

The following content is provided under a Creative Commons license. Your support will help MIT OpenCourseWare continue to offer high-quality educational resources for free. To make a donation or to view additional materials from hundreds of MIT courses, visit MIT OpenCourseWare at ocw.mit.edu.

**CATHERINE
DRENNAN:**

So that is electron affinity. But honestly, chemists don't really talk so much about electron affinity. They prefer to talk about electronegativity. And these are highly related terms. So this was also re-copied, although completely identical, I think, between the handouts. I just thought it was weird to have re-copied this and not this. So your handout for today is perfect on this point.

So electron negativity, the net ability of an atom to attract an electron from another atom. So you can see that electron affinity and electronegativity are very similar. In fact, all of these terms are highly related to each other. And this idea of electronegativity, of this as a term for a way of thinking about atoms initiated with Linus Pauling.

But here I have up a different picture. I have Robert Millikan. And the reason why I picked this picture is because he helped, a few years after Linus Pauling came up with this idea, coming up with an equation that help people think better about what electronegativity is.

And another reason why I picked his picture to put up here instead of Linus Pauling is that he was an MIT undergraduate, and he was a chemistry major. I'm not sure in 1917 or whenever-- that was the yearbook picture-- when he took-- I don't think it was 5.111 at that point. But at some point, he was here studying chemistry just like you.

And then when he got a faculty position later on, he did some beautiful work that had to do with bonding that he got the Nobel Prize. So some of you may have a Nobel Prize in chemistry one day. And so you want to make sure your yearbook picture is at least as good as this one for other generations of professors to show your picture and describe the work that you did to contribute to the field of chemistry. Oh, and he was born in Massachusetts too, so he is a native to this area in more than one way.

All right. So the way that he-- and this is a little bit more of a squishy definition. So electron negativity is proportional to a $1/2$ -- and IE stands for what?

AUDIENCE:

[INAUDIBLE].

CATHERINE

DRENNAN:

ionization energy. And EA-- our electron affinity that we just discussed. All right. So it's related to these other terms that we have already talked about. So let's think about then what this means. So we can consider an atom with high electronegativity and an atom with low electronegativity. And we want to think about whether an atom with high electronegativity is going to be an electron acceptor or an electron donor.

And that you can tell me, and that is going to be a clicker question. So you can try to grab your handout while clicking at the same time. All right. So we'll take 10 more seconds. Yes, 88%. That's great. Of course, if you looked and if you didn't believe it could be a donor, then that ruled out three of the four, but that's OK. Those are good things.

So yes, if it has high electronegativity, it's going to be an electron acceptor. And part of the reason for that is that it has a high affinity for electrons. And another part of the reason for that is that if you look at the equation, when you have a high ionization energy, something that has a high ionization energy is not going to be a good donor. So that wouldn't make sense. So both of those terms having high in both categories is consistent then with this trend.

So let's take a little bit more of a look at that and why this is true. So high electronegativity, an atom with high electronegativity is an electron acceptor, and then low would be a donor. And so if we think about this-- and this is our periodic table. And again, it's not going to be including our noble gases, which really don't want to be accepting or donating anything.

So in this corner then we had our high ionization energy, and we had a high electron affinity. And we saw last class we had high ionization energy. So it doesn't want to give up an electron, but it does want to accept one. So we have things that are going to be good acceptors. And down here, we have low ionization energy, so it's easy to donate an electron. Oh, let me just put these up, sorry.

So we have then if you're high and high up here, you have something that's a good acceptor, and it's going to have a high electronegativity. So high high means high over there. And then down here, we have low low, which means we have low electronegativity.

Low ionization energy-- it's easy to give something up. Low electron affinity-- it doesn't want electrons. It's happy to give up electrons. And if it gives up electrons, then you can get a complete octet. It can have a noble gas configuration. So on this side, you need electrons. This side, it's happy to give them up.

So if we look then just at a periodic table again, this makes sense. We gain an electron over here. We get our happy noble gas configuration. We lose electrons over here. We do the same thing. So that's a way to think about electronegativity.

All right. So why should we care about electronegativity? And that's because a lot of atoms that are electronegative are used in pharmaceutical molecules, and that this gives them special properties. So to hear in their own words, we're going to hear from a former uropper, Kateryna, talking about why you should care about electronegativity.

