We left off last Wednesday, last week Wednesday.

So it's been awhile.

So I wanted just to pick up where we left off.

And then cover some new things today.

But where we left off was with this diagram, OK?

And, you know, I kind of went through it.

And I went through it sort of quickly.

But I want to make sure you understand it very well.

Because this is such an important kind of diagram.

And we'll see it again.

And you may see it again in other classes.

This is a diagram of attraction and repulsion coming together in a balance to form a bond.

So, you know, last Wednesday, just to remind you, we made our first bond.

All right, we made the ionic bond.

And I showed you how that worked.

Because, in this case, the attraction, the force of attraction is the Coulomb potential, right?

Remember that goes as minus a constant times-- I'll just write it down because why not.

You're using it in your goody bag.

All right, so the Coulomb energy here goes as minus some constant times the charges, the two charges, divided by the distance between those charges.

That's an energy.

It's an energy of attraction.

But, you know, so here's the thing though.

All right, so if you have a plus and a minus charge, those are each 1, right, charge of electron.

And then the distance between them, when it's really far this, energy is pretty weak.

So that's this green curve here.

But as it gets closer and closer, that energy gets lower and lower.

And remember, happiness is lower energy.

All right, these atoms want to find their happy place.

And they see each other and they're like, well, we can be happier if we're closer, until a certain thing happens.

Which is then, they've got similar charges like the electron shells, negative charge.

Negative charge, so kind of coming too close.

That's the forces of repulsion.

And so that nets in a total energy curve, which is this white line here, the green and the red, and then you got the white line there, which is the total energy.

Because if they came too close, then you're going to be on this thing.

And you're just gonna be-- you know, all those electrons are going to be overlapping each other and very unhappy.

So this is what we do.

And we said, OK, this is the bond of an atom and another atom, an ion and another ion, a cation and an anion.

We did all this Wednesday.

I'm just getting us back in the mood.

And what I told you is that that bond energy is related to the lattice energy, all right?

I mean, now, technically, the lattice energy is-- and this is something you're playing around with-- the lattice energy would be the energy that it takes to go from those ions, sodium-plus in the gas phase, right?

We put that little guy there.

Now, that is something that's going keep coming, too.

It's in a notation I've been sneaking in.

Gas phase, all right, plus chlorine ion gas phase, and they came together to make NaCl-- ran out of room-- solid.

But I put it underneath, solid sodium chloride.

That's salt. Now, the lattice energy is the energy that it takes to go from these ions to the solid.

And as you can see, well, here I just did two ions.

But the lattice energy is going to be related to that bond energy, all right?

OK?

So the lattice energy of the solid that you make is going to be [INAUDIBLE].

And ionic solids, there's a whole bunch of materials that form solids this way.

They're ionic solids.

And they have properties that are general.

They're not always these properties.

But for the most part, these are the properties of our ionic solids, right?

They tend to be solid at room temperature.

They tend to have high melting and boiling points.

They're often transparent in the visible.

They're mostly electrical insulators.

They're hard and brittle, soluble in water.

But in your goody bag, you're seeing that's not always true.

Because sometimes, well, most of the time there, but if this lattice energy is too high, then it may not even dissolve.

It may still be an-- it just may be some really, really strong ionic solid.

But these are the general properties of ionic solids.

We'll come back as we learn about other solids Wednesday to the differences in properties between different types of solids.

OK, so that's where we ended.

Now, to do this, to get to here, we had to make ions.

And so we've been talking about ions.

And we're going to keep talking about ions because that's a very important part of how atoms see each other, all right?

Am I seeing you as a neutral atom with all your electrons in your shell or am I seeing you with a charge missing or not?

That makes a big difference in how they approach each other and therefore how they bond, all right?

This is a critical part of what we need to learn this semester.

And so, you know, we showed this, which is sometimes atoms actually don't mind losing an electron, don't mind.

Don't mind is like happiness level which has to do with energy.

And so if we want to look at this more formally, we would look at, for example, this chart from [? Avril ?] which is the first ionization energy.

