PROFESSOR: OK. Couple of announcements. First of all, the Wolf Lecture is tomorrow, right here in 10-250 at 4:00. As also, there will be the weekly quiz, the mini celebration. But there's going to be, a week from today, celebration part trois. Yes, third celebration of learning. And remember, what we're shooting for is the smiley face. None of this. And those of you who are over here on the grade point distribution curve, we want to move you to the right. So I'll say a little bit more later in the week about what the coverage will be. I'll be available for office hours later, and

I thought to try to cheer you up, and remind you that studying is not something you do only the night before the exam. You have to be studying on a regular basis. All right? So from the iconic film Ghostbusters, there's this one scene.

[FILM PLAYBACK]

For a moment, pretend that I don't know anything about metallurgy, engineering, or physics, and just tell me what the hell is going on.

You never studied.

[END FILM PLAYBACK]

You don't want to be a character in that scene. So what are you going to do about it? You're going to study.

All right. Let's get down to work. Today I'm going to take one day and I'm going to talk about organic chemistry. Now, I'm not going to pretend that after one lecture, you're going to walk out of here and know much organic chemistry. I'm going to give you, if you want to know organic chemistry, you're going to have to take 512, and many if you will. But I'm going to give you enough that it will set up for the subsequent units on polymers and on biochemistry. So we'll know enough that we can move forward.

So let's get into the definition of organic chemistry. It's the chemistry of compounds, containing both carbon and hydrogen. There can be, of course, other elements present, but you've got to have carbon and hydrogen to get yourself into Orgo. And what makes carbon and hydrogen so special? We've been studying these elements through the semester. But just to be reminded.

First of all, hydrogen is peculiar because it's got the lowest atomic number. It's got the lone proton, and a single electron. And when hydrogen ionizes, it forms the hydrogen ion, or the proton. That's all that's left. And this thing's very tiny, and it can form covalent bonds. It forms covalent bonds as primary bonds, and it can form hydrogen bonds as secondary bonds. So it's peculiar for that reason. It does not form ions. You cannot imagine something is tiny as a proton occupying a lattice site in an ionic crystal.

OK. Now when it comes to carbon, carbon is also peculiar. It's smack dab in the middle of the periodic table, in the second row. Four valence electrons, which means that it's very difficult to ionize. If it wants to achieve octet stability, it's going to have to form covalent compounds. It's not going to acquire four electrons or lose four electronics. In rare exceptional circumstances it, might, but commonly, it does not. And with that intermediate average valence electron energy, it's going to be capable of all sorts of bonding. And in particular, because of its small size, it's capable, with its small size and the intermediate value of average valence electron energy-- and this is critical-- is that it's capable of forming multiple bonds. And this is key, as we'll see later when things try to polymerize and form multiple bonds.

And this is not common, for example, not the case if you look underneath carbon. Silicon, germanium, for example. Same thing. Group 4, same kind of mid-stream average valence electron energy. But they're too big. Silicon and Germanium do not form double and triple bonds.

So the other thing is that it's capable of self-linking. It can link to itself, form carbon chains, and also with nitrogen, oxygen, sulfur. So this gives it-- and let's put up phosphorus as well. Phosphorus and sulfur. So with these kinds of links-- we've already seen oxygen links. We'll see, pretty soon, nitrogen links and so on.

There's millions of organic compounds, and today I'm just going to look at a few. I'm just going to give you some taxonomy and nomenclature, and we're going to walk through this. And here's the way to keep it straight. You know all this stuff. I'm just going to organize it in your minds for you.

We've already studied the three different types of hybridization. sp, sp2, and sp3. And we've already seen that if you're going to have a double bond, you need sp2 hybridization to reserve 1p orbital, to make that pi second bond. And if you want to make triple bonds, you make sp hybridization, thereby reserving two p orbitals to make two pi bonds. So this is sigma, sigma plus pi, sigma plus pi plus pi. That's it.

