

MITOCW | MIT3_091F18_lec07_300k

I want to start right where we left off, which is in filling atoms, filling electrons into atoms.

We've done a lot of work to know what it is that we're filling.

We had to go all the way to quantum to know what orbitals are, not just orbits.

Right.

And we talked about the four quantum numbers, right?

And so now we're going to use those, and we're going to use the exclusion principle from Pauli and a few other things.

And we're going to fill electrons into atoms.

So this is where we left off on Monday.

There they are, right?

Remember, you know, as I said on Monday, if we only had Bohr, what we would have is the left hand side here, n equals 1.

We would have just have n equals 1, 2, 3, 4, 5, but we wouldn't have any of this variation of l , right, like the p electrons, the d electrons.

That differentiation came in when we solved the equations of quantum mechanics, right, for an electron in feeling the potential of the proton.

And then the other thing that happens is these things don't line up necessarily exactly how you might imagine, because there are complicated effects, right?

It's not just a naive feeling.

Instead there's these effects like shielding.

Right, so electrons all the way out here in an atom, they don't necessarily see all the protons.

They're shielded.

And then to make matters even more complicated, these orbitals can wiggle in with nodes, right?

This is what we did Monday, with nodes all the way in so they can see-- some of these orbitals have a little bit of them that see the protons up close.

So we showed that with 2s.

That's called orbital penetration.

And that's why these things split up.

They have the same principle quantum number, n is 2, but this has a lower energy than the 2p.

But how is the electron moving within this orbital?

No.

It's not.

That's a classical way of thinking, right?

And it's so hard to stop thinking that way.

I know we want to think that way.

But it's not.

An electron is being in its orbital, because an orbital is a probability distribution.

All right?

So the electron, all we know from that orbital is that the electron has a probability of being here sometimes and a probability of being there other times.

That's what it tells us.

So we know that in that 2s orbital, sometimes it's really, really close to those protons, right?

That was that orbital penetration, the peak that you saw on Monday.

All right, so with all that knowledge just kind of getting us back in the mood here and our quantum numbers and our Pauli exclusion, we're ready to fill.

And the way that we're going to start is the way that most chemistry textbooks would start, which is

with a very simple rule, all right, called the aufbau principle.

And aufbau means-- so aufbau means filling up in German.

And basically the idea is that you fill-- you fill from the lowest levels, the lowest.

Those are also the ground state.

Right, the ground state would be the lowest, right?

That's the lowest energy, lowest levels, ground state level, energy, and up.

OK.

And you use what is called the $n + l$ rule.

The $n + l$ rule means that the ordering, the ordering of the orbital energies-- orbital energies-- and let's get a little more space here OK.

The ordering of the orbital energies increases with increasing $n + l$.

All right, you fill from the lowest energy up.

So as my $n + l$, those are the-- remember, that's the shape.

This is sort of related to the distance from the nucleus.

This is the principle quantum number.

This is related to the shape of the orbital.

Right l , and as that number increases, the energy, that's how you fill.

Oh, but there's the case where it might be the same.

All right, so we have to have another point here, which is when two orbitals-- I did it again.

I forgot my r .

Two orbitals have the same $n + l$, then the lower n has lower e .

OK, this is called the $n + l$ rule, but rules are made to be broken.

So as we're going to see, this is actually really only true maybe 80% of the time.

But we're going to use it, because it is a good framework, and it captures-- it gives us a way to start thinking about electron occupation and atoms.

And so if you take these two points, this filling up and the ordering of how you fill things, then you can draw this as a very simple picture.

This is from the textbook Averill, right, where what you do is you follow the arrows.

OK, so you go this arrow, I'm filling 1s, this arrow, then I go to 2s.

Right, it's not going anywhere.

So you go there, and you keep following.

OK, so 2s, 2p, so you'd fill 1s, 2s, then 2p, right, because of the $n + l$ rule is in effect.

And then, OK, 3s.

n is 3, l is 0.