Now, that's something that we talked about.

So the first ionization energy, which is how much energy it takes to remove an electron from the outer shell, the most outer shell electron, the first one you would take away.

That's the first ionization energy.

All right, and we saw, again, getting us back in the mood.

Oh, I've got colors there.

So if you started, for example, let's just take a look at this.

OK, so here we are at lithium.

I'm going to start here.

And that's first ionization energy.

And this is just the atomic number.

So there's lithium.

And remember, we looked at this last week, lithium and beryllium.

OK, so if you look at that, we said so for lithium, you had three electrons in the core.

So there's lithium, right?

And there's the lithium.

Oh, this is so fun because now we know what to call this-- 1s2.

All right, and then we have another shell.

These are just pictorial.

The electron is not in an orbit, it's in an orbital.

It's not moving around like this.

That's our classical minds fighting.

And it's got one more electron.

Why did I draw a dash?

It's got one more electron there in the 2s shell.

And that's lithium.

But now, we're going to go from lithium to beryllium and we talked about this.

So beryllium now is going to add, OK, it's going to add another proton.

It's got all of them.

But I'm showing the one that was added.

And it's also going to add an electron here.

But see, for beryllium this electron is really not shielded any more, not shielded more for the most part.

You know, it went into the same shell, the same N, all right, and the same L, same shape, same quantum number.

So there's nothing else shielding it.

And yet, there is a whole other proton.

And that's why beryllium is smaller.

This is all stuff we talked about Wednesday.

But see, look, that's also why it's harder to pull the outer electron off of beryllium-- there it is-- than lithium.

So those really go together.

All right, all the radius stuff we talked about holds, you know, that same kind of logic in terms of how these electrons are with respect to their nucleus holds for ionization energy.

Now, let's go to boron.

So if we go to boron, now, we're going over, look at what happened.

Look at that.

We went from there to there.

And the ionization energy got lower.

Well, that's also something we can understand.

Because now, we're putting it into the 2p.

OK, so if I do that, it would be, I'm going to put it under here, 2s2, right?

And now, if I go to boron, I'm going to put it here in a 2p shell.

So I'm starting a new L. Now, remember, that means the quantum number is still 2, but the shape of the orbital is different.

And now, we know also from something we learned last week that because of orbital penetration, the energies of the 2p and 2s orbitals are not the same.

All right, the s-energies are a little lower.

And it lets those electrons get a little bit closer to that positive charge that they want so badly.

So we know that this is going to be a little further out.

And so there's a little bit of shielding, a little shielding that you get for this, plus orbital penetration that we talked about.

And that puts it further out, all right?

And therefore, the attraction is a little less.

And therefore, the ionization energy went down.

And now, we're going to fill those up.

But it's the same thing as before.

Now, you're not going further out, you're just adding and you're not really screening while you add.

And so you're adding electrons into this 2p orbital.

But you're also adding all the positive charge.

Salute.

And so, the ionization energy is going to go up, except look at that, half filling, right there, half filling, a little extra stability.

Remember, we talked about exceptions, right?

Half filling.

Half filling.

And so, you see, now you can see, we can understand these trends that were being measured.

All right, and at the time, these ionization potentials, but now we really can understand them.

And you can group them.

And you can see if you group them, remember, we also grouped the periodic table a little bit by electron filling, quantum numbers.

All right, you had the s and p blocks.

Those are also often called the main group elements.

They're in the first couple of rows of those.

They're the most abundant elements.

But those trends, now, you fully understand from the quantum mechanical solution to the atom, all right, and the electron filling of those orbitals.

We can actually really understand it all.

And then the same thing happens with the d.

But it's a little more complicated.

By the way, these are often called transition metals.

The d-block, I told you about the d-block last week.

They're also called transition metals.

And the word metal is something we'll talk a lot about in a few weeks when we make metals.

But they're called transitions because, literally, they transition in the filling.

They transition.

You're filling d-orbitals, all right, transitioning.

And that makes it a little more complex in terms of what happens.

But the same trends hold.

The same trend holds.

