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PROFESSOR: OK. Settle down. Settle down. The weekend doesn't start until after the lecture. First we learn. All right. So couple of announcements, the big one being the reminder about that little celebration next Wednesday, coverage through the end of the lecture on Monday, there'll be no weekly quiz next Tuesday, and Hilary wanted me to announce that there's a job opening for someone who's in class anyways, and wants to-- if you take decent notes and your handwriting is legible, you want to make \$9 an hour, go see Hilary.

All right. Let's get into the lesson. Got a lot to cover today. Last day, we were looking at the formation of molecular orbitals by linear combination of atomic orbitals into molecular orbitals. So we had our SLI six-letter initialization, and we saw the formation of sigma and pi molecular orbitals.

And the way you distinguish between sigma and pi is based on electron distribution. Specifically, in a sigma molecular orbital, you have continuous electron density between the nuclei. That is to say, no holiday. So I've drawn a sigma orbital here where I've taken two P orbitals lying on their sides, so they're bonding longitudinally and when the bond is formed, the one nitrogen nucleus is here and the other nitrogen nucleus is here and you can see that there's continuous electron density, no holidays between the two nuclei. So this would be called a sigma bond.

The pi molecular orbital, on the other hand, has a nodal plane containing both nuclei. So since the nuclei are in a nodal plane, that means that you've got to have two lobes, one above the plane and one below the plane, and that's shown here. This is the example of nitrogen, where it forms both sigma and pi bonds and to form the pi bonds, you take the lateral smearing. And so this is what I call the 88. If you take the 88, you're going to make a pi bond. So that's lateral smearing of the p atomic orbitals.

And the last thing I want to make sure that you appreciate is that when you form multiple bonds, you have to have sigma and pi. You can only form one sigma bond because you have no holidays between the two nuclei so now you can't go and invade that same space with a second bond. So that means you're going to have to get beyond that zone and that means you have to go into pi. So if we've got only bond, it's a sigma. If you've got two, it's a sigma and a pi. If it's three, it's a sigma and two pis. That's what we saw in the case of nitrogen. I didn't complete it here, but there's two pi bonds because this is a triple bond. So there's a sigma and there's two pi's. One pi is shown here, and that's vertical to the board, and the other pi is the sigma x, which would be orthogonal coming in and out of the plane of the board. So those smear give us two of these and then we've got the sigma bond.

And we even extended this to hybrid systems, and towards the end of the lecture, we looked at ethylene, and I just wanted to come back to that for a second because I raced through it there. And so we wanted to look at C_2H_4 , which is ethylene. And in the case of C_2H_4 , we recognized that when we went through the Lewis structure, we had a molecule that looked like this, and the thing to note here is that we have a double bond. So once you see this double bond, then that means I'm going to need a mix of sigma and pi.

Now earlier we had studied methane-- methane, which was CH_4 , and to get methane, we hybridize sp^3 . sp^3 , which gave us four identical struts coming off the carbon. Well, we recognized that wouldn't work here because we needed to make a sigma and a pi so the gambit here was to reserve a p orbital. Reserve a p orbital. Don't combine all of the three p orbitals. Combine only two of the p's and that gave us sp^2 plus a p. And the box notation for that looked like this. Here's the p orbital that's not been hybridized. These are the sp^2 orbitals that are a mix of S, and I've taken two of the p's and I'm leaving one of the p's unmixed. And now I've got four electrons to put in here and that's where I ran out of gas last day. I have to put the fourth electron in here. These are still degenerate.

So now I've got still the capability of forming four bonds. You can see-- whenever you see carbon, there's an interesting-- just something to train your eye, if you see carbon in a structure, you need four sticks coming out of the carbon, or else there's a spelling error. So look here. This is carbon with four hydrogens. One, two, three, four. This is carbon one, two, three, four. Here's a carbon one, two, three, four. It's always got four struts. That's not what's different here. It's the arrangement of the struts. These are all four going at 109 degrees in three space. This one is going to have three of them identical, because we studied last day the principle that equal energies, equal bond energies, imply-- I want to get the spelling here. Equal bond energies imply equal spatial disposition. Equal spatial disposition of those bonds.