Now what we're going to do is to is go back and look at this in the context of hydrocarbons. So hydrocarbons is going to be now the focus here, and the hydrocarbons are the simplest of the organic compounds. So let's look at hydrocarbons. I'm going to look at their nomenclature and their properties.

But the structure behind everything, the treatment, is the type of hybridization. So these consist only of hydrogen

and carbon. Compounds containing only hydrogen and carbon. So I'm just going to go down that chart. And we'll go through.

So the first one on the left is the alkanes We'll look at the alkanes. And the alkanes are characterized by sp3 hybridization. And so this gives the maximum number of carbon linkages. Why? Because you can only form single bonds. So if I can only form single bonds, every bond goes to a different atom. Whereas if I form double bonds, then I'm burning up some of the capability. Because if I used two bonds to get to one neighboring element, instead of two single bonds each to two different neighboring elements, I don't have as many linkages.

So the maximum number of carbon linkages-- and because they have the maximum number, we're going to borrow terminology from the last unit on solutions. What's the maximum amount of solute you can get into a solution called? It's called the saturation value. And they use that term here. These are called saturated hydrocarbons, because they've got the maximum number of linkages, bonds. And they're all sigma bonds. All of this is a consequence of the choice of sp3. I could have led you through all of this, but we're putting it up quickly here.

All right. And so they have a common formula. CNH2N plus 2. I think I've got some-- yeah, this is taking from one of the other books. I know you've got a list like this in your own book. So here they are. So this is just CNH2N plus 2, and you go down, and there's the nomenclature. The meth, for historical reasons represents n equals 1.

So here's what the hidden message is. You name them by the carbon number plus the suffix A N E. So the -anes, the alk-anes, are all these sp3. So 1 is meth, 2 is eth, 3 is pro, and 4 is but. And they all have historical derivations. But- comes from butyric acid, which is the acid that's formed when butter turns rancid, and so on.

After you get up to N equals 4, it's just your Latin. So you just take your high school Latin, pentane, hexane, heptane, and so on. And if you didn't take high school Latin, this is a great chance to learn the ordinals in Latin. So there they are. Now, I'm not going to expect you to-- if I say decane, you're supposed to slam down C10H22. But I would expect that I could say decane, and I'd give you that formula and you'd be comfortable with it.

What's the other thing to note here, while we've got this up on the chart. You can see that all of these are symmetric molecules, and they're non-polar. They're perfectly symmetric non-polar. And what happens as the molecular mass increases, the melting point increases, the boiling point increases, and the state of matter transitions from gas at room temperature to liquid at room temperature to solid at room temperature.

So here's an excellent example of polarizability in action. Because the only way one methane can bond to another methane is by induced dipole-induced dipole interaction. So the only difference between methane and icosane is that icosane is a honking big molecule, non-polar, but very polarizable. And so as the polarizability increases, you

can see the effect here. OK, so that's good. So you get to learn your Latin ordinals.

Now these are also termed straight chains. They can be termed straight chains, but I want you to look carefully at what's going on here. So here's a diagram used, the first one on the left, you've seen methane many times. There's ethane and then propane. And what do you notice here? Because of the sp3 hybridization, all of these angles are 109 degrees, including the carbon-carbon angles. And when the thing is really short, you get this zig-zaggy nature. So because of the sp3 hybridization, you see the carbon doing this. This is carbon-carbon, and then the rest, forget it. I mean, this is the backbone.

But if I get up to icosane and I have about 20 of these things, forget the zig-zag. From here, yeah, I know there's a little bit of fine structure, but for all intents and purposes, this as a chain. Straight chain. And that's where you get the terminology from. But keep in mind that there is this zig-zagging back and forth, due to the sp3 hybridization. OK? So that's the first point.

The second point is that the bonds between the carbons are not fully specified. So for example, take the carbon here on the left. The sticks to go to the 3 hydrogens here depicted in white, and then the fourth stick goes to the carbon. And then off of that carbon comes 3 more sticks, and all of these are 109 degree angles.

