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**CATHERINE  
DRENNAN:**

That's today's handout. We have valence bond theory and hybridization. So some people ask, OK, now you're going to tell me everything you just learned. It's not really right and there's something else that's better. No. All of these theories are good theories. They all do a very good job predicting the properties of molecules, but they all have different strengths and weaknesses. And I think in terms of what they're useful for, molecular orbital theory is very good in terms of thinking about energy levels. It's very good about thinking about bond orders or predicting whether something's going to have an unpaired electron or not.

Valence bond theory and hybridization are really good in terms of thinking about shapes of molecules. So not so much about energy levels, but shapes. So all of these theories are very, very useful because we want to think about how atoms come together to form these molecules and what are the properties of the molecules. So these theories brought together really give us a wonderful picture of this. And I really like valence bond theory and hybridization because I like shape. I determine shapes of molecules, complicated molecules, for a living so I'm a big fan.

But I will say that when I taught this the same lecture last year, I announced to the class that I had had a dream where all these atomic orbitals were coming together and trying to make other kinds of orbitals. And I realized that, perhaps, that was a sign that on Friday I should start teaching thermodynamics which is what we're going to be doing. We're going to start on thermodynamics on Friday. And last night, I had another dream about orbitals. So I think this is some more orbitals and then we go to thermodynamics. And I remembered my dream because at that moment, my giant dog jumped on top of me as I was sleeping to wake me up and realize that thermodynamics needs to come pretty soon. OK.

But one more theory, valence bond. This is not so bad. OK. Bonds result from the pairing of unpaired electrons from the valence shell of atomic orbitals. That's it. That's it. So we have one, we bring in another so we can make molecular hydrogen, H<sub>2</sub>, because they each have one unpaired electron and they come together to form a bond. I like theories that you can put

on a magnet on your refrigerator. That's a good theory to me.

So also as part of valence bond theory, we have some names of bonds. And we've been talking about sigma molecular orbitals and pi molecular orbitals. And now, we're going to talk about sigma bonds and pi bonds. So we had orbitals in MO theory, valence bond theory, we now have bonds. Sigma orbital is cylindrically symmetric about the bonding axis. Thank goodness they didn't define them differently. That would have been a nightmare. So we have sigma orbitals that are cylindrically symmetrical about the bond axis and sigma bonds are cylindrically symmetrical about the bond axis. So no nodal plane along the bond axis. Good. We should be able to remember that.

So with pi bonds, we have electron density in two lobes with a single nodal plane along the bond axis. So again, with pi orbitals, we had more-- it wasn't cylindrically symmetric. So this we should be able to remember. A couple other things about sigma bonds and pi bonds, a single bond is a sigma bond. So when there's one bond, it's a sigma bond.

So what's a double bond? A double bond is a sigma bond plus one pi bond. So if it's double bond, it's got two types of bonds, sigma and pi. And what do you think a triple bond is? Sigma bond and two pi bonds. So you got a triple bond like nitrogen, you got two pi's. Hey, it's really a good life when you have a triple bond. All right. Single bonds always going to be sigma. Double, sigma and pi. Triple, sigma and two pi bonds. OK.

So now we're going to hybridize our orbitals. And we're going to talk about electron promotion, as well. So start with carbon, carbon based life. Carbon is really important and if you are an organic chemist, and by organic, it means studying things with carbon, you care a lot about hybridization. And the stuff I'm teaching you today, you'll see a lot if you go on to take Organic Chemistry 512. So carbon, such as one in methane, so we have our methane molecule here. The carbon has four unpaired-- can form bonds with four electrons, but to do so we need to do something with our electrons.

So carbon comes in, it has two electrons in its 2s and it has two electrons in its 2p's, p orbitals, but we want to form four bonds. And in covalent bond theory, every bond you bring an electron from one atom, an electron from the other, and they pair and that forms a bond. So we don't have four unpaired electrons to make four bonds with this configuration of electrons, so we can talk about promotion of an electron from here up there. And if we do that, now we have our four single unpaired electrons ready to make four bonds. And carbon does like to

make four bonds. It does it quite often. So that's electron promotion.