So if we have four bonds of equal energy, we have the tetrahedron. If we have three bonds of equal energy, how do I take three sticks coming out of a central carbon and make them uniformly distributed in space? They're going to lie in a plane and they're going to be 120 degrees apart. That's the consequence of having this structure. All of this follow and why did we get to this? Octet. So here we are. Here's the three of them and here the fourth bond is p, and it's normal to the plane of the board, and now we can take one of these and bond it to another one, bring them in along here. This is an sp^2 and an sp^2 and they bond to make a sigma bond. How do I know? Look, no holidays. Continuous electron density. And now in the plane normal to the board, I have lateral smearing, and I make a pi. So I've got a sigma and a pi. That gives me the double bond and let me just show you the cartoons here because we've got some nice artistic renderings.

So here we are. This is the top view of the three sp^2 's and then the p's are vertical. so we're going to move along

the z axis, bring two of these together-- there's the smearing of the sp²'s on axis. So this is a sigma bond, No electron density holidays from one nucleus to the other, and now we're going to smear the lobes of the p orbitals of the unmixed carbon. That gives us the double bond and there we are. This is the same thing taken from your book. This is the sigma orbital and then chemists like to do this. You see this sort of slashed line? It sort of indicates it's receding. It's going into the plane and then these are coming out. This hydrogen is coming out at us and that other one's going way, way in the back and this is in the plane and there's the lobes colored is they irritate me always.

By the way, they've chosen to make this the z-axis, which I don't like. The z-axis should be the axis along which they bond. And then here's acetylene. Acetylene is a compound that C₂H₂. C₂H₂, and I'm not going to go through this in class. You can try it when you have nothing else to do. and so it's got a triple bond. it's got a triple bond between the carbons and then hydrogens on either side. Again, look at the carbon: one, two, three, four. But now we're going to have to have two lateral smears, so that means I'm only going to take one of the p orbitals and leave two in reserve. so when I do that, I'm going to end up with this, starting with the box notation.

Here's carbon as it would exist without any hybridization. An S and three p orbitals. And what I want to do-- I want to have two pi's and a sigma, so I'm going to just capture one of these and reserve two p's as unmixed, so that will give me this. This is one S and one p, so this is called sp hybridization. And then I've got the two p's unmixed, so now I've got one, two, three, four. are both orthogonal to one another and these two are what? I've got two struts coming off the carbon.

How do I get two struts coming off a carbon symmetrically disposed in space? That's it. So look at the shape of the molecule. You see it here. There's the-- the sigma orbital is here and then we have vertical and then out on the plane. And so you first would say, wow, I understand that. That looks a lot like nitrogen. And then, of course, I looked at that, and I said, what about the right-hand rule? And according to the right-hand rule, the z is going in this direction. It's a little more subtle, but I pick up on stuff like that. I can't help it.

OK. So this is good. So what I've done here is gone through a number of instances where we start with electronic structure. sometimes we hybridize, sometimes we don't. But what follows here? When we start with the octet rule and we start playing with the energetics and forming the orbitals, can you see that there's a connection between electronic structure, which then dictates the bond disposition in space, which ultimately dictates the architecture of the molecule? Molecular architecture. So there's a way of doing this There's a way of doing this systematically.

I just got ahead of myself. Before we get to that, I wanted to just pause here for a second and reflect upon the properties of the covalent bond as contrasted to the properties of the ionic bond. So let me get to that first. Properties of covalent bond. OK. The first property I want to highlight is the fact that the bond is saturated.

Remember when we said that the ionic bond was saturated? So what does it mean? Saturated? It means that there's two atoms only. Once you've got the electron pair between the two atoms, the bond is formed and that's the end of the story, whereas with ionic bonding, you can keep by electrostatic attraction piling on and piling on. So this is the concept of saturation. So we have the shared electron pair. Shared electron pair and that's all there is to the story.