[VIDEO PLAYBACK]

- My name is Kateryna Kozyrytska. I come from Ukraine. And I'm interested in how microorganisms fight each other. Humans are very smart. They have found chemical ways to make new drugs. And so we spend a lot of time and money on making a new antibiotic, and then we put it into people and we hope for the best. Bacteria are also very, very smart. And they somehow learn to resist this new antibiotic that we just made.

Bugs, on the other hand, have been fighting each other with the same molecules for thousands and thousands of years, and we see no resistance developing there. So we want to learn what it is about the antibiotics that bugs make to fight each other that makes them so difficult to develop resistance to.

Normally, living organisms use 20 amino acids. But these bugs get very tricky in building their anti-other bugs molecules. So to add the functionality, bugs can do chemistry on the building blocks, the amino acids themselves. And so one of the things that they can do is chlorinate carbons, which activates these carbons for future chemistry.

The protein I study, the halogenase SyrB2, takes chloride ion from the environment and a molecule of threonine, the amino acid, and puts those together forming a chlorine-carbon bond. Since chlorine is so electronegative, when bonded to carbon it pulls electrons away from the carbon atom. And so it makes the carbon to which it is attached much more reactive toward other molecules. And this increased reactivity, at least partially, accounts for the antibiotic effect of the molecule.

I am hoping to figure out how SyrB2 positions all the atoms in such a way that they react in this very controlled, very appropriate manner, so that later we could maybe re-engineer this or other proteins to make them do chemistry that we want them to do. My hope is that we could

understand SyrB2 well enough that we'll be able to remake it into a protein that will actively participate in synthesis of new antibiotics.

[END PLAYBACK]

**CATHERINE
DRENNAN:**

Yeah. So that is Kateryna. And so how adding halogens, because they are so electronegative, is actually a very important area, and I'll give you a couple more examples. I also haven't watched that video in a little bit, and she said she was from the Ukraine. She's technically now part of Russia, I think, from part of Russia, so.

[LAUGHTER]

Anyway, that's a topic for a different day. So this is big business, actually, putting halogens on things. And if you become interested-- if you're taking an antibiotic or something, start looking at what the molecule is, start counting how many halogens are on that molecule, you will find a lot.

So a lot of antibiotics have halogens, either chlorides or fluorine, as shown here. Also a very common antidepressant has it, another example of something that is an anti-diabetic. Huge numbers of molecules have halogens added to them. Some of these are derived from natural products. So nature was making these molecules to kill other bugs, as you heard about in the video.

Other times, they came up with this molecule and they said, well, we need to make it a little bit different. It's being consumed too fast. It's broken down too fast in the body. Let's add some halogens. So sometimes it's sort of a man-made tailoring, but often it's a tailoring that nature came up with.

So why all these halogens? What's the benefit of having a carbon-fluorine bond instead of carbon-hydrogen? And one reason is that having a fluorine, this really electronegative atom, on, say, an aryl ring like this one, it actually sucks the electrons out of the ring and makes it what we call, or organic chemists like to call, electron poor. So it just kind of hauls those electrons away.

And when you make something electron poor, so by replacing C-H with C-F, that can make a potential drug molecule electron poor. And what this does often is make it harder to oxidize the molecule. So we're going to talk about oxidation-reduction much later in the semester, and

we'll come back to this idea.

But this turns out to be really important, because the way that the body metabolizes or breaks down these molecules is that it can oxidize it. And there are a number of enzymes in your liver which will oxidize and break apart these drugs. And so if you make it harder to oxidize, that makes the drug more stable in your body.

So if you want to take a drug, you want it to last for a while. And especially if it's something, an antibiotic, you want it to last till it kills all of the bacteria, not just half of them. And most medicines, you need them to be around for them to have their effect. So you want to tailor those molecules so that they don't get broken down as easily in the body.

And so this is one reason, and this is big business. And a lot of the times, adding those halogens actually involves pretty toxic chemicals. So some people, like Kateryna, are interested in using enzymes to do it instead. Some people who are still using organic synthesis-- example is Steve Buchwald's laboratory.