As you add more d-electrons, the ionization energy keeps going up.

OK, now, this is extremely important, this ionization energy.

The first ionization energy is very important because it tells you so much about the outer electron, whether an atom is going to be an anion or a cation.

But all of the ionization, you could throw a whole lot of energy at these atoms and they'll all come off, all right.

And, you know, so we played with this kind of energy.

This is visible light.

You had your goody bag, too.

You got a spectrometer.

You can see lines in here.

This is a continuous spectrum.

But you see, you've got this whole, all these different wavelengths, right?

And now you see this and you don't just think about colors, you think about energies, all right?

And so if you think that way, because I want to know how can I measure these things?

Ah, stop, better.

All right, and so, you know, if I look at energies, energies, I've got like the visible, which is around, oh, I don't know, say 2 to 6 EV.

That's pretty high energy.

And then you've got, well-- now, can I shine visible light and kick out, you know, a lithium electron?

Well, the answer is roughly yes.

You know, the lithium energy ionization, well, let's just go back and look.

That first ionization energy, 520 kilojoules per mole, I know that's 5.4 EV-ish.

Because I don't mind going back and forth, right?

Kilojoules per mole of electrons, Jules, EV, moles, you can go back and forth.

But that's right around here.

So I could shine visible light and potentially knock that out.

But see if I wanted the 1s electron of lithium, well, it's much, much lower in energy.

Because, remember, we talked about this.

It goes in and in each time I add charge.

And so it would take 122 electron volts to knock that out, right?

So but we have that kind of light, all right.

The UV would be something like, oh, 10 to 100.

And then, if we really want to blast these, we could go x-rays.

Oh, are we going to have fun with X-rays later in the semester.

And this would go to something like 100,000.

And so I got it.

You know, if I wanted to, I could shine light and ionize the whole thing.

And that is, in fact, one of the single most important experiments we do to characterize materials.

And so I want you to know about it.

It's called photoelectron spectroscopy, photoelectron spectroscopy.

And, for short, we will not write that out ever again.

I will write out PES, photoelectron spectroscopy.

Now, this is a characterization tool.

I'm basically saying, well, you know, I need energy.

I'm going to get it from some photon source.

But I'm going to get enough to just blast all the electrons out of this material, out of this atom.

Because, for now, our material, they're just atoms.

And I want to know about these atoms.

So let's take a look.

So what happens when I do this?

Well, let's see, first of all, I've got some energy that we know is h nu.

It's dependent on the frequency.

There are the frequencies.

There they are.

No, there they are.

OK, so now, I've got some atom here.

And what do I do?

Well, I look, did anything happen?

No.

What about now?

I don't know.

OK, what about now?

Oh, all of a sudden, they come flying out, electrons.

All right?

And I measure their energy.

It goes right back to the photoelectric effect that Einstein was doing.

He was shining visible-ish light onto metals and seeing what electron-- but now I'm shining all sorts of light onto an atom and I want them all to come out, not just the outer one.

And so we will.

You can see if I measure, so now I'm going to measure the kinetic energy of these.

And the ionization energy of that electron is going to equal whatever the energy of the incident photon was minus the kinetic energy of the electron.

That's what I want.

I want that ionization energy.

This is, just to be clear, ionization energy.

And like I said, I want it all.

Not just the first, I want the second and the third.

So now, you see.

So this is the experiment that you do.

What are the results?

Well, the results are really just what we've been doing.

But you turn it over.

So let's look at hydrogen, OK?

Now, what have we been doing?

What do I mean by that?

Well, what we've been doing is we've been going like this.

What's that?

OK, this is hydrogen 1s1.

I could do that or I could just label it the 1s and then you see the one electron right there.

Now, if I shine light on this and I measure, the way that you do a PES for hydrogen is you would look at the ionization energy on the x-axis.

And this would be the relative electron count on the y-axis.

So you're just counting, all right?

And what you would see in hydrogen is there would be a peak.

It would look like this.

That's it.

If all the electrons in hydrogen, 1, are in their ground state, 1s, then you would see a peak and this would be, you know that it would be at minus 13.6, right?