The other thing about the covalent bond is that it is directional. And so the bond spatial arrangement-- these are the positioners. You can't put atoms where there are no bonds and since the bonds have a direct spatial arrangement, then that's going to inform molecular architecture. Dictates molecular architecture.

So what I'd like to do next-- so for example, here's one that we know. You've seen this one. This is methane. This is sp^3 hybridization. So we learn that methane is tetrahedral. We say that the four corners of the hydrogen positions form a tetrahedron. So we call this molecule tetrahedral. So I could look at other molecules that have sp^3 hybridization. So if I look at something like titanium tetrachloride, which we met earlier to make titanium, this is titanium in the center and the four chlorines. It's the same molecular shape.

If you see leaded gasoline, the compound in unleaded gasoline is tetraethyl lead. Lead at the center, ethyl's at the four corners of the tetrahedron. So all of these-- and same over here. This one here, sp^2 , this means the molecule is planar. The carbons and the hydrogens lie in a plane. The whole molecule lies in the plane of the board. If I give you any molecule that's got sp^3 hybridization and has one, two, three, four, five, six atoms, in other words, all the bonds have atoms, the molecular is going to be a planar molecule with this, the 120 degrees.

So we've got a grand scheme here that allows us to codify this behavior and the connection between molecular structure and the electronics arrangement and it's written up here under the term Valence-Shell Electron-Pair-Repulsion Model. Let's write that down. I'm going to go through some examples.

Valence-Shell-- hyphenated-- Electron-Pair-Repulsion. That's the model, and it goes by the acronym VSEPR, which is a five-letter acronym. This is not an initialization because chemists reverse these, and they call it VSEPR, because the V and the S is kind of hard to pronounce unless you have Slavic blood. So it's called VSEPR, and that is the scheme that's up here.

So I'm going to go through some examples and show you how you can start with electronic structure and figure out what the molecular architecture is. So I'm going to follow this. So first thing we're going to do is write the Lewis structure. And the example I want to start with is sulfur hexafluoride. I want to-- so the question is, what is the molecular architecture of sulfur hexafluoride? Can we predict the shape? And this is an interesting molecule. It's very dense. It's about five times-- it's a gas at room temperature. It's about five times the density of air. It's got an atomic mass of about 150, and it's used in the light metal industry. It's used as a blanket gas for casting of

aluminum and magnesium to keep the air away from the liquid metal. Otherwise, the metal would oxidize.

But unfortunately, the sulfur fluorine bond is a really good infrared absorber, and when this stuff leaks out, it's a fantastic greenhouse gas trapper. It's got a greenhouse gas coefficient that's, I don't know, something like 10,000 times that of CO₂. So there's a lot of conversation about banning this. If we ban this, we're going to have to find a substitute. There is no readily available substitute. And what that'll mean is the cost of aluminum castings and magnesium castings will go up, which means that they won't be used in automobiles. Instead, cheaper steel will be used in automobiles, which means the mass of the automobile will go up, fuel economy will go down, tail pipe emissions will go up and greenhouse gas emissions will go up.

So how do we get ourselves out of this situation? Do we ban SF₆ or not? How do you go about thinking about that problem? Well, if you're in Washington today, it's the lobbyists with the biggest bag of money that comes in. That's what's going to dictate policy. But if we lived in an ideal world, we'd have people like you who understand the chemistry and what would be the analysis? You'd look at all the SF₆ emitted from the cast shops in a given year and compare that to the increase in tail pipe CO₂ emissions if you replaced the aluminum with steel, and then you figure out what the trade off is. And maybe what you'd do is set up a policy where you would gradually phase this out with-- heaven forbid-- solid research support to give us what we need. That way we get clean air and everything else. OK. So that's what we would like.