That's $n + 2$ is 3.

2p, 2p, n is 2, l is 1, same value.

But now I go to this one here, and when they have the same value, the lower n has lower energy.

Lower n filled first.

Filled first, because the ordering is to go from lowest energy up.

OK.

All right, so that's our rule, which will be broken.

Now, how does this look?

Well, if we do this, we're going to get a little notation here and that's important, if we do this, and we just do a few atoms here, so for hydrogen, right, so for hydrogen we'd have-- that's the 1s orbital, and we're putting in one electron in it.

Now, the notation here is something that we're going to use a lot.

So let me explain this.

This is n .

This is l .

That's the l in, what, spectroscopic notation, right?

Remember we talked about that Monday, right, s , p , d .

OK.

And then this is the number of electrons.

And so if we keep going with this notation, then in the case of helium, so for helium you would have $1s^2$, because I fill a second electron into the $1s$ orbital.

Now, this is where Pauli comes in.

Right, Paul said, OK, enough.

You're done with that orbital, because I've got two electrons.

They do have different quantum numbers.

They've got the same n .

Right, they've got the same l .

Their m sub l 's are the same because there is no range here for s , right.

Remember m sub l for s can only be, you know, zero.

But they have different spins, up and down.

And so sometimes, what is also really useful, and we'll go back and forth between these pictures, is you go back to what we drew, you know, with the Bohr model, where we draw this in terms of energy levels.

And so if you drew this in terms of energy levels, this is what'd you see.

So you'd have like energy.

And remember, in the hydrogen atom, this is minus 13.6 electron volts.

And so if you wanted to draw this in this way, you would say, OK, there is one electron there, and we'll - remember, we said you can draw electron spin with arrows.

Right, whereas here, if we drew this, then we draw-- I'm not putting the energy there.

We don't know it, because Bohr 13.6 good for hydrogen. Right, but here-- oh, and by the way, this would be the 1s orbital.

This would be the 1s orbital.

But in this case, I've got two electrons, one up, one down.

And now, Pauli says you're done.

All right.

And so if you go to lithium, then you have $1s^2 2s^1$.

And so if you drew this in terms of an energy diagram, you'd have the 1s, and we're putting that, and then you have the 2s somewhere up higher where we put one electron.

Those are-- that's how you would see these atoms in terms of the filling of the electrons into their orbitals.

Right.

OK, good.

Now, OK, let's keep going.

So what we'll do is we'll do this here.

So let's see.

So beryllium would be $1s^2 2s^2$.

And boron would be $1s^2 2s^2 2p^1$.

OK, this is looking good here, right?

I'm really following aufbau now.

Oh, I went that way, then I went that way, and then I went up here, and I started on this arrow, which is just a way, graphically, of showing those rules.

Good.

Now, we get to carbon.

And now we have to think about this, right, because for carbon, I've got $1s^2 2s^2 2p^2$.

That is correct.

All right.

But if I think about this in terms of the energy, $1s^2$, $2s$, right, well we know from before, right, we know from before, like $2p$.

Where is $2p$?

There it is.

It's got six electrons that it can take.

Because now m sub l can vary from minus 1, 0, plus 1.

Remember, that corresponded to the three different orientations of the p orbital.

OK.

But that means that I've got three, so I can fill this with 2, 2, and then the question is, what do I do?

I don't know what to do.

Do I do this?

Is that right?

No, somebody says.

And that person knows that the answer is no, because that person knows Hund.

Where's Hund?

There he is.

Would have been more dramatic had he appeared, but there he is, because Hund came up with another rule, which also is broken, sometimes.

Not as much as aufbau.

You know, so Hund said look, when you have a case like this, electrons in the same p, so the same like orbital, sub shell.

OK, it's not the same sub shell, but it's the same n 2p.

Let's draw that.

It is the same sub shell.

I'm saying opposite things.

2p, 2p.

Shell, sub shell.