That would be the energy that it takes to ionize it.

And this would be the 1s peak.

And that's what you measure.

Now, it gets more fun, because now I can also look at other elements.

So let's draw helium underneath it.

I'm gonna try to squeeze it in.

If I look at helium, well, helium is also 1s.

And its PES would be-- so this is now ionization energy.

This is the PES plot.

And the PES for helium would be also the 1s.

But the relative counts would be twice as high.

So the peak of helium would be twice as high as the peak of hydrogen.

Because, you know, for the same number of photons in, I'm getting that many more.

I've got both these electrons coming out.

So the relative peaks, I've got two more that can come out.

Now, if I go to-- OK, now, it gets really interesting, because now, oh, I even have room, which is making me very happy.

Because now I can do lithium right here.

And if you look at lithium, lithium goes like this, all right?

And if I look at a PES plot of lithium, the ionization energy and the relative electron count, OK, I'm going to have those 1s electrons filled.

And there's going to be another energy where the 2s electrons are.

So this is energy here.

These are the orbitals.

And somewhere over here is going to be a peak that's exactly half as high, exactly.

I mean, that's makes a lot of sense because there's twice as many electrons.

It's the relative electron count in the atom.

There's twice as many electrons coming out of that 1s orbital and the 2s.

So I'm going to get a half as high peak here, all right.

And I just told you, the lithium atom, this electron is 5 point something, 4-ish electron volts to remove.

This is 122.

So literally, for lithium, this would be 122.

This would be 5.4.

And you know, that's the ionization energy in EV.

And what you would do is you would draw the axis like that.

So you don't have to put this all the way over here, all right?

And all this means is I've broken up the axis and it's continuing to count here in a different scale, all right?

These PES plots are absolutely essential in understanding atoms and materials.

Because look at what you get.

You get, literally, the electron filling plot turned over.

It's incredible.

You get it turned over.

And you can write this as, oh, let's just do this.

You know, I could write this as 1s2.

We have so many options now, 2s1.

All right, I could write this as helium 2s1.

That would be like using the noble gas.

I could make what people, you know, there are so many-- I can make box plots, oh, box plots.

You can do this.

Some people really like putting things in boxes, 1s, 2s, box plots.

So these are all meaning, these all say the same thing-- Lithium.

They all say lithium, all right?

Photoelectron spectra.

So I could ask you questions like this.

There's one.

Ionization energy, megajoules per mole, relative number of electrons, this element has a charge of 2-plus and the PES shown below.

What is it?

So, OK, I know, 1s2, 2s2, and since this is going up three times as high, that sure looks like 2p6, doesn't it?

That looks like 2p6.

Remember, it's this turned over on its side, all right, where the peak height corresponds to the filling.

That's what that experiment gives you, 2p6.

Well, then it must-- But then you have your periodic table, ah-ha.

Gesundheit.

Neon, it must be neon.

No, who said no?

No.

Don't just-- Yes, it can't be neon because of the question.

It's got a charge of 2-plus.

Oh, tricky, tricky, tricky.

Magnesium, it must be magnesium.

That is the PES, right?

But it's magnesium 2-plus because I said it had a charge.

Thank you, shout out, I appreciate that.

OK, so that's the power of the photoelectron spectra.

And we'll be using it in the class.

And this tells us about how electrons leave atoms.

It tells us about ionization energies, not just the first, which is the one all the way out here to the right on the PES plot, but all of them.

Because I shine enough energy light to get them all out.

Now, you can also-- so that's about losing electrons-- you can also gain them.

And so I mentioned this a little bit, but just for completeness, I want to come back to this.

You have ionization energy, which is how much energy it takes to pull an electron out.

You also have electron affinity.

Because some atoms also might want electrons.

And so when you look at those plots, it actually all makes sense again.

Look at this, some atoms like chlorine-- this is electron affinity by atom number-- some atoms like chlorine really want another electron.

Why?

Because it's got an incomplete shell.

It's so close.

It just needs one more electron to fill that outer shell.