Now let's look at the Lewis structure of this thing because that's the first thing we're supposed to do according to VSEPR. So I'm going to put sulfur here and I know sulfur's got-- it's 3s² 3p⁴. So it's got six valence electrons. I'm going to put them symmetrically around the sulfur, and then I'm going to bring in the fluorine. I'll bring in one fluorine, and I know it's got seven valence electrons: one, two, three, four, five, six, seven. And the same thing with the other fluorines. So I'm going to put one, and then for the electron pair, I'm just going to draw a line. OK. You're smart. You can figure that out. That's just shorthand. OK. Here's another fluorine. One, two, three. So there's two, four, six, seven. There's one, two, three, four, five, six, seven. One, two, three, four, five, six, seven. One, two, three, four, five, six, seven. Beautiful.

OK. So now what I want to do is identify all of the bonding orbitals. So I see around the sulfur, I've got one, two, three, four, five, six, six bonding orbitals, six bonding domains, if you like. Well, that's 12 electrons. I said octet stability. Looks like sulfur took that idea and ran with it. And it turns out that this does happen on occasion in elements that are very electronegative. Very, very strong non-metals will form what is known as an expanded octet. And here we've got-- here's a list of elements.

[MICROPHONE ADJUSTMENT]

PROFESSOR: We need a high-tech tie clip. Pull out a little more, maybe it's constrained. It's tethered. the bond is

too tight. That's what it is. It's not flexible. It's covalent, you see.

All right. So here we are with the elements, and the ones shown in the in the grayish-blue here are the ones that on occasion might form an expanded octet. Not always, but if you do see 10 or 12 electrons around one of those elements, don't be alarmed. If you see 10 or 12 electrons around the potassium, that's probably a mistake.

OK. So we have six bonding domains and what else do we have? We have six bonding domains, and we have zero non-bonding domains, and in total, we have six electron domains, in other words, six electron pair systems around the central sulfur. And on the basis of the total of bonding and non-bonding, this is what dictates the skeletal structure. In this case, the skeletal structure is trivial. It's the same as the atomic structure because there are no non-bonding domains. And so I've got six of them. They're equivalent. So what's that going to mean in terms of shape?

Well, I'll put the sulfur in the center and I've got one, two, three, four, five, six. So two in one plane crossways and then two vertical. And then I put the fluorines-- one, two, three, four, five, six. So there's the molecular architecture and it's termed octahedral. I know at first this throws you because you see six struts. Where do you get the octahedral? Well, if we connect the four fluorines along with the sulfur, they lie in a plane, and if we were to pitch a tent here, we could pitch a tent to the upper fluorine, and we'd end up with four faces. So this is Greek. Eight faces. So there's four above and four below. I know it's sort of like a Jasper Johns painting. You say, octahedral. You hear octa, you're thinking eight and you're seeing six fluorines or seven. Anything but eight. So this is anything but eight, but it's eight faces.

Last thing-- let's look at the energetics. To hybridize like this, we needed to do something different because we know that the base structure of sulfur would be $3s^2 3p^4$. So this is $3s$ and this is $3p$. So that's $3s^2$, and then $3p^4$ would be one, two, three-- whoops! I violated the polyexclusion there. OK. There we go.

So I can't form six bonds with this. I only have two bonds. So you say, well, okay, let's make sp^3 . Well, if I make sp^3 , sp^3 will give me this. So now I got one, two, three, four, and then I'm going to put five, six. I'm still down with two bonds. I need to have six unpaired electrons. So I've already consumed my s 's and my p 's. Where else can I go shopping for orbitals? I go to the orbital store and the orbital store is over here. Here's $3d$. One, two, three, four. So I've got one, two, three, four, five d orbitals. So what I'm going to do is, I'm going to take the s , I'm going to take all three of the p and I'm going to lop off two of the d 's. So now I'm going to have one, two, three, four, five, six and I'll leave three d 's by themselves and here's now sp^3d^2 . And now I've got six electrons-- one, two, three, four, five, six and now I have six equal energy states, which means I have to have six struts equally disposed in space and everything makes sense. All right.