To form those four bonds, a 2s electron is promoted to an empty 2p orbital. And then, we can hybridize our orbitals and that means that we want to give all our orbitals some s and some p character. So here are our hybrid orbitals and let me show you the nomenclature. So we're talking about n equals two. So we have a two. We have s character and we have p character and we're using three p orbitals to make our hybrid orbitals. So we are going to make a  $2sp^3$  hybrid orbital. And we're going to make four of them because we've used four atomic orbitals to make them. So if we are using four, we need to make four.

So let's kind of take a look at what's going on here. And we'll say that these molecular orbitals differ only in terms of their orientation in space. So they don't have different shapes, they're just oriented differently. So here we have our 2s, remember it's symmetric, and we have our three p orbitals, and they're all the same except that they're all oriented differently in space. And when we bring these together, we form four hybrid orbitals and they kind of look like turtles, but they're turtles oriented differently in space, but otherwise they're the same. So those are our  $sp^3$  hybridized orbitals.

So carbon has this  $sp^3$  hybridized orbital and it has four unpaired electrons available to form bonds with four hydrogens. So let's bring our hydrogens in to form our bonds. And each hydrogen brings with it its one electron. So now we have two electrons in all four of our hybrid orbitals. And we can think about where the energy came from. I just moved that electron, I didn't think about it. I'm like, yeah, that just goes up here. So where did the energy come from to do that? And that is, it came from bonding. So this molecule now is more stable because it's bonded. Methane isn't quite a stable molecule. That's another problem in and of itself. So the bonding allows you to do that. You get back from this bonding.

So let's look at those bonds then that are formed that make that electron promotion worthwhile. And so you're forming a bond between the carbon and the hydrogen, you're forming for them, and you're forming single bonds, they're sigma bonds, and the bond is formed between the carbon's  $2sp^3$  orbital and the hydrogen's 1s orbital. Hydrogen can't hybridize, it's got one, 1s orbital. That's all it's got, can't do anything else. And that gives you a bond then, a sigma bond, that you'll see this a lot and you'll write this a lot.

This is how we're going to name that sigma bond. So we're going to say sigma. We're going to have a parentheses. Identify the element, it's carbon. N is 2. Type of orbital,  $sp^3$  comma

hydrogen, the name of the other element, and its orbital, which is 1s. So when it asks you to name the type of bond, this is the complete answer that we're looking for. And we'll have more practice on this.

Now we can also think about the shape that this molecule would have. What is the angle here between this hydrogen and that hydrogen and frankly, between any of the hydrogen carbon hydrogens? Yup, 109.5. And the name of that geometry? Tetrahedral, right. So  $sp^3$  gives you a tetrahedral based geometry here. All right.

So now let's get more complicated. Let's bring two carbons in. So we have ethane, two carbons, six hydrogens. So this also has its carbons are  $sp^3$ , and this is what we saw before for methane, but now I'm going to rotate this around and that's one carbon, but we need another carbon, but first we can think about this one carbon. So one of the carbons of ethane it would have this 109.5 angle. It has four unpaired electrons available in its four hybrid orbitals to form interactions, one with carbon and three of them with hydrogen. And then we need another one of these so we'll bring that in and it comes in with its set of hybrid orbitals and its set of electrons. And we form a bond between them. And the bond we're going to form is a single bond, a sigma bond.

And now let's bring in our hydrogens. So we had six hydrogens, three for each carbon. And so there are now two types of bonds. We have the carbon-carbon bond and we also have the carbon-hydrogen bonds. And so the carbon-carbon bond, which is a sigma bond, is sigma ( $C-C$  it has carbon--  $2sp^3$ , the other carbon is the same,  $C2sp^3$  and then the bracket. So that's that sigma bond. It's a single bond. And here is our ethane molecule.

And then we have our carbon hydrogen bonds, they're also sigma. Please don't give me pi bonds to hydrogen. It only has that one electron tapping with two electrons. It doesn't want to do anything complicated. It doesn't have p orbitals, just that one s. So sigma  $C2sp^3$ ,  $H1s$ . And now we have defined this molecule so we brought together two tetrahedral centers and formed this molecule with a single bond.

So let's talk about nitrogen. Nitrogen, also again, very important. So here we have five valence electrons. What about electron promotion? Should I do it? No. Because I mean, you could put it up here, but it can't make any more bonds so it doesn't really matter. So it doesn't occur because it would not increase the number of unpaired electrons to form bonds, but we can hybridize. So we can still hybridize our orbitals and we can get four hybrid orbitals, because

we're going to use our 2s and all three of our 2p orbitals. So we'll get the same set of hybrid orbitals. But this time, one of them has two electrons in it. So it's not ready to bond, it's happy according to valence bond theory. And these are our lone pairs.