But there are different possibilities, and how you fill them according to Hund is you maximize what's called the multiplicity, which means that you want the electrons to come in with the same spin in different orbitals, in different p orbitals.

That's how you fill them.

Why?

Well, that has to do with more quantum mechanics, and it has to do with something called exchange energy that's not part of what you need to know about.

And I'm not going to teach it here.

But you know, you can think about it simply as the electrons want to spread out, because they repel each other in the same orbital.

Right.

They want a lot of things, electrons.

They want to be close to protons.

But they want to be kind of not close to each other.

They need their space.

That's not a lot of things.

It's two things.

But that's what they want.

And this allows them to maximize that and lower their energy.

And remember, lowering your energy means happiness for electrons.

So this is the preferred way to fill, right?

And that's what Hund's rule tells us.

Now, if you go to silicon, and you say, well, OK, what does silicon look like?

All right, I'm going to put silicon-- no, don't do that.

I'm going to put silicon right underneath carbon, because I think it's an interesting comparison.

All right, so silicon is OK-- $1s^2 2s^2 2p^6$.

They're all filled.

Right, and then you go to $3s^2$ and you go to $3p^2$.

And this looks very different than carbon, but actually, you can abbreviate the notation, and we very often do in these filling notation, and we do it this way.

We say, well, this has a neon core.

So you go to the nearest noble gas, not with more electrons, because they're not there, with less electrons.

You go to the nearest noble gas, and say well, OK, that's basically neon.

All right, this is neon.

And so you can write the notation is this-- it's the same atom, just a slightly abbreviated way of writing it, where I would write, OK, this is a neon core with $3s^2$ and $3p^2$.

Well, you can write that with a helium core, $2s^2 2p^2$.

And what happens is you wind up having a nice distinction here, because these are core electrons.

Remember, we're filling electrons here.

That's electrons.

And these are called valence electrons, and we'll be talking about valence chemistry for the whole semester and what these valence electrons do, right?

And in fact, these valence electrons, these ones on the outside closest to the out-- furthest away, these are the ones that have all of the chemistry that we care about in this class, chemistry.

It's so important that I went all caps.

I'm not shouting, but I went all caps.

Whereas the core electrons are kind of mostly inert.

They're mostly inert, chemically inert.

OK, and so sometimes it's very useful to write the notation this way to see what do you got in the valence, what's going on?

Right, what does this valence-- because then you start to see similarities.

The valence for lithium and the valence for sodium look really similar.

There's an outer s electron.

That's the quantum number, the principle quantum number is different.

But the valence looks very similar, one electron in an s orbital.

And that allows us to think intuitively about similarities between elements, because if they have the same valence, and the valence is responsible, not same, but similar, similar shapes of orbitals, different principal numbers, but if that's similar, then maybe the chemistry those elements do is also similar.

This is a helpful way to think about it.

OK.

I told you that rules are meant to be broken, especially when we're talking about $n + l$, and it happens all the time.

And there's two main reasons or important reasons, and then there's a whole bunch of other very complicated ones, that two I want you to know about, which are exceptions to this aufbau filling, OK, which are exceptions to that, come in the following situations.

That you get stability-- you get stability.

That means lower energy for the whole system.

The system lowers its energy and can be more stable when either an orbital is fully filled-- there it is, ns^2 , np^6 , nd^{10} .

You know those are fully filled, because that's how many electrons you can put in each one, or half filled exactly.

And so that's why in chromium if you predict with aufbau, you would get $4s^2 3d^4$, but actually you get $4s^1 3d^5$, because you can half fill both of those, and that adds to the stability.

In copper, the predicted one would be $4s^2 3d^9$.

But actually what you see is a fully filled d shell, which adds to stability and a half filled s shell.

So it's willing to make this trade off of a fully filled s shell for half filled.

That's not a big trade off to get the d, fully filled.

OK, now there's more complicated effects, all sorts of effects that also lead to violations of the $n + l$ rule.

I really hesitate to call it a rule, given that, you know, there's a lot of examples where it's not about

the half filling or the full filling.