Is this polar or nonpolar? Is this molecule polar or nonpolar? Nonpolar, good. Let's look at it carefully. What do we have? I've got sulfur fluorine. Whenever you see fluorine, that's the biggest electron hog on the Periodic Table so you know you're going to have a dipole moment on the bond. But I have six such bonds and they're symmetrically disposed in space. All of the fluorines lie on the surface of a sphere so the center of local negative charge is right on top of the sulfur nucleus, which is where the center of locally positive charge is. No electron displacement, no net dipole moment. OK. This is working. Let's try another one. Let's see. What do I have here?

I got bromine pentafluoride. So start-- let's go back to the VSEPR rules. There we go. OK. VSEPR rules. So I put bromine in the center and it's got seven valence electrons-- one, two, three, four, five, six-- and I'll put the seventh up here. And I'm going to bring in the fluorines-- one, two, three, four, five and just as before, there's one, two, three, four, five, six, seven. So you get a bond with each fluorine. A bond with each fluorine.

All right. So what do we have now? How many electron domains? One, two, three, four, five, six. Six electron domains as before, but one of them's nonbonding. I get to use the colored chalk again. All right. So this is one nonbonding domain and five bonding domains. One, two, three, four, five. One bond to each of the fluorines, five bonding, and the total is six electron domains, so that means octahedral structure. The total number of electron domains dictates skeletal structure. So man said put six. So I put the bromine here, and I go, one, two, three, four, five, six. That's the skeletal structure. Those are the bonds and space. Now I know one of those is a nonbonding domain. So where do I put it? Well, they're all equivalent. It doesn't make any difference. Nobody knows where the real x is in this world. So just for grins and chuckles, I'll make-- I'll put the nonbonding pair down here. And then I put the fluorines at the other struts.

And so now what's the shape of the molecule? You don't see these electrons. And in fact, this is only done by chemistry professors, the other ones, not me. I only do this to let you know what the books are doing because we all know that those nonbonding electrons are in here. They're right up tucked against the bromine. So if you walked in the room, what do you see? The bromine and the four fluorines lie in a plane, and then you've got this thing sticking up. So the molecule is a pyramid with a square base. So it's called square pyramidal.

And how about polar or nonpolar? Is there a net dipole moment? Yes or no? Yes, because we got the-- in the plane with the bromine, everything is centered, right? The fluorines pull out, but they all pull out symmetrically, but then there's the fluorine pulling up. So this is polar. This is polar and it's got a dipole setup like this. Good.

All right. Let's do one more. I'm having so much fun. I'm going to do as many as I can until I run out of time. This is the best part because colored chalk. That's the end. All right. Let's do this one. Tetrafluoroiodate. IF_4^- . This is cool because it's got a net charge, but it's a covalent anion. So same thing. Lewis structure. So we're going to put iodine first. Iodine and it's just like bromine. It's got one, two, three, four, five, six-- I'm going to put a seventh

electron here because I'm anticipating. Just to mix things up. See, last time I put the seventh electron up. I could put the seventh electron down, but it's got a net charge. How'd it get the net charge? By accepting an electron. So that electron's got to be here, too. So I'm going to indicate that that extra electron's here as well. So iodine's got eight already. So we're off to a roaring start. It's got eight and it hasn't even started bonding yet. So we know where we're going here. We're going to expanded octet land.

All right. So we need four more fluorine. I'm going to put four fluorines-- one, two, three, four-- and as before, so we've got bonds. One, two, three, four. And we've still got some other orbitals lying next to the iodine and they bond to nothing. So we have two nonbonding, and we have four bonding for a total of six electron domains. Again, octahedral structure. Octahedral skeleton. So I'm going to put iodine here. One, two, three, four, five, six. All right. Now comes the part-- see, we said-- this VSEPR consists of valence-shell. I've been doing a lot of valence-shell thought here, but I haven't said anything about electron-pair repulsion.