All of that specified, and nothing more. There's nothing saying that these 2 hydrogens in the back have to line up. These two hydrogens out front have to the line up. That's all free to rotate. And so we have that degree of freedom there. The result is that you can have something that zig-zags back and forth. So for example, this is called eclipse, because if I were to stand here, these two hydrogens are on a direct line. So the front hydrogen covers the back hydrogen. It eclipses. This hydrogen out front to your and my right is on the same line is this hydrogen to your and my left. And so if I were to look here from the end, this hydrogen eclipses, and so on.

So if I get something eclipsed, I'm going to have a perfect straight line. If I have something that's staggered, it's free to zig-zag, but it's free to zig-zag and meander, and this is what happens. So here's 2 C17H36 molecules, and the top confirmation is more or less straight. These are more or less eclipsed configurations. Whereas the lower one, you can see, there's a fair bit of stagger, and as a result, the backbone can actually loop back on itself.

But in both cases, we still call these straight chains. They're still straight chain. Because as you start at one end and go from carbon to carbon, you linearly go to the other end. It doesn't mean it's straight like straight as an arrow. It just means there's no branches.

So now I'm going to distinguish this from the other one by branching. So let's just get this up here. So we have straight chains, is one case. Straight chains. And the straight chains can either be staggered or eclipsed. And the closer you get to eclipsed, the closer you get to-- the more eclipsed, the straighter. I think you'll understand what

this means. Staggered allows it to make large loops-- I'll just put here, straight but looped. You can use that to describe some of your friends, too. All right. Now-- it's a Monday, it's a joke! Wake up! Wake up!

All right. Now we can look at another one. Branched chains. And now, this will be a major departure, whether the chain is super-straight or does a lot of looping back and forth, the branched chain is in marked contrast. Now let me give you an example of the branch chain. Best to just do it.

So here's the branch chain. I'm going to use a butane. So you know B U T is 4. So that means 4 carbons, and A means 4 carbons, sp3 hybridized. So that's it. So now I'm going to write this. Here's butane. C, C, C, C. And I'm going to write this as just a straight chain, instead of doing this. I'm going to contrast. I'm going to do 1, 2, 3, 4. 1, 2, 3, 4. 1, 2, 3, 4. 1, 2, 3, 4. OK? So there we are.

So this is, instead of writing all this stuff, this is butane, and this is butane. But now here's the thing. Four stick rule. Whenever you're doing Orgo, four stick rule. That means, four sticks out of every carbon. So here's the carbon-carbon bond, so I need three more sticks. 1, 2, 3. And I know these are sp3 hybridized, but just, in compressed notation, you can do this. And I'm not even going to write all the hydrogens. Here I'm going to write the hydrogens because it's our first time through, and it's Monday, and everybody needs a little bit of TLC.

But now the TLC is gone. This is it, all right? I'm not even going to put the H's on the end. The stick coming off of carbon means there's a hydrogen here. All right? So how about this one? 1, 2, 3, 4, 1, 2, 3, 4, 1, 2, 3, 4. And now let's count the hydrogens. 1, 2, 3, 4, 5, 6, 7, 8, 9, 10. $C_4H_{10}$. Butane. All right.

Now there's another way to write this. I could go like so. 1, 2, 3. And I could put a carbon up here. This is a branch, you see? It's not in a straight line. And I'm not talking about eclipsed. Because if I start at 1n and go down the backbone, I get to the other end, I only hit three. I never got to this one. But let's see. You know, maybe this is different.

So four stick rule. 1, 2, 3, 4. 1, 2, 3, 4. 1, 2, 3, 4. 1, 2, 3, 4. Now let's count the hydrogens. 1, 2, 3, 4, 4, 6, 7, 8, 9, 10. This is $C_4H_{10}$. It's a different molecule, you can see. With what you've already learned in 3091, this has a different dipole moment, doesn't it? It's got a different dipole moment. It's going to have a different polarizability. But it's a butane, it's a $C_4H_{10}$.