And it wants it.

And that's the electron affinity is how much does it want it.

Now, want-- happiness, happiness-- lower energy.

That's what this is.

So if you're not going to lower your energy with another electron, then you're just saying no.

You just say no, 0.

I will not take an electron because if you give me one, my energy is going to go up.

And I will be a less happy atom.

Now, that makes sense, too, right?

Filled shells.

Nitrogen, nitrogen has got a half-filled shell.

"No" to that electron, I like my half-filled shell.

It gives me a little added stability, a little extra kick.

If you give me that electron, you know, I'm gonna not have that added stability.

I don't want it, 0.

All right, 0.

Losing electrons, gaining electrons, losing electrons, gaining electrons, why does it matter?

Why does it matter?

Well, we made a solid last week.

Today, the answer is obvious.

It matters because of Danish wind, obviously.

43%, in 2014, 43% of all the electrical energy in Denmark came from wind.

It's higher now.

But see, now, this is a quarter of the Danish wind.

I only gave you a two minute hemodialysis why this matters on Wednesday.

So I'm going extended today just a little bit.

Danish wind, people are very dependable.

So over a three-month period, this is how much we need, energy, electricity.

But look at the wind supply.

Sometimes it's really there for us.

Sometimes it's not at all there.

Sometimes it's predictable.

Sometimes it's not at all predictable.

With that much of your electricity coming from this type of unreliable resource, that is a huge challenge.

I mean, even if you go out to Arizona and you talk about solar subsidies, Arizona is sunny all the time.

No, not sunny all the time, most of the time, more than Boston.

But this is the sun in Arizona.

That's the power you're getting from it.

And look at this, these are just clouds.

Because Arizona does have clouds.

And they passed by and they blocked the sun.

Do you know what a nightmare this is for a grid operator?

If a lot of your customers get their energy in this way, you know, and all of a sudden, half of the your supply of energy just turns off, I mean, that's a huge problem.

And this is one of the most limiting factors for increasing, to a large extent, the amount of renewables we have on our grid.

And so, of course, I know a lot of you are thinking, well, just store it.

And that's what we need to do.

But it turns out that really one of the only ways we have to store energy at this large scale is pumped hydro.

And you see, we're pumping water up a hill.

When I have access, I pump water up a hill.

And then when I need, I roll it back down and I turn a turbine with it.

I'm literally just trading energy, potential energy.

And then I bring it back, kinetic energy.

And then I make electricity.

The problem is that, you know, well, as you can see, pumped hydro is going to be good where there's hydro, where there's water.

So that's limiting.

But it also is a very low areal density.

And there are also a lot of environmental challenges with making this work in a way that doesn't harm the environment.

So there are a lot of issues with scaling up pumped hydro, a lot of issues as a storage.

And so I ask, well, what else can we pump up hills?

And we know what the answer is because it's what we're talking about-- ions.

Where are they?

Ions, we can pump ions up hills.

There's my picture.

Look at that.

Ions, and you know what this is, it's a battery.

A battery is two different materials, two different metals, a and b, where one of them has an ion, I don't know, like lithium, for example, that can go back and forth.

And the electrolyte is this thing in between that only allows that ion through, OK?

Now, here's the thing.

So if I'm a metal and I lose a positively charged atom, well, then I got to stay neutral.

And the only way to stay neutral is to pump an electron out, all right?

So if I want to draw electrons out of this, that's fine as long as I draw ions out of it.

And then both of them can do work.

And then they come back and they roll down an energy hill.

That's what they're doing and going back and forth, shuttling back and forth, back and forth.

They're rolling down a hill, literally, of energy.

You know, when it's in one metal, it's higher in energy.

And then, when I, you know, plug my phone in, it rolls down that energy hill.

And as it does, it travels across to the other metal, gets lower in energy, and the electron has got to come around and do work for me.

Because otherwise, it wouldn't stay neutral.

That's what a battery is.

It's all about ions, all right?

It's all about ions.

It's like a ski lift.

I like to think of it, you know, it's like these ions are getting into a ski lift and they're just getting pumped up the mountain.