Now I'm going to show you how electron-pair repulsion comes into play. So I can put four fluorines around here in two different ways. One way is to put the four fluorines into play and then I have the-- these are called lone pairs. Why are they lone? Because they don't bond to a second. So sometimes chemists call these lone pairs. They're all alone. It's Friday and all alone. So here they are. So that's one way to put them, but there's another way to put them, and that's this one here. I'm going to go way over to the other side, and here's the other way.

The other way to put them is to put the pairs in the same plane next to each other, because if I put the pairs opposite each other, that's equivalent. It doesn't matter. It just-- whether on the x-axis or y-axis. So if you think about this as a globe, you think about the molecule as a globe, so in this case over there, I put all the fluorines on the equatorial position. You can think of this plane at the waistline as the equator. So these are all at the equatorial position, and the electron pairs are at the poles. So the lone pairs are at polar positions. In this case, the lone pairs are lying in equatorial positions.

Now what do you know about electrons when they get close together? They repel. And it's a really strong repulsion. That's the Bourne exponent. Electron-pair repulsion is really strong and it'll jack the energy way up. So if you've got a choice of putting electrons here close together or here farther apart, which one is going to be the lower energy condition? This, because this minimizes electron-pair repulsion, and in fact, this is the one that's favored. So by using valence-shell thought, bonding, nonbonding skeletal structure, and then finally electron-pair repulsion, you conclude that this is the correct form, and so you don't see these electron pairs because we know they're really tucked in tight here.

What do you see if you do a spectroscopic analysis of this thing? All five atoms lie in a plane. So this molecule is planar. It's planar, just as the sp^2 ethylene was planar, but this is planar for a different reason. Is this polar or

nonpolar? See, it's negative. It's not negative. If I were a cation, I'd be drawn in like this. So is it polar or nonpolar? Where's the net displacement? The fluorines are pulling away from the iodine, but they're pulling away symmetrically. So there's no net dipole moment. It's cool, huh? No net dipole moment, but net negative charge. I love it. It just-- it's the best. OK.

So now let's look at some others. We've looked at CF_4 . We know CF_4 . CF_4 , you could do this one in your head. This would be just like methane, right? This would be tetrahedral. I'm not even going to write that out. This tetrahedral and it would be sp^3 hybridization and we just did IF_4 minus. See the stoichiometry. It's four fluorines. Four fluorines with something and it's not the same architecture. So stoichiometry doesn't dictate molecular architecture. Very important. You can say, well, it's a net charge. A net charge doesn't do-- it's this. This is square planar. Yeah, so I'm going to add that. It's not only planar, it's planar and it's a square, whereas the other one is trigonal planar. So this is square planar.

So I'm going to do another F_4 , just to show you, make the point that this thing doesn't-- so I'm going to do sulfur hexafluoride. You say, gee, sulfur hexafluoride. I bet it's just like carbon because carbon, sulfur-- no. Let's look carefully. So I'm going to put the sulfur here and we're going to put-- how many electrons? One, two, three, four, five, six. OK. And then we're going to bring in the fluorines, one, two, three, four. So each of the fluorines is going to form a bond. Do I have to do this? Can I just do this? Yeah, that's enough. I'm not-- I'm raising-- we're accelerating now. Now I'm going to stop talking and I'll just think about it and you'll get it.

All right. So here we go. Bonding domains. One with the fluorine, two with the fluorines, three with the fluorine, four with the fluorine and there's two electrons, and they'll just sit together in a nonbonding domain. So how many domains? One, two, three, four, five. Five electron domains. Four bonding, one nonbonding. It all adds up. Good. But now I got to put five struts. This gives me the skeletal structure. As before, skeletal structure. So I've got to-- how do I put five struts around something uniformly in space? One, two and I'll put three in the plane. Three in the plane. So three at the equator and two at the poles. That's the skeletal structure. This is called a trigonal bipyramid. Because I can make a pyramid above, right? I got a tent with three corners. So that's a pyramid with three faces. So that's called trigonal. It's trigonal-- three-- and then I've got it above and below. So that's trigonal bipyramid. See, you're going to learn all these new words. When you go out tonight, you're going to be able to impress people to the point they'll walk away from you. All right.