In fact, if you go to almost any top chemistry department, I think there's someone who's trying to find out new methods of putting halogens on molecules. It's a very important area in designing new molecules and improving them. OK. So that is electronegativity. So one atom added to a big number of atoms can change the property of the molecule by sucking electrons away.

And now we're going to talk about atomic and ionic radius and also isoelectric atoms. So these trends are pretty good. We're back to some pretty good-- fewer glitches. So what is the atomic radius? So here we have $2r$, 2 times the atomic radius. So the atomic radius is defined by the value of r that has about 90% of the electron density. I mean, technically, an electron could be infinitely or close to infinitely far away, but pretty much most of them are going to be about within 90%. We call that the radius of the atom.

So they're trends. Trends are pretty good. So across the periodic table, what matters is the Z effective. See, everything we've learned about it comes back. If you didn't learn it on exam 1, still learn it because we're going to use it again. Across the periodic table, Z effective matters. Down the periodic table, n , or the principle quantum number, matters. So how do these matter? So across the periodic table, Z effective is going to do what? Increase or decrease?

AUDIENCE: Increase.

CATHERINE

DRENNAN:

So it's going to increase across the periodic table, and this results in a decrease in the atomic radius. So again, it increases going across the periodic table because you're adding both electrons and protons. But the electrons are not giving you complete shielding, so you're not canceling out every proton with every electron.

So overall, you get an increase in the Z effective. And because of that, when you have this increased Z , it's kind of pulling the electrons in. And then it isn't until you go down the periodic table when n increases that you start to see the radius increase. So there, the electrons are getting farther away, and so you are getting this bigger thing.

So I like to think about it as sort of the mom at the park. There are some moms, their kids are sort of running everywhere, but other ones are sort of hauling their kids in. They have this positive force that seems to keep them all sort of in the general area. And so they're shrinking the size of their kids' play area with this force that they're exerting.

But if the kids get too far away, it's like, yeah, they're not going to hear you call. They're not going to hear you jump up and down. They're not going to see you. And so they're just going to be out there, and the radius of your kids is going to be farther away.

So we can look at these trends. If we're over here in the beginning, we're starting on this side. We're going across the periodic table. We go down, then we have a jump up when we increase n . And then we go down again, then we have a jump up when we increase n . We go down again, except over here there's a little glitch.

Those d electrons, they're back. They're going to give you a couple glitches. I love d electrons. We're going to talk about them more around Thanksgiving. That's my favorite part of the course. Those d electrons are always causing trouble. OK. Then we go up again, and then we go down, and go up, and go down. Those are pretty good. Those are pretty good trends.

All right. OK. So ions. Ions are different than their neutral parent, once again. So we saw this before, that when you start filling the $3d$, the energy levels change. So ions can have different properties than their neutral parents. And so if we have two kinds of ions, we can have cations, which are positively charged. And so a positively charged ion will have lost an electron, and so it's going to be smaller than its parent.

And so we can see here lithium. And then in the center, that's lithium plus. So when you lose the electron, the radius actually shrinks quite a bit. It's like that electron was just really kind of

causing a bigger radius. And when it's finally gone, you're at a smaller size over there.

Anions-- negatively charged ions. So they're gaining an electron, and their radius is larger than their parent. And so you can see over here, we have oxygen in the center. Oxygen minus 2 is much larger. And again, we can see some of the other trends. Some of them are the same.

The ionic radius also will increase when we're going down a group, so when n is increasing, so from lithium to sodium. We have an increase from fluorine to the top of the periodic table to chlorine. So we still, as we increase n , increase in size. But you have to think about the ion-- did it lose an electron, or did it gain an electron-- to think about how its size changed with respect to its parent.

So why does this matter? There's one example of-- if you're interested in biology or anything to do with the brain or neurons, then you should care about ion channels. So there are channels in membranes that bring ions in, and this is really important. So ion channels are in muscle cells and in neurons. So if you want to move or think, something that MIT students generally like to do both of those things, you need ion channels to do that. So ion channels should be important to you.

And you want ion channels. They regulate the influx of ions into the cell and allow for really rapid responses, which is also really important. And amazingly, they're highly selective for certain ions. So this is important. It needs to be tightly regulated to say how much sodium you have in there or how much potassium that you have coming in. And if the ion channel took potassium when it was supposed to take sodium, that would not be good.