And then when you plug it in, they're ready to ski down and they go down.

And that's just cruising across this electrolyte.

And I don't know what the electron is in this analogy.

But that's all a battery is.

You plug it in and the ski lift takes it up when you charge it.

And now, you power your phone on and it rolls back down.

Well, see the thing is that batteries have seen a Moore's law of themselves.

If you look at the Moore's law for batteries, it doubles.

So this is the energy storage of batteries over the last 150-ish years.

It doubles every 60 years.

That's not a very good Moore's law.

But that's completely changed.

That's completely changed recently.

And the reason is all about ion shuttling materials and ion storing materials.

This is why this has happened.

This is why we've had a revolution in electrochemical energy storage, all right.

Because 150 years ago, we only made batteries out of about 10 different materials.

And today, there's well over 80 that are commercialized.

There's many hundreds in research labs.

What is making the difference is that those materials now allow the shuttling to happen.

It allows the shuttling to happen more easily, maybe faster.

And it allows more of them to be stored per volume.

It's all about the chemistry that houses those ions.

It's all about the chemistry of the ions.

So that's why this matters.

Now, we roll things down hills to power our world.

I told you this already in the first or second lecture.

We roll things down hills all the time, all right?

So like methane is the core ingredient in natural gas.

And what we do-- and we talked about combustion already-- is we light that methane on fire.

But see, the thing is that what nature has done is it has put all this stuff up the hill for us.

So over tens of millions of years, nature has pushed the chemistry up a hill.

That is literally what it's done in energy.

Because, you know, it's made it so that if I light this on fire, it will go down the hill.

It will roll down the hill and give off energy.

Nature has done that for us.

But see, we need to be able to do this as well as nature.

And so in the last minute of why this matters, I want to explain why this is so challenging.

We need to be able to match.

Because if you look at the energy per weight versus the energy per volume of a material, that's a very important metric.

We plot this all the time if you work on energy materials-- the volumetric versus the gravimetric energy density.

This is good, a lot of energy per weight and volume.

Well, look at where gasoline is.

I want to do some math here.

Because I think this is very important.

Because I want you to see, now, if I take one liter of gasoline, so if I take one liter of gasoline, now, the cost is around $1.

Well, it fluctuates, but, you know, in this country right around now it's about $1.

OK, now, if I look at how much energy is in that gasoline, the energy stored in those bonds that nature has pushed up a hill over millions of years, the energy stored is about 33 megajoules.

OK, good.

Now, I'm going to make a comparison.

One MIT professor, one MIT professor operates at around 60 watts.

Now, this is something like-- well, this is equal to 60 joules per second.

Now, I know some professors who can operate a little higher than that and maybe some a little lower.

But that's like the average-- 60 joules per second.

Now, the thing is, if I want to get 33 megajoules out of this professor, then I get, at this rate, 33 megajoules takes the prof 153 hours.

Now, here's why this matters.

Because 153 hours will cost about $1,530 at the MIT professor salary, which is around $10 per hour.

And, OK, but look at this.

I dug something out of the ground that nature spent 100 million years making, fine, OK.

And I burned it.

And I got it for a $1.

I got the same amount of energy that it would have taken me literally $1,530 to pay for it.

That's our challenge.

Now, you can plot other things here.

But look at this. ethanol, wood, OK, liquid hydrogen, we're still trying to compress it.

But look at this, batteries are way down there.

This is why this matters.

Because we're not even close to being done.

We're not even close.

See that great uptick in electrochemical storage.

We need another order or two of magnitude still in storing energy, efficiencies, costs, et cetera.

So that's my "why this matters" for today.

Now, back to ions and electrons and atoms.

What I want to take the last 15 minutes of class is to introduce a new way of looking at this.

And some of you may have seen the Lewis dots, OK?

I'm going to introduce them today.

And then we're going to make a whole bunch of molecular structures on Wednesday.

And the reason this is so important is that it gives us a sense of these outer electrons.