So now let's start putting the fluorines-- again, we've got choices here. They're not equivalent. So one possibility is to put the lone pair at the equator and then the fluorines go here, and the other possibility is to put the-- let's see. I've got room up there. The other possibility is-- let's make the trigonal bipyramid-- is to put the lone pair at a polar position and put all the fluorines in the equatorial plane. And so now we have to think about this for a second and try to convince ourselves which situation has greater electron-pair repulsion. So the way to think about it is, in

which position does that lone pair come into more contact than otherwise?

The other thing to know is-- I didn't say this earlier, but it just occurred to me-- is that if you've got a choice of bonding electrons versus nonbonding electrons, nonbonding electrons want more room. Why? Because they're not tightly confined. When you get something in a bond, it's on axis between the two atoms, and so things are much more tightly confined. So these things want a lot of room. So if I'm sitting here, I interact with this fluorine and this fluorine-- these two are far away. If I'm here, I interact with three fluorines. So it turns out the equatorial position is favored. Give more space nonbonding-- place nonbonding domains at equatorial positions in trigonal bipyramid. That's what number five means. If somebody asked you to test the microphone, instead of saying number nine, number nine or one, two, three, say, place nonbonding domains at equatorial positions and a trigonal bipyramid. And they'll ask you to step away from the microphone. OK.

So this structure here is, in fact, the shape of the molecule. You don't see this. And this molecule is called seesaw, like a teeter totter. You remember, when you were kids. You don't remember that.

OK. So how about polar versus nonpolar? It's obvious. If it's asymmetric and you've got the fluorines and the non-pair, this is definitely polar. And where's the net dipole? Where's the negative side? Well, along this axis, it's pulled right to the center. So along this axis, clearly the fluorines are pulling so you have something that looks like this. And how about the energetics? How do we get five of these? If we've got five struts, it means we must have five hybridized orbitals of equal energy. Well, we got six before and before that, we had four. Before that, we had three. Well, now we're going to get five.

So how do we make five orbitals? So in this case I start-- I want s, p, d, and I want five. So one plus three is four plus one is five and that gives me sp^3d . Gives me five and then, what do we get? One, two, three. It's good. OK. Now-- so we've covered a lot. All right.

So some parting comments. We've been sticking other things-- what we could do is we could do this one. We could take carbon and I'm going to go now sp^3 . only instead of putting hydrogens, I'm going to put carbons here. So what happens now? One, two, three, four. I'll put another carbon here. One, two, three, four, et cetera. So this is sp^3 , all carbon, and this is called diamond. This is called diamond sp^3 . So these are all sigma bonds, all sigma bonds everywhere, sp^3 . And these are very, very strong bonds and these are very tight bonds-- and one of the properties of diamond is that it has a very high refractive index because high energy-- and these are tight, tight bonds. So the electrons aren't easily excited. It's transparent, divisible light, which means that when light shines on this, there's no expectation, but there is bending. So it's very, very high refractive index and a high refractive index is what makes this such a troublesome compound because it concentrates light.

So you can imagine there's somebody in a dimly lit cafe, and she's sitting over in the corner with a candle in front,

and she's wearing a diamond stud, and it takes those few eV of energy. It concentrates, it shoots it in the room, and some unsuspecting fellow gets hit in the eye, is attracted to the person wearing the earrings, and that's when the trouble begins.