So these channels are designed by nature to be highly selective, and so they care to be highly selective. And you're talking about a plus 1 maybe or another plus 1, they have to think about the radius. So why don't you tell me what the differential is in the radius from smallest to largest for these three different ones. All right, 10 more seconds. OK, great. So most people got that right.

Let's just kind of take a look at that. So here we want to think about the neutral, and then this one has one less electron. So that's going to be smaller. And then when you compare potassium with sodium, you have to think about n . And so this is down farther, so that's going to be bigger than sodium. So here we're thinking about the difference in electron configuration, and here we're thinking about n .

All right. So amazingly, these channels have it figured out, and so look at this. This is about significant figures. Too many people lost points on significant figures, I have to say, on the exam, so make sure you learn them.

But if you had, say-- you say, oh, what's the difference between 1 to 3 significant figures? The difference is sort of the potassium versus the sodium radius. So 1.38×10^{-10} , 1.02×10^{-10} -- those seem like pretty small numbers. Does it really matter if it's 0.2 versus 0.38? And the answer is, yes, you'd be dead if nature could not distinguish between these significant figures for you.

So these channels are designed to be selective at that kind of atomic scale, and only let one ion in. And so Rod MacKinnon, who's a crystallographer, won a Nobel Prize for solving some of the structures of these ion channels. And this just shows a ribbon drawing, and this shows an all atom drawing of a channel, and there's an ion going through. That's its hole, and its radius is perfect for that ion.

And the other one, even though it's not that many significant figures different, doesn't fit. And that's pretty amazing. Nature is truly amazing. That's the hole. It makes a perfect hole just for the one kind of ion that it's supposed to accept.

All right. So now, there are one more definition that we're going to do. There are things that can have the same electron configuration. Those are called isoelectronic, and let's think about those. They don't necessarily have the same size, but they have the same electron configuration. And I'm just going to write these out.

So when we think around other ones near neon, noble gas, that would have that exact configuration-- so how do we get fluorine to have that configuration? What does it need to do-- gain or lose an electron, and how many? What would its state be? So what do I write-- what's the thing for fluorine that is going to be the same electron configuration just in terms of its charge? I'd write F what?

AUDIENCE: Minus.

CATHERINE DRENNAN: Minus. For O, what am I going to write? For oxygen, what am I going to write?

AUDIENCE: 2 minus.

CATHERINE 2 minus. And, say, nitrogen? We'll stop there. What's that?
DRENNAN:

AUDIENCE: 3 minus.

CATHERINE 3 minus. Great. Let's go on the other side. What about for sodium? What does sodium have to
DRENNAN: do to have that configuration?

AUDIENCE: Plus.

CATHERINE Plus. What about Mg?
DRENNAN:

AUDIENCE: 2 plus.

CATHERINE 2 plus. Aluminum?
DRENNAN:

AUDIENCE: 3 plus.

CATHERINE OK. And silicon-- 4 plus. So you get the idea. And now, we can think about which will have
DRENNAN: bigger and which will have smaller radii as well. So are these going to be bigger or smaller?

AUDIENCE: [INAUDIBLE]

CATHERINE Right. So they're going to have larger radii than their parents because they've all gained. And
DRENNAN: then over here, these will be smaller since they've lost electrons compared to their parent ion.

OK, so let's just do one, which should be very fast, clicker question. It is a clicker competition after all, so we've got to get in some extra clicker questions. And this should be very fast, I think.

All right, let's just do 10 more seconds. You have a periodic table up here in case you need it. Yeah, OK, that's not going to distinguish the recitations very much. Yeah, so you just have to look at what is nearby and think about how many electrons it needs to gain or lose to have the same configuration.

All right, bonds-- now, we're up to bonds. There are three types of bonds we will discuss today. We probably won't get to them all today. After all, the handout stuff. But anyway, then we will discuss in the class over time.

Now, some of you have probably figured out-- almost everyone has probably figured out that one of the things I love to do is teach chemistry. I love to teach chemistry. Some of you have come to my office hours or pizza forums or even paid attention to some of my slides about office hours may realize that I love dogs.

So I love teaching chemistry, and I love dogs. What is the most amazing thing that you can think of? Dogs teaching chemistry.