It's about complicated interactions that have to do sometimes, all the way with things like relativity, relativistic effects, literally, of these electrons, which will not be on any quiz or exam, with also the complexities that happen with orbital penetration as you add more and more electrons and the levels are kind of closer together.

They're a lot harder to see clearly separate energies than in the first two, three rows, as you go down in the periodic table, those levels are closer and closer together.

So anomalies happen.

But they're not really-- I'm not sure we should call them anomalies, because these are all exceptions to aufbau.

So like I said, I want you to know aufbau, because it's a great way to start, but I also want you to know that sometimes things switch around because of the effects that we've been talking about.

I don't expect you to memorize all the exceptions to aufbau, OK?

But I do want you to know about these two key factors, which are half filling and fully filling leading to enhanced stability.

All right, good.

OK, so we got valence.

We got filling.

Now, right away, right away, all right-- so by the way, configurations are in here.

It's a beautiful thing.

When I look at this, I now can look at this line in here that shows-- that shows using this notation, the filling of electrons for every single atom, every single atom.

And it's a beautiful thing because it tells us so much about the periodic table.

It tells us so much about why the periodic table is what it is.

It was arranged.

Right, it was arranged because of things like the combination, remember Mendeleev, mass and properties.

But now we see why.

We see why.

It's incredible, because it goes back to counting.

It literally goes back to the quantum mechanical derivations of orbitals and the principles of filling them, and then the filling them themselves.

And you see, you know, why does the lanthanides and actinides, why did those things run for 14?

Because that's how many electrons go into the f orbitals.

Right, that's how many electrons go into 4f and 5f.

That's why these are 14.

Why are these two columns here?

Why do those have similar properties as you go down?

Right, I mentioned lithium and sodium, right.

Well, because of what we just said.

You're increasing the principle quantum number, but the valence chemistry is very similar.

They're all s.

And we even call them sometimes these blocks by the valence chemistry that's getting filled.

Right, so those first two columns are the s block.

Here over here, they're the p block.

It's not that these don't have p electrons in them.

It's that these are the electrons that are getting added in this part, in this region, whereas here,

you've got d electrons.

And now we know exactly why this is 10.

I mean, we know even more than this, because aufbau says, why is this here and not here?

Right, because of the ordering of the filling, of the energies of the filling.

These three d's come after 4s.

That's why they're here.

Right, that's why they belong here.

We now know so much more.

We have so much more insight into the ordering of the periodic table, because of electron filling.

And we'll be making those connections.

We'll be making those connections a lot.

Right, but I just wanted to kind of show you the power of seeing things in terms of filling.

The one connection I'll tell you about today has to do with diameter.

OK, so we'll be making connections all throughout the class between electron filling and properties.

Right, but let's start with a simple one, how big is an atom?

Well, now you can understand some things.

Right, so for example, if I plot here or if Averill plots it, and I take that from Averill and show it to you, the probability-- remember we plotted this before for 1s, 2s.

Now, what they plot here, OK, the 1s orbital for helium, that's got two electrons in it.

There it is right there, the 1s, and then the 2s and 2p combined.

Right, everything is sort of thrown together here.

They're not separated for neon and then argon.

So these are the first three noble gases, helium, neon, argon.

OK.

And a couple of things.

One thing you can see is that the s electrons in helium are furthest out.

And then the neon ones are closer in, and look at that.

The argon ones are even closer.

Why?

Why are the 1s electrons getting closer in as I add more electrons?

Right.

Well, it has to do with what we talked about Monday.

Those 1s electrons are not screened.

So what they see as you go from helium to neon to argon is more and more positive charge.

In fact, I've got the numbers here, if you look at the 1s electron energy, remember for hydrogen, 13.6.

Now remember, I didn't write a value there.

But there's two protons.

So that 1s electron for helium.

So if we take 1s electron energy, so much abbreviation.

For helium, it's 24.5 eV.

That's how much it would take.