So this is called butane, right? And this one, over here, historically, was called isobutane. Because it's an isomer. It's got the same chemical composition, but it's got a different makeup. It's got a different structural makeup. It's got a different-- I'm going to use the material science-y word. Different structure. It's got a different molecular structure.

But the chemists going back 50 years ago didn't use these terms. They called it different constituents. So I'm

going to use that as well. Different molecular structure or different constituents. And what do we mean by constituents? Now comes the color chalk.

So we're going to do, is I want to circle the makeup of this. So at this end, I've got-- this is a CH3 here, correct? There's a CH3 over here. And then in between, I've got CH2 and a CH2. Now let's go over here and diagram this one. The isobutane, I've got a CH3 at the end, and a CH3 at the opposite end. So far, same as the butane.

But now look. There's a third CH3 here, and then there's this thing, which is a CH. You could argue it's a CH2 that got methylated, but, you know. So I'm going to put a CH2 here that's been methylated. And I'll tell you what that means. OK? So it's really a CH2 group that got methylated. So it's called isobutane, but you can see, it's got different-- all I'm doing here is making the justification for the different constitution, different dielectric constants, and so on and so forth. So this is where the chemists must come in. They would say, different constituents. And so these are isomers that have the same chemical composition, and are different on the basis of their constitution. So they call these constitutional isomers. This is a constitutional isomer. Isobutane is a constitutional isomer.

Now the last piece is nomenclature. They don't use this today. Because isobutane was a term that came out of chemical industry. Because it looks like butane, has got the same molecular weight, but it's very different.

Today we use the IUPAC notation. Remember this? International Union of Pure and Applied Chemistry. They're the ones, when one of you discovers and stabilizes element 113, And you want to name it after dadadadada, you have to get past the IUPAC committee. OK? So their nomenclature committee would say, this is an alkane, but it's only three carbon units long. So it is not a butane. It is a propane. It is a type of propane. But it's not simple propane, because there's a methyl hanging off the side. So in point of fact, this, in IUPAC nomenclature, is a methyl propane.

And furthermore, if you want to get really pedantic, and IUPAC loves to get pedantic, is what they'll do, is they'll number these, starting from left to right. So this is the number 1 carbon, this is the number 2 carbon, and this is number 3 carbon. And the methyl group is pendent on the number 2 carbon. So if we want to get really, really, really pedantic, and who wouldn't, on a Monday morning, what we would do, is we would call this 2-methylpropane. So it's a propane, 3 carbon units, there's a methyl group hanging off of number 2. Now, I don't expect you to be able to do these on sight, but I would like you to be appreciative of what this is. So I'll give it to you on an exam.

All right. So we've got the isomers. And I think I've got some-- I think I've got this. No. OK, this is good.

Oh. There's one other thing I want to talk about with alkanes. There's another thing that we need to know, and that's about the radicals. The radicals that form. The radical is a species with one or more unpaired electrons.

Right? This is broken bond. One pair electrons. That's equivalent to broken bond. OK.

So this is highly reactive. Because this unpaired electron is going to go look for a mate. So if you take a look at CH4-- now I'm going to put the hydrogens in. So this is methane, CH4.

I'm going to break this bond in the lower right corner, and just have an unpaired electron. And this is called the methyl radical. Very reactive. So you can see that this methyl radical could go up and stick on to the fourth strut of that number 2 carbon, and thereby form the carbon-carbon bond. So that's why they call this methylpropane. So the methyl radical.

And you can write it this way, H3C with dot indicating unpaired electrons. But I want you to be literate with the way chemists will write this in various compact notations. So you might even see it written this way. Now, that doesn't mean that the unpaired electron is on a hydrogen. This is just a way of saying, you know, CH4 is methane, CH3 is methyl, and who cares. There's only one possible meaning to this, so it's all good. All right?

Let's do two unpaired electrons. So I'm going to take two unpaired electrons. Again, this is out of the alkane, so I'll put a hydrogen above and below. So that would be like this one here. You see the CH2? So it's got carbon-carbon bonds on either side, but in point of fact, this is part of a zig-zag. So I'm going to break bonds here and here. All right?