[VIDEO PLAYBACK]

CATHERINE So now, I'm going to let dogs--

DRENNAN:

- Welcome to "Dogs Teaching Chemistry!"

CATHERINE --tell you about bonding.

DRENNAN:

- --is chemical bonding. Chemical bonds are what holds atoms together. A chemical bond is an attraction between atoms that allows the formation of a chemical substance. The electrons that participate in a chemical bond are called valence electrons. These are electrons that are found in an atom's outermost shell.

Let's take a look at the types of chemical bonds that can be formed between atoms. An ionic bond is formed when one of the atoms will lose its electron to the other atom. This results in a positively charged ion called a cation and a negatively charged ion called an anion. Positive and negative attract, and the result is an ionic bond.

Covalent chemical bonds involve the sharing of a pair of valence electrons by two atoms. There is also what is called polar covalent bonds. These are covalent bonds in which the sharing of the electron pair is unequal. The result is a bond where the electron pair is displaced toward the more electronegative atom. Thanks for watching, and we'll see you guys next time.

[END PLAYBACK]

[APPLAUSE]

CATHERINE
DRENNAN:

There's, I think, one other one, but that is totally the best one. And everything they say is exactly right, so it's really exciting. I even love it when they had the two-- they put two balls in there that they were sharing. It's just really very well done.

OK, so now, you've probably filled in some of your notes here, but in case you missed some of them, I will tell you exactly what the dogs just told you. The dogs had it completely correct. So ionic bonds is the transfer of an electron, as you saw, and then the generation of a cation and an anion that are attracted to each other due to the charge.

So the bonding comes from that attraction between the positively charged and the negatively charged atom, and an example that you're probably all familiar with is table salt, NaCl. And so you have Na plus and Cl minus that are attracted to each other and form these bonds, which creates table salt. So let's see how far we can get in thinking about, really, this interaction between ionic bonds, and we'll see if we can get through ionic bonds. We might have to wait until covalent bonds until Monday, but let's see if we can finish this.

So the formation of NaCl from neutral Na and neutral Cl will first involve forming your cations and your anions. So you have Na going to Na plus plus an electron, and so here, you're talking about a process where the energy is going to be equal to the ionization energy again. So we're not moving far away from these terms because you're talking about ionizing a neutral atom to Na plus, and there's a value for that.

And then we're talking about neutral Cl, neutral chlorine, going to Cl minus, and so it's gaining an electron. So here, the process you're talking about is the electron affinity, and so the energy change here is equal to the negative electron affinity, which is minus 349 in this case. So this is a favorable process here to gain this extra electron, and so overall, the ΔE here is negative.

So now, if we're going to talk about this process here, we have two of these, so we're going to go-- we're going to put these guys together. So we need both of those to ionize, and so we can add up what energy difference we should expect to form Na plus and Cl minus together from their parents. And so we have, now, a plus-- we've added these two together-- a plus 145 kilojoules per mol.

So this seems weird. It's plus, and so now, we're seeing that the formation of these ions from their neutral atoms has this positive value, which means it requires energy. But we think about

NaCl as being this natural table salt thing, so why is there so much table salt if this requires energy to do it? And the answer is that this is only part of the process.

You need to form your cations and anions, but then you have energy of them coming together. So they're attracted to each other, and that's a really important part of forming the bond. And they're attracted by a simple coulombic relationship here. So the attraction between the positively charged and the negatively charged ion has an energy of minus 589 kilojoules per mol, so overall then, if you consider both forming Na plus and Cl minus and the attraction between them, we have a negative ΔE 444.

So the net energy here is in favor of forming NaCl. We have a decrease in energy. This is a stable compound, so let's look at where this number comes from. So we just put this out. That's the coulomb thing, but let's actually calculate this and see where that number comes from.

So we're back to coulombic equations again. We never get very far away. They turn out to be very important in chemistry.

So we have the coulombic potential. We have z 's, our charge on our ions. We have the absolute value of the charge of an electron squared over 4π , our permittivity constant in r , our distance between those ions.

And for any Cl the bond length, or the distance between Na plus and Cl minus, is 2.36 angstroms, so we can use that and just plug it into the equation. And we have plus 1 for the sodium, minus 1 for chloride, so overall, this will be a negative term. And if you work out the math, it's minus 9.774 times 10 to the minus 19th joules, and we have three significant figures. What's limiting our significant figures?