It gives us a sense of how those outer electrons look for the atoms, and then, very importantly, how they come together in bonds.

OK, how they come together in bonds.

So if I just very, very simply ask, you know, what does the Lewis dot structure look like?

Well, first of all, the number of dots is equal to the number of valence electrons.

And this is equal to the last digit of the group, so of the element group.

I'll show you that in a second.

OK, so there's Lewis.

Now, Lewis, you know, so there's one of his drawings in one of his original papers in the early 1900s.

He actually was a professor here, as well.

And then he went and did a lot of this work out in Berkeley.

And so they named a hall after him.

And his contribution was absolutely profound because in thinking about atoms with dots to represent electrons in the outer valence, it gave us a way to so easily think about bonding.

Now, we can think about ionic bonding for sure, which is what we've done.

But more importantly, what Lewis helps us with is to think about covalent bonding.

And that's the subject of Wednesday, which is a whole other type of bonding that we're going to talk about.

OK, so now, the last digit of the element group, OK, good.

So there it predicts number of bonds formed by most elements in their compounds, good.

And there is the dots.

Look at that example, fluorine, or using noble gas notation, helium, that really saved us time.

I didn't have to write 1s2.

Actually, this looks like four strokes of the pen.

1s2 would have been three.

I don't know, honestly.

But anyway, OK, there is fluorine.

And look, one dot?

No.

Two dots?

No, because you add a dot for the valence.

So fluorine has seven dots until you reach the valence, until you reach the valence.

That is the key of the Lewis picture.

And we talked about this before.

We talked about this before.

Because it showed there were three really important things that come out of the Lewis picture.

One is it tells us, as I've just told you, it tells us about bond formation.

It tells us about bond formation.

And it makes the assumption that the valence electrons are what matter for chemistry.

And this is critical.

And I've alluded to this before.

You now have a real sense of this, though.

From these energy diagrams, you have a sense of this.

Because look, if I want to rip an electron off of lithium, I got to spend 5.4 electron volts.

If I want to rip the next one off, I got to spend over 100 electron volts.

And we said, well, we'll do PES with all these energies.

It doesn't matter.

I'll get 'em all.

But what if I just want to say, well, what electrons are going to be available to participate in something like a bond?

Which ones?

It's going to be these ones.

It's not going to be the ones all the way in close to the core that don't want to be bothered and they've got so much energy.

Just to get their attention takes 100 EV.

And even that's not a-- But here, a little visible light and I got lithium talking to me.

Well, chemically, it's the same.

That's light, right?

These ones are the ones that want to and can participate in bonding.

The ones all the way down in energy, way, way, way low, they're inert.

They're inert, all right?

So most of chemistry happens in the valence.

And Lewis really nailed that.

So these dots are only valence electron dots.

And there's one more rule.

Now, OK, so there is the seven dots for fluorine.

And then, OK, sometimes you'll see this.

Oh, this is just the number of dots.

It's this classification of the columns, all right, the groups in a pair.

But, no, that is not OK, not standard.

Remember, the IUPAC, which throws the best, most lively parties and conferences, they decide on how to name elements.

And many, many other aspects of our lives are decided by them.

And they said, no, that's an old classification.

It's confusing.

So we're going to go with the standard, which is that we still count 1 to 18, 1 to 18.

1, 2-- s-block.

13 to 18-- p-block.

All right?

OK, that's cool.

3 to 12-- d-block.

We're going back and forth.

But you can just take the last digit there, IUPAC standard.

Even the periodic table talks about it.

It shows it.

It shows the old classification there.

And then it says, no, don't use it, it's confusing.

Well, then why do they show it?

Because some textbooks still show it.

So I want you to be aware of that.

OK, now, you can use Lewis, you can use this idea of only thinking about the valence to understand when atoms have similar chemistry, all right?

So if the valence of carbon is the same as the valence of silicon, you've bumped down one quantum number, but they both have 4 electrons out there in the valence.

In this case, it's 2.

In this case, it's 3, in terms of the quantum number.

But they both have those same kinds of and same number of electrons.

So you can maybe expect there to be similar chemistry.