So this is very, very dangerous, very dangerous, and it's a consequence of these strong sigma bonds. But there's another way. There's another way that we can bond carbon. We can make sp^2 hybrids. So if we make sp^2 hybrids, that now looks like this. The carbon lies in the plane, 120 degrees, which means we're going to end up with something that looks like this, et cetera, et cetera. And this is graphite. This is graphite. And now we have a mix of sigma bonds in the plane that are very strong. In fact, they're every bit as strong as these bonds, but there's a fourth. I said, you always have to find four struts off the carbons. So where's the fourth strut? It's the pi bond that comes out of the board. So there's a pi bond here and a pi bond here and a pi bond here and what we're going to learn later is that these things are delocalized.

Because of the spacing here, the electrons can form a bond here that then forms a bond here, but it can actually resonate in such a way that the electrons can move everywhere throughout the crystal of graphite and this is what gives graphite its electrical conductivity, whereas diamond is a strict insulator. So these are very strong bonds in the plane, but not between planes, which gives graphite its lubricity. It's a dry-lock lubricant and so on. OK. So this is very good. So we have a mix of sigma and pi bonds. And it's an absorber and it's dark. It's an absorber so it appears black, whereas this diamond is transparent to visible light.

And if you look in your book, this is all categorized for you, everything that we've done here with VSEPR. All right. Look at this. All nice big system. OK. So here's diamond, which I was trying to draw over here, four struts sp^3 hybridized. this is graphite sp^2 , and these are pi bonds in between. OK. Now suppose this fellow asks this girl out, and things get serious and they get-- I'm going to use an adverb here-- really serious and he decides maybe he wants to make a commitment and he wants to buy her a ring, an engagement ring, and the custom is that the engagement ring has a stone in it.

Now being an MIT student, he knows that he wants to have symbolism in that stone. Now the custom is to get diamond, but it turns out that if you look carefully, the stable form of carbon at room temperature and atmospheric pressure is not diamond, it's graphite. So if he wants to give his love this stone that symbolizes eternal commitment, wouldn't he choose the stable form and not the one that's metastable and doesn't represent the natural state at room temperature? That's the truth. That's the truth. Graphite is the stable form. I leave it up to the young man to persuade the young lady that the gemstone is going to have graphite in it and not diamond. OK. Now something very interesting happened.

What I wanted to play for you as you're leaving today is an old Beatles song, Lucy in the Sky With Diamonds. And

what happened-- before I get to it, what happened was-- and this is even in today's Le Tech if you look on the second page, "New Fossil Skeleton from africa Predates Lucy." It was just discovered. I'm not kidding you. This isn't a joke. I heard this on NPR last night. So it's a female skeleton of a pre-hominid, 4.4 million years old, which is a million years older than Lucy. Lucy was discovered in Ethiopia in 1974 and considered to be 3.18 million years old. This new one is 4.4 million years old, about four feet tall, and it just turns the clock even farther back.

The interesting thing is the name. How did the skeleton discovered in 1974 get the name Lucy? Interesting story. There were two American archaeologists that were head of the team that did the dig, and they were so excited the night they found this pre-hominid skeleton. They knew it was a female, and they hadn't dated it yet, but they knew that it was roughly way, way back, but it was definitely female. So they were back at the campfire, kicking back, drinking, probably heavily, and singing songs, and one of the songs they were singing was Lucy in the Sky With Diamonds and that night, they named the skeleton Lucy.

But the story doesn't end there. Where did this Lucy in the Sky With Diamonds come from? John Lennon, one of the-- you may have heard of this group or not, but there was this group called The Beatles. So one of the lead composers was John Lennon. His six-year-old son came home from school one day with a piece of artwork, and the father asks, what is this? He says, it's a painting of my classmate Lucy, and it was Lucy and he had diamonds in the sky. So it was Lucy in the Sky with Diamonds, and then they wrote the song about it. The real Lucy just died this week. She had Lupus. She died at the age of 46. And all of these things have come together just because we had this lesson today about covalent bonding, which took us to this, which then took us to Ethiopia, which then takes this up here. and so with that, I think we'll go for an early dismissal, and I'll wish you a pleasant weekend.