AUDIENCE: [INAUDIBLE]

CATHERINE DRENNAN: What-- yeah, the bottom length is limiting it here. I'm just going to talk about significant figures for a while, until Test 2 when you can demonstrate I can stop talking about significant figures. All right, so then we want to convert to kilojoules per mol because that was the number I gave you, was in kilojoules per mol, so we have our conversion factor between joules and kilojoules. And I guess I should mention up here-- coulombs cancel, and our meters cancel, so we're left in joules up there. Units are also important.

Then we can use Avogadro's number, because we're given kilojoules per mol, and we can get out the number I gave you before. So this number really just comes right out of this equation minus 589 kilojoules per mol. So we have our ionization energy to tell us about forming ions. We have our electron affinity, and now, we have a coulombic relationship.

So I said before-- this is what I showed you before-- that we have this attraction that's favorable. We have forming the ions, which had a positive energy associated with it, but overall, this process has a lower energy. It forms a bond. But this is just based on this calculation.

So we can ask what is the experimental measurement for this interaction, and we note that it's somewhat different. So we have, instead of minus 444, we have minus 411 kilojoules per mol. So why the difference?

So again, our ionic model, which just considers ionization energy, electron affinity, and that positive coulombic interaction in our experimental result-- so problems with this model that we did up here include that we only thought about favorable interactions. There are also going to be some that are not favorable; protons against other protons, repulsion, electron electron repulsion. So there are some negatives. It's never all positive in any relationship, whether it's sodium chloride or anything else.

Always some negatives, and we ignore those. And the result of this is that you're going to get a larger ΔE predicted than the experimental value. So it seems like this is more favorable, like that's a stronger bond, a stronger interaction; but really, if there's some repulsion, that's overestimated. It's really going to be a lower value, which is what you see in the experiment.

Also, we just said that sodium plus was one point charge and Cl minus was another. It's more complicated. Their interactions are more complicated than that.

And we ignored quantum mechanics, but in doing that, we did pretty well. If we had one significant figure, we would have been perfect, so to one significant figure, these approximations work really well. OK, that's it for today, and I will see you Monday. Have a great weekend.

All right, let's just take 10 more seconds on the clicker question. Great, so people are getting the hang of this. If you hadn't yet, there's still time.

So as you're going across the periodic table, you are increasing z and increasing the z

effective as well, because you don't have total shielding, so aluminum is the correct answer. It has a lower z effective, and so therefore, a smaller ionization energy. The electrons aren't held as tightly.

All right, so we're talking about chemical bonds, and it seems like an appropriate topic to talk about when we're also talking about bonding as a community, so that seems like a good thing. So a chemical bond is an arrangement of atoms so that they come together in such a way that they're lower in energy than they were when they were apart. So they're more stable together than they were apart, and that's a chemical bond. So this is page five of the handout from last time. And excitingly, we have Lecture 10 handouts today, so there's lots of things working today.

All right, so this is lower in energy, i.e. more negative, when these atoms come together. So a chemical bond-- as you saw last time with those wonderful dogs sharing a pull toy, a covalent bond is a bond where the electrons are shared between these two atoms, and each atom is giving up one bond to share. So we can think about this more graphically of what is happening, and we have this little plot on your notes, where you're going to be filling in a bunch of details.

So we have the internuclear distance, r , the distance between the two nuclei. And we're back to hydrogen for the moment, so we're going to talk about a bond between two H atoms. And on the axis over here, we have energy. So we have energy versus the distance between these two hydrogens.

So at 0 energy, we just have the hydrogens by themselves. They're not interacting with each other in any kind of way that lowers either one of their energies. There's no interaction, no energy change. They're not interacting.

Down here at this dash line, we do have a bond, so we formed H_2 . The hydrogen atoms are interacting with each other, and this is lower in energy. So what does this plot look like then if you draw energy versus this distance? So up here, it's above and higher energy, above 0-- this is unfavorable-- going down to this dashed line and then going back up to 0.