But we can form three bonds with these guys so let's look at an example,  $\text{NH}_3$ . So now we have our lone pair, it's in this orbital up here, and then we have three orbitals available for bonding, each with an unpaired electron ready for the three atoms of hydrogen to come in. So we bring in our three atoms of hydrogen, each came with an electron, and now you can tell me with a clicker about the angle and the geometry of this molecule.

Let's just do 10 more seconds.

All right. So this is back to VSEPR again. So we have an angle here. It's based on an  $sp^3$  system, one lone pair, three bonded atoms. So it's based on our 109.5, but those lone pairs make for bad roommates and they're pressing all of these hydrogens together and so the angle is less than 109.5 and we name this structure based on the atoms we see, not the lone pairs, so this is trigonal pyramidal. And so here we have it here so we're naming it without thinking about the position of those lone pairs that are pressing down on the bond so it looks trigonal, like a triangle, but it's also a little pyramid. So VSEPR-- VSEPR and hybridization, they just go right together. It's awesome. OK.

So we can also name the type of bond. So our nitrogen had  $2sp^3$  hybridization and our hydrogen just 1s, it's a sigma bond, it's a single bond. So we named that sigma  $N2sp^3$ ,  $H1s$ . So nitrogen, now we're going to go back to carbon-- sorry, to oxygen-- and think about hybridization of oxygen.

Oxygen. Should I do an electron promotion? No. It's not going to help me. It's not going to create any more electrons available to form bonds, but I can hybridize and I can get the same four hybrid orbitals, our four  $2sp^3$  orbitals, but now two of them have two electrons in them and two are available to form bonds. Oxygen loves to form bonds with hydrogen and form water, most of the planet is water. There's a lot of water and water is really important for life so it's great that oxygen and hydrogen get along so well.

So the oxygen, again, has two lone pairs which are here. You bring in our hydrogens and they come with one electron. And again, now, it's still a steric number of four systems, so it's less than 109.5, and it's actually a lot less than the nitrogen because you have those two lone pairs that are just taken up so much room and squeezing together these hydrogen atoms over here

creating this 104.5 angle.

So here we have our oxygen molecule with its two lone pairs and its two hydrogens, and what's the name of that geometry? Bent. And again, we have these polar bonds that create a dipole so it's a polar molecule, which is very important in life. And we can name that bond. It's a sigma bond. It's made up of oxygen  $O2sp^2$ ,  $H1s$ . All right. So that's  $sp^3$  hybridization. Now let's talk about  $sp^2$  hybridization.

So  $sp^2$  hybridization. So back to our atomic orbitals. And now, we're not going hybridized all of our orbitals. We're just going hybridize our  $2s$  and two of our  $p$  orbitals. So we'll hybridized these guys and we will form three hybrid orbitals and we will still have one un-hybridized orbital. We will have  $2p_y$  left alone. So let's see what this does. How is this hybridization useful?

So let's talk about boron. Boron has three unpaired electrons, but they are not all available right now to form bonds, according to valence bond theory. So here we do want to do an electron promotion to put one of them up here so that now all three are available to form bonds. And we can again, hybridize these three atomic orbitals and form three hybrid orbitals. So we have three  $2sp^2$  hybrid orbitals and then we still have our  $2p_y$  orbitals so don't forget to mark it. It seems lonely. It's over here, but it's going to be important later so don't feel bad for it, yet. All right.

So boron-- let's think about these hybrid orbitals and how this gives us the structure that we know occurs when we have boron. So boron now has its three  $sp^2$  orbitals and these are going to lie in a plane and they're going to be as far apart from each other as they can to minimize electron repulsion. And if you're in a plane, then you need-- far apart as you could be is 120 degrees. And this is what gives us our trigonal planar geometry. So we saw that boron formed these trigonal planar complexes before and again, they're trigonal planar because they're like a triangle and they're in a plane, trigonal planar.

And we can now bring in our hydrogens. The hydrogens come with an electron so we have an electron for them and there we have our structure. We can also name that bond. So again we have single bonds. So sigma B, for boron,  $2sp^2$ ,  $H1s$ , and there are three of those.