So now I've got the possibility of putting a bond on either side. And this one here is called methylene. This is methylene. This is the methylene radical, and it's got two unpaired electrons. So we could write that as H2C with 2 unpaired electrons.

And, you know, these are electrons. They're not just anywhere. They're sitting in an orbital, and that orbital is as at 109 degrees from the other bond. So it has to be sitting there in reserve, in waiting. So if I bring another carbon here, I get the carbon-carbon bond, and I start forming the zig-zag.

And there's one more, a common one that we'll need to know. And that's from ethane, which is C2H6, right? Let's start with ethane, C2H6, and I'm going to make a radical out of that. It will be C2H5, like this. To which I can add a hydroxyl, and now I'll make ethyl alcohol, and then it's party time. But so this is how you start. All right? This is the ethyl radical. Starting from methane.

And that's pretty much it. That's about all I want you to really take away with the alkanes. C4H10. Yeah, there you go. So there you can see, to the left, the zig-zagging of the linear and the butane, and then the methyl propane. That's good.

OK. Now let's go to sp2. Now we're going to look at the alkenes. So alkenes represent-- the alkenes-- this is sp2

hybridization. And they contain-- this means there's the possibility of a carbon-carbon double bond. And they're also known as unsaturated. Because they don't optimize. The fact that you spent two of your bonds going between two atoms instead of taking one bond to one carbon and another bond somewhere else, means that you're undermaximized. So unsaturated hydrocarbons. OK? And the formula for these is CNH2N, where obviously N is greater than 1. We can't form a carbon-carton double bond with only one carbon.

So you know, the prototypical one we looked at was this one here, just to C2. That's the first one. So it's going to be C2H4. So carbon-carbon double bond. There's a sigma bond here, and a pi bond. And they lie in a plane. The molecule is planar. 120 degrees here. So we'll put the hydrogens on it-- you know already, I don't have to put the hydrogens there. That's good enough.

All right. So what's this one called? This is called ethylene, historically. Ethylene. But the IUPAC notation is similar to this one. It's carbon number plus -ene, E N E. So strictly speaking, this is what everybody knows. But IUPAC insists-- and I don't know. As you publish, some editors are really fastidious, and they'll force you to change if you use this kind of terminology. So you're going to have to call it ethene. Or in the UK they call it ethene. But everybody knows it is ethylene. OK. So there it is.

And you only need 1 carbon-carbon bond. So you can have long chains where the position isn't fixed. So let's look at a couple of examples. So let's look at butene. OK, let me just write that down. That's good. The position of carbon-carbon is not fixed. It just has to be somewhere in the molecule, which consists only of hydrogen and carbon

So here's a couple of examples. I'll start with this one. This is going to be-- so this is 1, 2, 3, 4, so that's a but-something. And it's got a double bond, so that's a butene. It's all hydrocarbon. So that's butene here. But I could also-- well, let's finish this. How many? This is 1, 2, 3, 4. See, I even tried to indicate the 120 degrees here. This is going to be 1, 2, 3, 4. 1, 2, 3, 4, 1, 2, 3, 4. All right. So there's the butene.

But there's another place I could put the carbon-carbon double bond. I could do it this way. Make a carbon-carbon single bond, move the carbon-carbon double bond over to the number 2 position. this is 1, 2, 3, 4. So this is going to be 1-butene, and this one here is going to be 1, 2. So we'll make this 2-butene. And let's keep counting here. 1, 2, 3, 4, 1, 2, 3, 4, 1, 2, 3, 4, 1, 2, 3, 4. And I've proven to you, I hope, by now that I can count to four.

So this is 2-butene, right? 2-butene. And if you go through this using the same kind of diagrammatic analysis as over here, you're going to come to the conclusion that even though these both have the same composition, which is going to be C4H8, they're both C4H8, but as you diagram them, you'll find that they have different constituents. So these, in fact, are constitutional isomers of one another. Constitutional isomers. Yeah, that's pretty good.