So let's think about what's happening here, and there's a bunch of different kinds of interactions you can have between those two hydrogen atoms. There are repulsive interactions, nuclear nuclear repulsion, electron electron repulsion; and there are positive

interactions like the electron nuclear, the positive and negative. So up here, these atoms are very close together, and that is-- they're really too close, and that's unfavorable. Two objects trying to occupy the same space at the same time is unfavorable.

But as you start separating out these atoms, then they become more comfortable, and you get to a distance where you have the sharing of electrons. They're next to each other. They're interacting in a positive way, and they're sharing.

But then if you bring them too far apart, they're no longer communicating with each other. We don't even know where one of them is at this point. They're just kind of infinitely far apart. No interaction. So we go from too close to just right to too far away.

So this distance here that has this minimum energy-- OK, so we have this minimum energy here, and that's the distance at which you have this really favorable interaction, this bond length. That's the bond length between those two hydrogens right here. Now, we can think about this energy difference over here, and it has a special name. So this is the dissociation energy, or $\Delta E_{\text{sub } d}$. Sometimes, it's just called big capital D in your textbook.

So this is the energy that's needed to dissociate those atoms. So if the atoms come together, and they're lower in energy, if you then want to break them apart, you need to put energy in so that they can be broken apart. And that's called the dissociation energy. And if it's a really big number, that means it's very hard to dissociate them, and if it's a small number, it's pretty easy. So this is our dissociation energy here.

So for hydrogen, this value is 424 kilojoules per mol. And if we were putting it on this axis here, it would be right down here, so it's going to be a negative value on this axis. We're below 0, so the negative of the dissociation energy is found down here, minus 424 kilojoules per mol. And if you wanted to break that bond, you would need to put in energy, so dissociation energy is positive. It's the energy you need to put in to break those bonds.

So we can think about this plot now and consider looking at the plot and evaluating what you can and cannot say about different kinds of compounds, their bond lengths, and also their dissociation energies. So we good people? Mostly got this written down? And today, for some of the people who wanted to come late, we can post versions of this, too.

All right, so now let's consider this plot and ask which bond is stronger? So we have hydrazine, and we also have molecular nitrogen, N_2 . And in your dashed line, you have nitrogen, N_2 , and

in the solid line, you have hydrazine.

So can you just look at this plot and tell me which is stronger? Is nitrogen or hydrazine stronger? And you can just yell out what you think.

AUDIENCE: [INAUDIBLE]

CATHERINE DRENNAN: Yeah. So nitrogen is going to be stronger here, and people knew that because it's a deeper well. So you go way farther down, there's more stabilization. It's a lower energy, a lower negative value of energy, so that means it's going to have a greater dissociation energy. You'd have to put in more energy to dissociate nitrogen nitrogen than these two submolecules here in hydrozine.

All right, so we can also look at this plot and ask the question, which has a shorter bond? What do you think? Which bond is shorter?

AUDIENCE: [INAUDIBLE]

CATHERINE DRENNAN: Nitrogen as well, right. So nitrogen is also the shorter bond, and we know this because this is increasing distance. And this is closer to the axis, so this has a shorter distance. And later, we're going to be doing Lewis structures and other things of nitrogen and discover it has a triple bond, which you may already know, and we'll talk more about nitrogen's amazing triple bond as we go along.

And so that's a very short bond and a very, very strong bond as well. So you should be able to look at these plots and evaluate what kind of dissociation energy it would have, is it bigger or smaller, and also what kind of distance you expect between them. And you should be able to draw these kinds of plots on the exam, if asked, in just kind of simple detail. Nothing too fancy.

All right, so in terms of bond strengths, carbon monoxide has one of the strongest bonds, so it has a very large dissociation energy. And iodide, I₂, has one of the weaker ones. And later in the semester, we're going to be doing a demo that shows why that's kind of cool, that weak bond leads to some cool, cool demos. OK, so those are covalent bonds. And, Ashley, could you just close that door, please?

All right, so let's finish polar covalent, and then we'll have our moment of silence, and maybe it'll be silent in the hallway by then as well. All right, so polar covalent bonds-- so last time, the dog showed you that you can have equal sharing and unequal sharing. And of those of you

who've watched dogs play with pull toys, most of the time, the sharing is pretty unequal, and so whenever you see that again, you can think polar covalent and tell your friends, and they'll be like, "I don't know what happened to you at MIT. Those are two dogs playing. What are you talking about?"