Carbon-- carbon can also do this. We talked about carbon being  $sp^3$ . Carbon can also be  $sp^2$ . Hybridized carbon is amazing that's why life is based on carbon. Carbon can do lots of things. So again, we're going hybridized two  $p$  orbitals, one  $s$  orbital to give three hybrid

orbitals and we have our  $2p_y$  over here in the corner, but don't feel bad for it, it's going to do something useful. So we now have three electrons in these hybrid orbitals and now we have one electron in our  $2p_y$  un-hybridized orbital, as well.

So let's see what carbon, with this kind of arrangement of orbitals, can do. And again, we're going to have trigonal planar geometry for our  $2sp^2$  hybrid orbitals. So we have carbon there and these are all in a plane, but now coming out of a plane toward us is this  $2p_y$  orbital. So it's coming out 90 degrees away from the trigonal planar geometry.

So an example of  $sp^2$  hybridization is in this molecule,  $C_2H_4$ , and it has a double bond, which means if it's a double bond, it has what kind of bonds in it? Sigma and pi, right. So one sigma, one pi bond. So here now, and this is the trigonal planar geometry. It's supposed to be in a plane, but you can't really see it if it's really in a plane. But 90 degrees away from that plane is our  $2p_y$  orbital. We brought in our two hydrogens so this carbon here, is carbon is there, two hydrogens are there. This would be 120 degrees.

Now we're going to bring in another one. That's the one over here. It comes in with its carbon. It comes in with its two hydrogens forming these single bond, sigma bonds, between the carbon and those hydrogens. And now we're going to form a carbon-carbon bond. This carbon-carbon bond is a sigma bond and so it's  $C_2sp^2$ ,  $C_2sp^2$ , but we're not done. We said this is a double bond, so that's our sigma bond, but we need our pi bond.

And now  $p_y$ , our un-hybridized orbital, is extremely excited because it can form the pi bond. So we form a pi bond and that's formed by our  $C_2p_y$ ,  $C_2p_y$  un-hybridized orbital. And we also have four CH bonds and those are single bonds, those are sigma bonds, and so they're formed by our  $C_2sp^2$  carbon and hydrogen  $1s$ . And there are four of those. So that's an example of  $sp^2$  hybridization.

And one thing that's very important, and here you can see what that molecule looks like-- doesn't all fall apart-- so this is a double bond, these smaller kits don't let me make double bonds, so I have a sign double bond. And you can see the angles and the geometry of this molecule. And another property of something with the double bond like this is that it's not really free to rotate. So when you have these two kind of points of attachment, when we have these orbitals forming between your on hybridized p orbitals, that does not allow for rotation around the double bond. So if you're an organic chemist wanting to make a molecule that's going to be rigid, if you put a lot of double bonds in it, it can't twist and turn very well. It's often very rigid

which is useful. So we'll stop here and we'll finish up on Friday sp 2 hybridization.

For the clicker question, the bone over there on that is also the same as the one on the board. The one on the board is written with atoms in it and it has squiggly lines to abbreviate so make sure that your answer is consistent with the picture on the board, as well. How we doing? OK. All right. Let's just take 10 more seconds. Remember this is a clicker competition so we want to get the right answer in for your recitation.

**AUDIENCE:** [SIDE CONVERSATIONS]

**CATHERINE** All right.

**DRENNAN:**

That's pretty good because that's the right answer. OK. So let's just take a look at this for a minute. First let me explain. Let's settle down, quiet down. Let me just explain the diagram, too, because you'll be seeing these diagrams. So when you just have a bond, a line, and there's no atom indicated, that means it's carbon. Organic chemists, I think, came up with this rule. Carbon, they just said if nothing's indicated, of course, it's carbon. Carbon is such an important element, we don't really need to say more about it than that.

So you could interpret this diagram, you have a carbon double bonded to another carbon. And then up here, there's a carbon in that ring so I just put, in this diagram, carbon with squiggles, you'll see that sometimes. That means that there's more atoms there, but I'm too lazy to draw them. And on this side, there's a carbon, but there's more atoms there and I'm too lazy to draw it, another squiggly. And then we have the double bond so there's a carbon down here, as well. It wasn't indicated, just the line in the drawing. And you have to predict how many hydrogens Hydrogens are often not indicated. This one is indicated. There are other hydrogens in this drawing that are not indicated. You need to figure out where they go and the material we're doing now is going to help you do that. And then I also drew something and another squiggly because I was too lazy to draw the rest.

