason I can get from point A to point B there is because I'm not breaking this apart, and that's how thin I can get-- 27-millimeter diameter when I pull it.

And I thought that was a really good example of why this matters, because maybe you've noticed, but wires are really important to us.

We use them a lot.

And so, I you just saw in a cartoon with electrons.

Here is a cartoon that I thought was interesting that shows a wire-pulling setup.

So, now this is-- Die, reducing the diameter.

That's a die.

Further reduction can take place using extra dies.

The drawn wire is finally wound on a spool.

This transformation elongates the wire.

One meter of 5.5-millimeter wire rod can be drawn to 30 meters with one millimeter diameter or 484 meters with 0.25 millimeters diameter.

OK.

Now, that is a cartoon, and I think it's always good.

When you see cartoons like this, you don't know.

Go look up what the real deal is.

What does it really look like to draw wires?

And I found it pretty cool.

So this is an actual wire-drawing plant, and this is what it looks like.

Now, there's no audio here, but I just wanted to show you the magnitude.

Those are huge towers that are melting metals to give you this first rod, and then it goes through one after the other.

And look at it being pulled-- being pulled in at some diameter, and it comes out in a different diameter.

But I wanted to show you the scope and size of this.

This is just one of them.

There's a person, and so you get a sense of the size, and then you can also see something-- when you go visit these factories, and I have these factories, there's always a lot of this happening.

Why?

Things are getting really hot.

When you're spinning things like a wire at a very high velocity, things heat up, so you get water all over the place.

So there's water being sprayed in factories all over the place, because you've got to keep stuff cool, or there's water baths that you run these things through.

And then you've got to think about, well, does this thing interact with water?

How does the water play a role?

It all comes down to the electrons.

And once they've actually drawn the wire out over this huge factory, they can actually weave it back together.

So I found this.

This is the same factory.

Now these things here are doing the opposite.

So these are huge drums, about the size of a person, and each of these got a wire that was drawn.

And now, again, because of the ductility, you can also move the wires, play with the wires, and you can weave them together.

You can weave them right back together into a thread that looks like that, which we all have seen I think many, many times.

All of this processing, all of this manufacturability, has to do with the electron sea.

That's why we have wires OK.

Now, a couple more small points.

So we're going to go back to this picture because this is how we started.

And this is the 1-D-- remember that 1-D model that we used to introduce band theory-- and I want to go back to this, because what you can see is that if I think about this as bonding and anti-bonding orbitals all kind of mixed in here, then you can also get a sense, just like in the dimers that we make-- N<sub>2</sub>, O<sub>2</sub>, remember NO?

Bond order.

It counts how many electrons are in an anti-bonding orbital.

Just because it's in an anti-bonding orbital doesn't mean it's not there in the system.

It's the same with these metal bands.

Some of them may be good for bonding.

Some of them may not.

But just like in the dimers, the bond order is related to the strength of that bond.

Remember?

And in NO, it's like a bond order of 2 and 1/2, and then if I took an electron out, it actually got stronger.

The bonding got stronger.

It's the same thing.

You can think about this the same way in metals.

You have a band.

Some of these orbitals are not going to be bonding orbitals, and some of them are going to be much more about the bonding.

And so that's going to relate to things like boiling point.

Metals have boiling points too.

So if you look now, go back to the periodic table, remember the highest electrical conductivity was here.

Well, that had to do with that right there, with the position, with the filling of the orbitals to give me the freest electron I can imagine, after those d-orbitals are filled, that s.

But if I'm thinking about whether I filled anti-bonding orbitals or not, because now I'm thinking about the strength of the bonds, you see a different trend, and you see something that you would expect.

First of all, you see a huge variability, but you see that the ones that are strongest-- the highest boiling points-- are going to be the ones where you fill that band-- those bands-- roughly halfway.

And this is also why, in a metallic bond, in a metallic bond, you've got this enormous range.

So you've got-- it's the weakest for elements with near-empty, near empty, valence subshells, valence subshells, and I would say, like OK, caesium is an example, or near-full, like, for example, in mercury.

Now mercury is actually a liquid at room temperature because of this.

There it is.

There it is.

Why?

Because I-- remember the beryllium dimer.



I filled all these electrons in there, and they're causing this really weak bonding, this really weak bonding, because I had to put all of them into the anti-bonding orbitals.

I couldn't do this trick that right over here, right one column next to it, these guys can have tricks, where they-- now this is bonding strength.

This is not electroconductivity.

But here, they're filling the d, and then the s is free.

Not here.

Here they're all filled, bonding and anti-bonding.

No tricks.

Now, if you look at the melting temperature-- that's one of the strongest elements would be tungsten.

The strongest would be example tungsten, where you've got a range of melting-- the melting point of tungsten-- and this is not the same as the boiling point-- the melting point of tungsten is 3,680 degrees C.

That's a liquid at room temperature.

And the melting point of cesium is 28 degrees C. Look at that range.

And it all has to do with how these electrons are filling the metallic bands.

And more than that, because it also has to do with the way the atoms come together, and that is going to be the topic of next week.

Now that would have been a perfect line to end on because I said the words, "next week." I just want to make one more point.

I can't stop because there is one more thing I want to show very quickly in one minute.

And that is-- going back to this picture, if you're thinking about filling, and I've got like my 3s-- remember, I've got like my 3s-- and let's put 2p here and then 3s and 3p.

Now remember, for metals, for a lot of these metals, what ends up happening is what I showed you

before on Wednesday, which is that you get actually an overlap of this, so you get a band that would be the 3s/3p band.

Sodium, magnesium, aluminum-- these all come together like this to form a single band.

Now remember, I don't need you to know how to do this.

We would tell you if it does this or not.

And we talked about that.

I'm not going to ask you to know how these come together, do they form a continuous band or not?

But in this case, they do, for that row.

And so you know.

I've got eight electrons, eight-electron capacity in that band per atom.

That's the capacity of the band.

So now I want to talk about, how is this band filled?

Well, for sodium, I'm going to fill it 1/8th, because it's got just one electron that it's putting into that band.

And for magnesium, I'm going to fill it with two electrons.

And aluminum would be three.

And so on.

So it goes back to the same picture as before.

We're filling the bands, but now we've got 10 to the 23rd and 24th atoms, and so we think about filling the bands as what's the occupancy per atom you can put in there?

And so I could tell you if I had a mole of sodium, how many electrons I have in this band, right?

It's a mole.

And how did you fill it?

And that ties into what I just talked about, which is how that relates to bonding strength.

OK.

Now we can stop.

On Monday, we will start talking about crystals.

Have a very good weekend.