There's even more. It gets even better. I want to focus on the bottom one, the 2-butene, and I want to write it out with a little bit more detail. So let's start with that carbon-carbon double bond. And we know the carbon-carbon double bond requires planar structures at 120 degrees from one another. So let's look at that one. And what we've got here, is on the carbon, I've got on the one side, I've got a hydrogen strut, right? And on the other side, I've got a methyl. I've put a methyl here, like that. OK?

Or actually, let's get-- OK, I'll do that. I was-- what I wanted to do, to get really-- OK, watch this. All right. So what I'm going to do on this side, this side, same thing. The carbon has a hydrogen strut down here, and on the other side, you've got the methyl. OK?

So that looks good. But there's another way to do this, and still keep the carbon-carbon bond in the center. So I'm going to do that over here. So 1, 2, 1, 2. And in this instance, I'll put the same thing I have on the left side, H3C. So so far, so good. I've got the same molecule. But now what I'm going to do, just for grins and chuckles, I'm going to put the hydrogen in the upper position, and the methyl in the lower position.

Now, these are not constitutional isomers. They have the same constitution, right? Carbon-carbon double bond, two methyl groups, two hydrogens. Carbon-carbon double bond, two methyl groups, two hydrogens. So they're constitutionally identical, yet these have very different properties. Can you see?

I mean, the dipole moment is going to be different. Look. Here's the charge, all up on one side, and here the charge is more symmetrically straddled. Which means, they're going to have different melting points. They're going to have different boiling points. And they have the same chemical formula, and they have the same chemical constitution. So they're different kinds. There's some kind of isomers-- what's the difference here?

It's not the constitution, it's the way things are arranged in space! This is great material science. Spatial orientation is everything. And what do you call it when you're listening to some loudspeakers and a really good audio system, and you can close your eyes, and the violins are over here, and the cello is here, and you can hear the percussion way in the back. What is that called?

Stereophonic sound. Stereo phonic, meaning you can render spacial resolution. So these are called stereoisomers. One's a stereoisomer of the other. So they're spacially distinguishable. So they have-- for example, I went and looked this up.

Oh. Let's give some names to this. In this case, all the methyls are on the same side, and the hydrogens are on the same side. So I can a dividing plane here, if you like, parallel to the floor. And I'll just cut right through the center of the carbons, parallel to the floor. And so in one instance, all the methyls are on the same side, and here the methyls are on opposite sides of the double bond.

So when they're on the same side, this is called cis-butene. And let's be super pedantic. It was a 2-butene, so now it's a cis-2-butene. This is terrific. And this is a 2-butene, but it's on opposite side. So this is called a trans-2-butene. Cis-2-butene and trans-2-butene.

And I looked up-- so let's see, which one do you think has tighter, which one's going to be more tightly packed? Which one's going to have the greater density? Which one's going to pack better? Cis or trans? The density here is 0.627, and the density here is 0.61. OK?

Now, which one's going to have the higher boiling point? Which one's going to have the higher boiling point? Gee, that should be a no-brainer. Once you know which one's got the higher density. Why does it have the higher density? Because it's more nearest neighbors is more tightly packed. So I would bet on the one with the higher density. And this one boils at, boiling point is 3.7 degrees Celsius, and this one, the boiling point is-- hang on, I've got this backwards. This is 3.7. Boiling point is 1 degree Celsius. So those are the data as they come up.

So stereoisomer, spacially distinguishable. By the way, we can form multiple bonds. We can form multiple double bonds. So they're -enes as well. So if we have two double bonds in the molecule, so we'll call that a diene. And we can have-- you can keep going with the Latin, but common ones you might find is the triene. And this will come back later when we look at polymers.

So if we have two bonds, we will have-- You know, some of that hard luggage that you're seeing that's coming back, they got the roller boards, but instead of the fabric, now they're coming back with the really hard polymer? That's ABS, and the b is butadiene. So let's look at a-- here's a pentadiene. I'll give you a pentadiene. So that's going to be 1, 2, 3, 4, 5. It's five carbons, for starters. And I'm going to put the double bonds on the 1 and the 3 carbon. It'll be a penta-, it's five carbons. And it's an -ene, because there's going to be some double bonds here. And it's a di-, because I'm going to put double bonds at the 1 and the 3. So 1 gets a double bond, and 3 gets a double bond. So this is a pentadiene. OK?