It turns out, in this case, there is very different chemistry.

But we'll learn about that.

Carbon-- hybridisation.

Silicon-- no.

We'll get to that.

But still, this is how we see silicon now.

We don't see it this way with Lewis.

We see it this way.

This is how we see carbon.

And so I want you to start seeing elements this way.

OK, and so you can look at periodic tables and take images randomly found online with Lewis dots.

And you can start thinking of atoms as their Lewis-dot representation.

And like I said, on Wednesday we're going to make a whole bunch of structures, molecules, with Lewis dots and show how that teaches us, not only about how to draw the bonds, but, literally, about which bonding-- You know, if I have sodium and chlorine, it's pretty obvious.

If there is a bond, it's between sodium and chlorine, all right?

But as I go to more and more atoms and I'm trying to form molecules, it's not so obvious what is bonded to what and how.

All right, and thinking about this and thinking about the third part of what Lewis did that was so important is what allows us to draw a whole bunch of molecules-- Octet rule, which is that atoms like-- we now know "like" means lower energy-- to reach 8 electrons in their valence.

OK, well, we can already see this for the ionic bond that we've talked about.

And you can write this.

And I want you to write ionic bonds using Lewis.

But in this case, it's actually pretty straightforward, right?

If I had, for example-- let's get some more board space over here.

If I had, for example, an ionic bond between cesium and fluorine, so if I had cesium and fluorine, then the way I would think about this in the Lewis world would be this way and this way.

Because I've got 7 valence electrons in fluorine and 1 in cesium.

And then, if I write this as an ionic bond, the way I would write it is cesium-plus with fluorine having taken that electron.

And so you would write this still with the minus sign there.

That emphasizes.

It emphasizes what happened.

But I see my Lewis here.

You know, I see it there.

But I'm emphasizing that this is an ionic bond that's happened.

Because fluorine wasn't supposed to have 8, but then it did.

It's got a super-high electron affinity.

Cesium has a very low ionization energy of its first electron.

And so fluorine just took it.

It literally just took it.

I'll take that.

Thank you very much.

You know, I'm not being judgmental.

But ionic bonds are kind of, you know, the relationship in a bond-- and I don't want to judge-- but the relationship is not even.

It's not even.

And that's OK.

But fluorine said, take, take, give.

And cesium atom, OK, fine.

Now, on Wednesday when we do covalent bonds, it's not so-- then there's a lot more sharing that happens.

And that's why we can model it the way that I-- yes, I still left there-- we can model it that way.

Because it is like a charged atom and another charged atom, all right?

And it's because fluorine took it.

Fluorine took it.

It's OK.

Any kind of relationship, you know, they get along.

They make it work.

Now, what we can also do just in thinking about this, just in thinking about this, I can go to another example like calcium.

And if I think about calcium and fluorine, well, I know again fluorine likes to be f-minus.

But see, calcium, when I think about it as a Lewis atom, you know, calcium likes to lose both of those.

Calcium, often, because, remember, they want to get to the Octet rule.

They want to reach their 8 electrons.

But calcium has got 2, what do I do?

Either somebody gives me 6, not going to happen, or I can lose 2.

And so what you get is that calcium will go, you know, to Ca2-plus.

But fluorine is still just f-minus.

Right from this I know that the balance stoichiometry, that is to say how many calcium atoms I need to stabilize this whole ionic bond situation, is going to be 2, right?

I got to go to calcium F2.

And if I were going to write this with emphasizing the ionic bond, I might put brackets around just so it's really easy to see, 2-plus.

There's no dots left.

Because now I'm writing it in a way where I've emphasized that-- gesundheit-- that F took it.

It took it.

And it is minus.

And there's two of them, you see?

So that would be like thinking about an ionic bond, all right, but from the Lewis-dot vantage point where I really see that fluorine took the charge, OK?

OK, so I think-- oh, ho, you could go further.

You can go further.

I think we'll stop there.

And on Wednesday, like I said, we're going Lewis all the way and we're doing covalent molecules, covalent bonds.