So we can put minus, if you want to compare with this, right, this minus 13.6.

That's also how much it would take to ionize that 1s electron.

That's where it sits.

But if you look at neon, neon is minus 869.5 eV.

And argon is minus 3,206 eV.

That is a lot of energy.

That is a lot more energy.

That is much, much happier.

If energy-- if lower energy is happiness, those 1s electrons are super happy, and they get closer to all that positive charge.

But you can also understand why, like neon is larger than helium, and argon, even though these energies of the 1s are closer and closer, these are still going to be larger, because as you go down in a column, you add a whole shell.

Right, so I can't-- right, so remember, here's my plus charge.

Here's my minus charge.

I can't-- this is 1s.

Now I can't add anything in here.

I got to add stuff out here and go out.

Remember those principle quantum numbers, 1 to 2?

Right, we don't need the s here, just 1 to 2 to 3, they're going to push you further and further out.

We talked about that Monday.

So now the next electrons that come in in neon's case in 2s or 2p or in argon's case, 3s and 3p, those are going to come much further out.

That's why when you look at size, which is what's shown here, you see this trend.

All right.

So you see this trend that as you go down a column, here they are.

Here they are.

So this is the size in the periodic table, and these are calculated.

These are calculated.

Why are they calculated?

Because you can see, this should be pretty hard to measure.

Right, these are hard to measure unless the atoms are in like a solid, and they're bonded together.

So measured atomic radii are often taking half of a bond length.

But we can also calculate them.

And that's why these are calculated atomic radii.

So these have no bonding.

They're just isolated, and they're calculated and there's helium, neon, and argon.

OK, fine.

So we just understood that.

And we understood this arrow that the radii get larger as you go down the periodic table.

But what about going across?

Why?

I'm adding stuff to the atom.

I'm filling lithium here.

I'm adding an electron and from lithium there to beryllium.

How can it get smaller?

Right.

How can it get smaller?

So that has to do-- what did I do?

OK, I'm going over here.

All right.

So that has to do with the fact-- with the same things that we've been talking about.

It has to do with the fact that as I go from lithium.

So OK, let's do this.

If I have hydrogen, and I go to lithium, you know that it's going to be larger for the reasons I just said.

But if I go from lithium over to beryllium, it gets smaller.

And the reason is because of exactly this.

If I go from hydrogen to lithium, I add protons.

But I shield those protons with electrons.

OK, but if I am at lithium, let's see, lithium is here.

So I've got one, two, three.

And I've got an electron here.

Right, I've got two electrons in the 1s orbital, and then I've got one electron in the 2s orbital.

OK, so that's lithium.

But now I go to beryllium.

And if I'm beryllium, I've got four.

OK, and so now I've got those 2s electrons here.

But see-- I'm sorry, 1s.

But in the 2s orbital, I can put a second electron.

Right, and so what happens is those two electrons in the 2s orbital, they don't really screen each other.

They are screened by the 1s orbital.

So if I now-- because I went-- even though I added an electron, I also have a lot more charge here.

Right, and because these two don't screen each other, but these two do, right, these are going to be closer in, because they're more attracted.

They're only screened by these two electrons.

So the additional positive charge wins.

If I go down in quantum number, then I get a whole shell filled to screen.

But here, I'm just adding another electron to the same orbital to the 2s orbital.

But I've got a hold other proton, and the shielding, the screening is the same.

That's why beryllium is smaller than lithium, even though you've added an electron.

And it's why the trend goes like this, increasing that way, decreasing that way.

Now, I like dancing.

And so I like this dancing analogy, because I love thinking about atoms as dancing.

I told you about the disco and the periodic table.

But you can take that all the way to the atom.

So you know, you can imagine that you're an electron, and you see a proton.

OK, so dance pairs are happy.

Dancers are happy if they have more people to dance with.

And so if I'm hydrogen, and I'm an electron, and I see a proton, protons can only dance with electrons.

And then I say, OK, let's dance.