So this is unequal sharing of electrons between two atoms, and this happens when those atoms have a very different electronegativity. So in general, a bond between two atoms is considered a polar covalent bond when the difference in electric negativity between the atoms is greater than 0.4 and less than 1.7, and that's Linus Pauling's scale and works quite well. So if we look at this little plot, we could see that carbon hydrogen bond only has a 0.4 difference, so that would not be considered a polar covalent bond. But nitrogen hydrogen is a difference that is greater than 0.4, so that would be a polar bond. And so you can use these values to think about whether you're going to have unequal sharing, and the more electronegative element is going to be pulling on those electrons.

We also can use this to think about polar molecules, and this is kind of a little bit of a flash forward to Friday's lecture. We're going to talk about shapes of molecules because a polar molecule has to have polar bonds, but also has to have those bonds arranged in such a way that there's a net difference, there's a net dipole, there's a net kind of pulling of those electrons in a different way. So here is carbon dioxide, a molecule that causes global warming. Yes, I said it on the videotape. Yes, I believe that human beings are responsible for some of the global warming, and we should do something about it.

This molecule does have polar bonds, so we have carbon in the middle and oxygen here. So carbon and oxygen have an electronegativity difference greater than 0.4, but it's not really a polar molecule. It's a non-polar molecule, and that's because of its shape. So shape matters.

So we have pulling of electrons one way, but we have equal and opposite pulling of electrons other way. This would be really cool if the dogs could have done this as well. So in this case, we have a non-polar molecule that has polar bonds. Now, there's only very few cases where this would actually be true, and you really need to think about the shape. And so that's why we're going to talk about shape.

Another molecule that also has polar bonds is water. So here, we have polar bonds between the oxygen and the hydrogen, so oxygen and hydrogen, greater than 0.4. But in this case, the shape of water is such that they don't cancel out, and you do have a net dipole. You do have a

net charge on that molecule, makes it a polar molecule. So we need to know about the shape, and we need to know about electronegativity.

So in large, organic molecules, sometimes we just talk about the number of polar bonds and then think about whether that's likely to be a polar molecule or not. You can't always think about the shape of something really complicated and what direction all of the pulling of electrons is going, but we can at least count polar bonds. So here are two vitamins, vitamin A and vitamin B-9, which I think is also B-10 and a number of other B's.

Its name is folic acid. They kept finding it again and again, so there's a whole gap of b-vitamins where they're like, oh, B-9, folic acid; B-10, also folic acid; B-11, I think also folic acid. B-12 something different, though.

So anyway, this is a very important B vitamin, and I'm actually going to come back to this molecule later in the course. But we can think about how many polar bonds it has, and that's a clicker question. And there, you have the molecules up here to look at. All right, let's just take 10 more seconds. It takes a while to count, but I think you probably can answer it without maybe fully counting all of them.

All right, so over here now, the answer is the folic acid, and I'll just highlight. You might not have found all of these polar bonds, but you should have at least seen that this one really didn't have many. Vitamin A only has one. So we have polar bonds down here between carbon and oxygen over here, carbon and oxygen, nitrogen, hydrogen, carbon oxygen again, nitrogen with hydrogen, oxygen with hydrogen, nitrogen with hydrogen over here.

So folic acid is quite polar, and if we're going to think, now, about whether it is a water soluble vitamin or a fat soluble vitamin-- which is something that a lot of times your supplements will tell you about. If it's water soluble-- and we'll talk about this more later-- like dissolves likes. So water likes polar molecules, which makes folic acid water soluble and makes vitamin A fat soluble. It's not very polar, doesn't dissolve very well in water.

And this kind of turns out to be important, in that if you read your vitamin supplements, if you take vitamin supplements, it will often tell you interesting things like how many hundred times over the daily recommended allowance this vitamin tablet is, and if things are water soluble, it doesn't matter so much that you're taking way more than your body actually needs. You just have a very expensive pee. But if it's fat soluble, then it's going to stay in your body, and you don't need it, and it can be a little bit toxic. So try to think about the vitamins. Not everything--

even though vitamins are good for you, they're not good in every kind of amount that you could take, so you will now use your knowledge of polar bonds to figure out whether you should be taking certain vitamin supplements.