So these are different kinds of diagrams that you'll see that all kind of mean there's more than one way to kind of write the same structure. So this particular molecule was used to treat schizophrenia in the 1950s and key to the usefulness of the molecule was that double bond. As we talked about last time, double bonds restrict movement. You can't twist around the double bond. And so if you had exchange and you had this group over there and the hydrogen

over here, it wouldn't be an active molecule. So this double bond fixes the orientation of those other atoms such that it was an active molecule and could be used as a pharmaceutical to treat schizophrenia.

So in terms of the bonds then, we have a double bond which means we have one sigma and one pi bond. And so the sigma bond down here, we had to know what the hybridization was. And here, those carbons are bonded to three other atoms and so it would be  $sp^2$  carbons and also with the double bond  $sp^2$ . And then we also have a pi bond and pi bonds are made up of non-hybridized orbitals, our  $p_y$  or our  $p_x$ . And so those are the ones that make up the pi bond. In all the other variations, some you had two sigmas, that's not right, we have a sigma and a pi so most people figure that out. They picked the ones that had those categories for the most part. And then you had to pay attention to whether it was  $sp^2$  or  $sp^3$ . And then here, this one, the pi bond is not made up of hybridized orbitals, it's made up of the atomic orbital leftover.

So a lot to look at that particular problem, but this is really good practice for the exam, which is coming up a week from Monday. There's going to be lots of hybridization and today, we're going to post extra problems for the exam so you have, really, a whole week to start getting ready for this exam and to keep up with the new material. And so extra problems and an old exam are also going to be posted later today. All right.

So let's just finish with hybridization now and this is good because this is all stuff that's going to be on the exam. And also, the instructions for the exam are attached to today's handouts and, of course, remember no makeup exams and clicker competition. So I'm not really going to go through anything in the instructions. It's very similar to last time so you can take a look at that and see on the material, the material starts with the periodic table, periodic trends, and goes through the material that I'm going to finish with partway through today's lecture. So at the end of hybridization, that's the end of exam two material.

So we're going to finish our lecture notes from last time so pull those out. And then we're going to move on to thermodynamics. So once we hit thermodynamics, that's exam three material, so we're almost done with exam two material, a week in advance to get ready for the exam. So that's great. Lots of time to review exam two material and let's see if we can have an A average on this exam. That would make me really, really really, happy. I would wear my periodic table leggings again if we could get an A average on the exam. I'm just saying, I'd be very excited. OK.

So we better finish up that material so that you can get started, get ready for this. So we talking about valence bond theory and hybridization and forming these hybrid orbitals. And valence bond theory is this idea that if you have a single electron in an orbital, it's available to form a bond and bonding happens when two atoms bring together single electrons and those pair up to form a bond. So we talked about electron promotion before and let's just review what that meant. So if you have an empty orbital, you can promote one of your electrons to that empty orbital. And so now we have four valence electrons after electron promotion so that we have more possibility of forming bonds.

Now if you don't have an empty orbital, you can't promote your electron. If you do have an empty orbital, you can promote it more single electrons available for bonding. If you don't have an empty orbital, there's nothing to do with that. So that's the trick to electron promotion. All right.

So now we have one electron in each of the four valence atomic orbitals that we have for carbon, but we're going to only hybridize two of them. We saw already last time that we can hybridize all four orbitals and have  $sp^3$ . We can hybridize just three of those orbitals and have  $sp^2$ . Now we're going to see that we can hybridize just two of those orbitals and have  $sp$ . So carbon is really amazing. It can do all three of these kinds of hybridization. That's why carbon based life forms are able to exist and do so much.

So we're going to now hybridize our  $2s$  and our two  $p_z$ .  $Z$  is just special and so it gets to hybridize with the  $2s$  leaving two of the other orbitals just by themselves. So we're going to form two hybrid orbitals. Again, if we hybridize two atomic orbitals, we're going to form two hybrid orbitals. And if we hybridized  $2s$  and  $2p_z$ , we're going to get hybrid  $2sp$  orbitals. And we'll have our  $2p_x$  and our  $2p_y$  just the same as always. All right.