And you kind of start moving, you know, a little bit.

And but now, I'm going to go to helium.

So there's another dancer in the middle, and someone comes on, and says, oh OK, that's cool, right.

I can-- Now, either one of us can dance with either one of those.

All right, so I've got two different people.

I'm happier closer in.

We can get close, because there's choices for who you can dance with.

All right.

So that's why you get helium much smaller than hydrogen.

More people to dance with.

Now you're lithium, and you come along, and I'm a lithium electron here, and I'm like I want to dance with somebody.

Who's there?

And there's only one person, because these two people are totally occupied.

I only have one possibility for people to dance with if I'm this lithium electron.

Right, but now I come in, and I'm beryllium, and these two electrons, and they're both like, well, OK, you guys are blocking, whatever.

That's not cool.

But there's four people in the middle.

So there's always two or three people, right, that we can dance with.

That wasn't like a move.

I don't know what that was.

But I'm like I can dance with you or I can dance with you.

And more dancing choices, happier.

Happier, lower energy.

So that's how I like to see atoms.

OK, so that's one example of filling and screening and dancing, where this kind of picture of the atoms, this electron filling of the orbitals explains what we see.

OK, now, we talked about ionization, and we talked about how sometimes those outer electrons can be lost, or sometimes, maybe an atom would gain an outer electron.

If an electron is lost from an atom, called an ion.

And so a charged atom is an ion, and a cation is a positive charge, positive charge.

Abbreviations.

Right, and so here's sodium, and it's got all those dancers in there, and it has lost an electron.

There it is.

It's going to lose-- remember we talked about the first ionization potential.

We talked about that last week.

It's going to lose that lowest-- sorry, the highest energy electron, the least bound electron, the one that's all the way out there.

And so if it does, it's Na^+ .

OK, gesundheit.

Now, you can also-- this is from Averill.

You can also lose more than that.

All right, magnesium might lose two electrons.

So the gray part here is the ion, and the blue part is the original neutral atom, all right.

And you can see that if sodium loses one electron, and magnesium loses two, those actually have the same number of electrons.

They're different atoms.

Those are called isoelectronic.

Well, you can also gain electrons.

So if you write that down.

So you know, if you gain electrons, does anybody know what those are called?

Anion, OK, negative charge would be in an anion.

OK, and I can ask question.

If I lose an electron, it makes sense that I got smaller.

This is the radius.

And if I gain, it make sense that I got bigger.

And I can mix it up.

So I could ask you a question, like what's the ordering-- what's the ordering of these atoms.

Right, I could say, well, if I had Mg 2 plus and Ca 2 plus, and Ca, what's the ordering of their size?

You now know how to answer questions like this, right?

If I lose electrons, I'm going to get smaller.

If I gain electrons, I'm going to get bigger.

But I also know that for example, Ca 2 plus is smaller than Ca, and Ca is bigger than Mg.

Right.

And so Ca 2 plus is going to be bigger than Mg 2 plus.

Right, that makes a lot of sense.

Done.

You can order them now.

All right, so Ca must be bigger than Ca²⁺, and that must be bigger than Mg²⁺.

So we can think about now adding in ions to the mix.

And we're doing this for a reason.

All right, we're doing this for a reason.

But because, you see, some atoms really don't mind doing this.

They really don't.

Sodium is one of them.

They don't mind losing atoms.

I mean, yeah, the electron is there.

But it's kind of there as sort of like, I could be here, I can be there.

I'm good.

Whereas those 1s electrons, you know the 1s electrons are argon, are like you just try to take me out.

All right, no way.

But we will.

We will ionize those next week.

But for now, we're taking these outer ones out that may not be strongly balanced.

They're happy, losing or gaining.

And how do you know?

How do you know?

If you have ions, how do you know, you know, whether you have them?

Well, here's the deal, you now know a way to tell because in this goodie bag, which I'll talk about in a few minutes, you actually have the most accurate measuring device ever made.

You have a scale.

That's not-- no.

It's in here somewhere.

You have a conductivity meter, somewhere in here, which I cannot find.

And the conductivity meter is exactly that.

There it is.

Oh, look at this.

This is incredible.

So you can put this into a solution, and it literally just measures-- it's got two electrodes that go in there, and it measures whether there's any conduction, and you can see, well, if I've got charge species in the solution, ions, then this will tell you, because they'll help conduct electricity.

So you now have a way of knowing when you have ions or not.

OK, so that's good.

But you see, the thing is something very important happens when atoms come together-- when atoms-- don't measure anything yet.

But take it to lunch.

Take it to lunch.

You don't know.

People talk about electrolytes.

They talk a big game.

All right.

Now, the thing is when atoms react, when-- this sounds like the beginning of a novel.

When atoms react, right, sometimes they will gain or lose.

When atoms react, they may want to gain or lose the amount of electrons to get them to the closest noble gas.

OK, closest noble gas.

Now, that's really important.

We just talked about how you use the noble gas notation-- did I just erase it?

I did.

Because that could be chemically inert.

And we can only go down when we do the notation, but now we can go up, because I could add electrons to get to the nearest noble gas, in terms of reacting.

And look at this.

If sodium is an atom that doesn't mind losing electrons and chlorine is an atom that really wants to gain electrons, and they come together, and they come together, then you can form a bond.

Oh, oh boy.

Here we are I guaranteed you.

I told you on Monday, we would get here.

I said, it's our three week anniversary, and I think we're going to bond.

First bond, first bond.

And it's ionic.

It's ionic.

Because this is a bond of an electrostatic attraction.

This is the bond of electrostatic attraction between two oppositely charged atoms.

So between two oppositely charged atoms.

And so if I've got a positive and a negative, they're going to be attracted.

How are they going to be attracted?

Well, actually that's an attraction that we know all about.

Oh there it is.

Hold up.

We just made our first bond.

Let's talk about it.

What do I want to say?

I want to talk about that attraction.

Because that's actually attraction that we know very well.

It's a Coulomb attraction.

Right, and so the way this works is if I'm an atom with positive charge, and I'm an atom with negative charge.

All right, so let's take that same example, and this is the energy.

And this is say, zero.

And this is the distance between the two atoms.

If I'm out here, then maybe the distance between sodium and chlorine is very far.

But you see there's a Coulomb attraction.

And so the attraction is Coulomb, which is going to be minus a constant times the charges on those ions over the distance between them.

All right.

So that-- if I draw that attraction, this is the charges on ions, charges on ions.

And if I draw that, the Coulomb attraction looks like this.

That's it.

Right, that's 1 over r .

But see, the thing is if they get too close, then those core, you know, then things start to repel.

It's like, whoa, whoa.

Hold on.

I'm not going to go any further, because I don't like having the same charges getting too close, right.

I don't like that.

Your electrons are getting too close to my electrons, and don't make me go nuclear.

And so what you get is you get something like this.

And this would be the potential energy of those two ions.

They can't get super duper close.

But when they're far away, they're kind of free, and then as they get a little closer, they start feeling that attraction, that Coulomb attraction.

And this energy here, this energy has a name.

It's called the lattice energy, because for the solid, right?

So if I make a solid out of these types of ions, and now I break it all up, and I go all the way out to just one ion, so I go from the isolated ions to a solid of these ionic bonds, then I gain this energy.

And it takes that energy to break them back up.

That's called the lattice energy.

And you can see right away, there are some really kind of interesting observations you can make related to this simple relationship of a Coulomb attraction for an ionic bond.

All right, so here's two ionic solids.

There's sodium chloride.

There's magnesium oxide.

There's Na plus and F minus.

OK, and notice the Coulomb energy here.

And notice the differences in the lattice energy.

This is like four times that.

That is exactly what you get from this, because in this case Q_1 and Q_2 are both 1, right, 1 and negative 1.

And here, Q_1 and Q_2 are both 2, and negative 2.