So we can think about this in terms of shapes, as well. So we have, again, our spherically symmetric  $s$  orbitals and our  $p$  orbitals and we have the three of our  $p$  orbitals that are the same shape they just differ in orientation in space. And so we're just going hybridize our  $2p_z$  and our  $2s$  and so we'll have our kind of funny looking I think of them as turtle shaped or hybrid orbitals and then we also have our  $2p_x$  and  $2p_y$  orbitals the same as always.

So what are we going to do with our two  $sp$  orbitals and our one  $2p_x$  and one  $2p_y$ ? Well, we can form a pretty cool molecule with it. So we're going to form something that has a carbon-carbon triple bond. So this is  $C_2H_2$ . So now in cyan is the  $sp$  orbital, the hybrid orbital, that is

formed on this carbon here. And then we have a  $2p_x$  orbital here in the plane of the screen. And we have a  $2p_y$  orbital coming out toward us. And, of course, our  $2p_z$  orbital had been hybridized with the  $2s$ .

So here we have this structure. We're going to bring in our other carbon, and the other carbon has the same situation going on. And we can form a bond between the two carbons with our  $sp$  orbitals. And we can form, also, with the  $sp$  hybrid orbitals, bonds to hydrogen. So we have two hydrogens, one over here and one over there. So what is the angle between these hydrogens here? Yeah. So that's  $180^\circ$  and, again, we have an example here of the molecule we're going to build that's going to have a triple bond.

So now let's name those types of bonds or as sometimes in problem set, it will say describe the symmetry of the bond. And what it means by that is the following. It means that, that's either say name the type of bond or describe the symmetry. There's multiple ways to ask the question and this is the answer to those questions. So the bond that's formed, the first one that's formed, between the two carbons is a sigma bond and it's formed between the  $sp$  orbital. So  $C2sp$ ,  $C2sp$ . That's the first one. But this is a triple bond so we have two more bonds to form. And this is where our atomic  $p_x$  and  $p_y$  orbitals will come in. So we're going to form the next bond, which is what? A sigma or pi? Pi bond and that can be between our  $x$ ,  $p_x$ , orbitals, so pi  $C2p_x$ ,  $C2p_x$ .

And now, we have the  $2p_y$  orbitals, as well, and that allows us to form our triple bond. So we're also going to have a bond pi  $C2p_y$ ,  $C2p_y$ . So again, with the triple bond, we're going to have one sigma and two pi bonds. The sigma is formed from the hybrid orbitals and the pi bonds are formed by the  $2p_x$  and  $2p_y$  orbitals. So carbon, really impressive. Carbon can form of these three types of hybrid orbitals. It can form molecules with single bonds, double bonds, and triple bonds.

So let's just have a little cheat sheet to think about that. So again, this is for carbon hydrocarbon molecules, like we've looked at so far, that have two carbons in them. So let's look at carbon in  $C_2H_6$ , so that's over here. What is going to be the hybridization when you have a carbon that has a bond to another carbon and a bond to three hydrogens here? What hybridization?  $sp^3$ . That's right. And it's going to have what kind of a bond-- single, double, or triple? It's going to have a single bond and it's going to have tetrahedral geometry around both of the carbons. So both of these carbons are going to have tetrahedral geometry, which is not a blank in your note, but what's the angle?  $109.5^\circ$ , right. Thank you.

So we have carbons  $C_2H_2$  are going to be  $sp^2$  hybridized. And what kind of a bond are they going to have between the two carbons? That will be a double bond. And what is the geometry of that? Right. Trigonal planar. And here, you have to pretend this is a double bond, my model kit didn't come with double bond possibilities and I have to hold it very carefully, but if I hold it very y-- oh-- the bonds are still there. You'll see that the angles are 120 and so this is trigonal planar geometry at each carbon. We didn't tape that one. You can see there's scotch tape all over the others. It was not a happy molecule. OK.

So now,  $C_2H_2$ , what kind of hybridization?  $sp$ , that is our friend the triple bond and we're going to have linear geometry and 180. So both carbons have linear geometry. That works. It's always triple bonds are much more stable, they don't fall apart as much. OK. So that's a cheat sheet for carbon. Now if you're thinking about nitrogen or oxygen, those often have lone pairs on them. Carbon likes to form all bonds, it doesn't care double, triple, single, whatever, but it doesn't really have a lot of lone pairs on it. But oxygen, nitrogen have lone pairs and whenever you have lone pairs, you have to worry about what the geometry is because the geometry gets named based on the atoms that you do see, not the lone pairs. So this cheat sheet works for carbon without lone pairs. If you have lone pairs, you've got to go back to your vesper and think about what the names of the geometries are. OK.

So rules, and I posted this on Steller for the problem set that was due today, and so very simple for determining hybridization. And this is the kind of equation that will not be on an equation sheet for an exam, you just need to know that. So in determining hybridization of an atom in a complex molecule, you're going to be thinking about the number of bonded atoms plus the number of lone pairs is going to be equal to the number of hybrid orbitals. So now, clicker question, what is the hybridization of an atom that has exactly two hybrid orbitals? All right. 10 seconds. Yes. Right.  $sp$ . So we can take a look at that two hybrid orbitals are formed by one at 1s orbital and 1p orbital and if you have two things bonded and no lone pairs, that's what you would get. Three hybrid orbitals would be  $sp^2$  and four would be  $sp^3$ .

So again, you're going to just be thinking in these problems about how many atoms are bonded to that central atom and how many lone pairs do you have. And that's going to then let you figure out what your hybridization is. And we have one exception which is that if an atom has a single bond and it's terminal on the edge of the molecule, then we're not going to hybridize it. So we can now take a look at an example of this and this is going to be another-- yeah, just keep your clickers out. We've got a whole bunch of clicker questions coming at you kind of in a

row here. And if we have this molecule, it has a central carbon and three terminal atoms. Now help me figure out what kinds of bonds this will form. So which one of these has the correct bond types for this molecule? All right. Make a decision. Let's just take 10 more seconds. Interesting.

I think some time I re-poll, but I think that we'll just kind of go over this one and then we'll-- do you want to go ahead and show the answer and then-- this is--

**AUDIENCE:** [SIDE CONVERSATIONS]

**CATHERINE DRENNAN:** And if it wasn't a clicker competition, I might have you discuss it more and re-poll, but it's a competition. So let's go over it. So this one isn't in your notes so if you want to write it at the bottom of the page, we'll go over what the answer is. Hopefully there's a typo in there, but we'll see when we go through. All right.

So let's take a look at this molecule. Hydrogen is terminal and single bonded, but we've already talked about hydrogen so we kind of knew that. Oxygen is terminal, but it's double bonded so we need to hybridize it. Cl is terminal and single bonded so we don't hybridize this and we don't hybridize hydrogen, never hybridized hydrogen. OK.

So let's look at the kind of bonds that are formed. So we have a sigma bond, single bond, this carbon is  $C2sp^2$ . It's bonded to three different things and it has no lone pairs so that makes it three that there's three things so we have three hybrid orbitals, which is  $sp^2$ . Our hydrogen is just  $1s$ , it's always just  $1s$ . That's all it is. So let's look at this bond now. So we have a single bond between our carbon that is  $2sp^2$  and then, we also have this oxygen. We do hybridize it because it has a double bond and it has two sets of lone pairs and it's bonded to one atom, so it has three hybrid orbitals, so it's  $sp^2$  just like the carbon. And then we have a pi bond and the pi bond is made up of atomic orbitals, either  $2p_x$  or  $2p_y$ .

Chlorine is single bonded so we're not going to hybridize it because it is single bonded and it's terminal. So it's a single bond, it's from this carbon that's  $C2sp^2$ , we already saw that, and then the chlorine is going to be terminal and so it's  $Cl3p_z$  and so that's a non-hybridized orbital. So good practice for the clicker. I think that one could help, but we're going to have more practice now.

I threw in a bunch of extra problems so that one was extra and now, let's do the one that is in the notes from last time, which is vitamin C. So I'll give you another minute if everyone has that

one down. OK. So let's look at vitamin C. So vitamin C is needed to form collagen in your body. Without enough vitamin C in your diet, you could be in trouble. So it doesn't happen too much anymore because there's vitamin supplements and all sorts of things, but often, vitamin C deficiency is associated with sailors who went out to sea and didn't have a healthy diet and they became deficient in vitamin C and got scurvy. And so then they had to figure out they had to eat oranges or other things that were rich in vitamin C. In terms of who should be concerned about vitamin C deficiency, us, primates, we don't make vitamin C, so we have to get it in our diet and also, Guinea pigs. Most other animals make it. I don't really know why-- maybe this is why Guinea pigs are called Guinea pigs, they're good for scurvy experiments because they don't make vitamin C. All right.