So it should be four times, and that's, in fact, you can measure how much energy does it take to break this thing up, four times as much.

That's the lattice energy.

So we know something about this.

And you can see trends.

Right, so these are now the radii.

We just talked about these.

Right, radii for ions.

And this is the lattice energy.

And you can see that if I'm the smallest one, and the smallest one there, I have the highest lattice energy.

And that's the $1/r$ part.

So you got the charge and how close they can get.

And you now know from-- just from the trends we've just talked about, the radii trends we just talked about, you know how you might compare how much one ionic solid is bonded, compared to another, because you can talk about how much charge it might lose or gain.

It likes to get close to the nearest noble gas.

And you can also think about their size or relative sizes.

So like the largest one with the largest has a larger r , and so weaker lattice energy.

And boy, is this important.

It changes all sorts of things about the properties of these solids, and that's why, we developed the goodie bag to help you explore that.

So you have different solids.

They're not all ionic, because you're talking about another kind of solid next week.

Right, but you've got some ionic solids.

And look, like, you know, so a lot of them dissolve, but if the lattice energy is so high, maybe it doesn't dissolve.

Or maybe it's harder to dissolve.

How do you know?

Well, you could see it if you stir it.

All right.

But you also may be able to measure.

Are there ions in solution?

If there are ions in solution, than it could be a dissolved ionic solid that you've made.

And so-- oh hardness, Mohs' scale.

You can look that up.

Did something scratch something else?

Right.

That's what that's related to.

Melting point, right, the properties depend on this ionic bond.

OK so there's the goodie bag, and then I'll do my why this matters in the last couple of minutes of class.

All right, so there's the goodie bag.

Oh, you've got a conductivity meter.

You've got a scale.

You've got a scale.

I used to give that scale out with white powder right before Thanksgiving for a different goodie bag, and it was a really bad idea.

But now it was-- it was citric acid, lime juice.

But people traveled like that.

And so we've learned.

We've learned.

And so we give it to you now, but please keep it.

It's the world's most accurate scale.

Why does this matter?

Because look-- you're going to need the scale again.

So please keep it.

Here you go.

Here's another chart of ionic bonds.

These are lattice energies.

OK, these are lattice energies, oh kilojoules per mole, it's just per mole of stuff.

Right, don't get confused by that.

Joules, electron volts, energies, per mole is per mole of atoms, right?

OK, so we can isolate it down to one bond or a mole of bonds.

All right, 15,000 for aluminum.

Look at that.

That's why aluminum is in your toothpaste.

Aluminum oxide is in your toothpaste.

It's in your pans.

It's in sandpaper.

It is a very hard material.

It is a very hard material.

Why?

Because of this.

Literally.

Those Q's are really high, right?

Why does this matter?

Well, I'm going to give you an example in hemodialysis.

Hemodialysis is something that people have to go into a hospital.

People in their kidneys, don't clean their blood efficiently or enough have to have it cleaned by a machine, and 650,000 patients suffer from this and go to a hospital for literally four hours, three

times a week.

It devastates their weekly schedule.

And if you could make this portable, you would change lives in a really big way.

One of the big-- there are people out there trying to do this.

This is an example of a design for a portable hemodialysis machine.

One of the single biggest drawbacks in making it portable is making a filter, because you're filtering blood.

You're filtering toxins out of the blood.

It's something that the body can't do.

So you need to do it for them.

But the filters get mucked up and gunked up, and they're not strong enough, and they can't be cleaned.

And look at this one.

Here is a filter made out of aluminum.

All right, and these things are what's needed.

We need super strong, super resilient new filters that can filter things like blood and toxins out of the blood.

So very, very tiny sizes that we can make over large areas and uniformly.

This is the kind of starting material that we need to make new filters, and it all comes back down.

Why would that be a good one if we can overcome it's brittleness?

Because of the Q's.

Because of the atom sizes, because of everything that we talked about today.

So have a great long weekend, and see you guys next week.