So let's look at this vitamin C molecule and think about what type of molecule it is and this is a quicker question. So we have to remember back more material that's going to be on exam two. Does that look like a polar or non-polar molecule and what's true about polar and non-polar molecules? All right 10 more seconds. Great. So it is polar and it, therefore, water soluble and so you know that because if there's atoms in there that have differences of electronegativity of greater than 0.4, carbon and oxygen of a difference in electronegativity of greater than 0.4, oxygen hydrogen also, electronegativity differences greater than 0.4. So we have a lot of polar bonds and they're not canceling each other out. It's not a symmetric molecule so therefore, it would be a polar molecule and water soluble. OK. Great. So you're good on your polar covalent bonds which is also going to be on exam two. All right.

So let's go back to hybridization and have a little more practice on that. So don't put your clickers away. Why don't you tell me the hybridization of carbon a labeled up here and in your notes. All right. 10 more seconds. All right. So we know what clicker questions are going to determine the winners. OK. So carbon a was  $sp^3$  hybridized. So if we look at it over here, it has bonded to four things so there's four which makes it  $sp^3$ . OK. So let's just do the rest and you can yell these out. Carbon labeled b, what kind of hybridization for carbon b?

**AUDIENCE:**  $Sp^3$ .

**CATHERINE DRENNAN:**  $Sp^3$ . Carbon c?

**DRENNAN:**

**AUDIENCE:**  $Sp^3$ .

**CATHERINE DRENNAN:** Sp 3. Again, you just want to count how many bonds you have going on or lone pairs, but carbon doesn't usually like to have lone pairs. What about carbon d?

**AUDIENCE:** sp 2.

**CATHERINE DRENNAN:** Sp 2. Right. It only has-- if we look at that one over here, I'm supposed to point to this one-- so carbon d over here, it has three atoms that it's bound to. Carbon e? sp 2 and carbon f?

**AUDIENCE:** sp 2.

**CATHERINE DRENNAN:** Sp 2. Right. So now that we did that, we can use this information when we think about the bonds that are formed between these carbons and the other atoms. So let's look at bonding now. So if we look at carbon b, two hydrogen, that's going to be a sigma bond, and you told me that carbon b was sp 3 so we write that. So describe the symmetry around the bond, name the bond, C2sp 3, H1s, we do not hybridize hydrogen.

So now, for the bond between b and a, again, a sigma bond. We already looked at the fact that carbon b is 2sp 3, carbon a was the same. Now if we look at the difference between b and c, b was C2sp 3 and then c is also the same. Remember to write the twos, remember to write the hybridization, remember to write the element, remember to write sigma for the single bond. Grading these questions on the exam is not fun. You've got to remember to have all those things in there so if you get them all in there, it makes everyone very happy. OK.

Now let's look at carbon b to the oxygen. It's also a single bond, so sigma. We know that carbon b is C2sp 3. The oxygen here is also going to be sp 3 because it has two bonded atoms and two sets of lone pairs. OK. One more clicker. All right. 10 more seconds. Great. Yup. So that is correct and if we take a look at that over here, we have carbon d, it has bonded to three things so it's sp 2 and the oxygen is bonded to two atoms and two lone pairs so it's sp 3.

We can keep going and finish up between d and c now, we have-- oops, sorry, d and c up here, we have d which is 2sp 2 bonded to three things, c has bonded to four things, it's C2sp 3. And then finally, d to e, we have two bonds. We have a sigma bond so that's between our two-- these two carbons here are hybridized orbitals and again, it's 2sp 2, 2sp 2. And it's a double bond so we have one sigma and one pi bond and the pi bond is between non-hybridized orbitals, so it's C2py, C2py or you could have used x, I don't really care about that.

All right. Good practice. I think you're getting the hang of this. Again, there will be more

practice problems on hybridization posted today to get you ready for the exam and also to figure out these bonds. Once you get the hang of this, it's really pretty trivial and good points for an exam.