And then the last thing is-- pardon me. Last thing is radicals. So here's the common radical. We'll start with ethylene, and I'm going to make an ethylene radical here. I'll break off the hydrogen at this point, and I can write that as CH2CH. And we can't call this ethyl radical, because ethyl was already used for C2H5. So we need to indicate that there's a double bond, and so this radical is known as vinyl. This is the vinyl radical.

And so you could react this with chlorine, for example, in which case you'd make vinyl chloride, and then you'd polymerize that, and you make PVC. And we're going to see that in the next coming days. Or you could you can put alcohol on there. I've already told you that. I don't need to remind you twice about that, I bet.

OK. So now-- oh, there's one other one. There's one other one that will come up if you take a lot of biochemistry,

and that's the one that comes off of propene. So if you start with propene, propene's going to look like this. 1, 2, 3. So there's propene. 3. There's a saying double bond here, and I'm going to put 1, 2, 3 in here. 1, 2, 3, 1, 2, 3. 1, 2, 3, 4. OK. So this is the radical that comes from propene. So this one is derived from ethylene, and this one here is derived from propylene.

So again, as in this case, we couldn't call this ethyl. We can't call this propyl, because that's the one that you get from the alkane. So instead, this one is called allyl. A L L Y L. And if you take 7012 at some point, Professor Weinberg will refer to these pendant groups. He likes to pronounce this all-YL. You'll hear him talk about, and the all-YLS are over here and here. That's what we're talking about.

OK, good. So this is $C_3H_7$ dot. Good.

All right. Moving right along. Now I want to turn to the alkynes. That's the sp hybridization. sp. And that means there's going to be at least one carbon-carbon triple bond. One carbon-carbon triple bond. So the formula there is $C_NH_{2N}$ minus 2, and obviously N has to be greater than 1, otherwise you can't form a bond.

And there's very little that you need to remember about this one. The dominant one that you'll come upon is $C_2H_2$. So according IUPAC nomenclature, the two carbons means it has to be F, and these are all alkynes, so this should be ethyne.

And if you ask for ethyne, nobody knows what you're talking about, although it's very correct. This is known as acetylene, which is a fuel gas for welding, cutting, torches and so on. You've got two lines, you've got a big tank of acetylene, you've got a big tank of oxygen. And this is used because there is enormous energy stored in this carbon-carbon double bond. And so when you combust this with the stoichiometric amount of oxygen, you can cut through steel, you can use this to make very elaborate glassware with pure silica, making fused quartz glassware. Very, very high energy contained in here.

And just to give one more example, let's look at something that's got four carbons in it. So it's going to have four carbons and a triple bond. So if it's 4, it's going to have the B U T but-, and it's going to need an -yne. So this will be a butyne. And depending on where we want to put the triple bond, I don't know, if you want to put a triple bond here, there's 1, 2, 3, 4. So this we could call 2-butyne, and it's $C_4H_6$, If you go through the stoichiometry.

Let's try it, just for practice. So this carbon on the end is 1, 2, 3, 4. Look at the second carbon. It's got one strut bonding to the neighboring carbon, and three struts bonding to the neighboring carbon on the right. It's saturated now. Nothing comes of without carbon. Look at this one. 1, 2, 3, 4. Nothing comes off of this one. And then finally 1, 2, 3. So there's three hydrogens here, there's three hydrogens here, there's your sixth hydrogen. $C_4H_6$. OK. That's good.

All right. There's one other set that I want to look at, and these are called the arynes. These are aromatic hydrocarbons. Some people refer to them as arynes.

And the most notable one is C6H6, which is benzine. And benzine's molecular weight was determined long ago, in the early 1800s. And it was Kekule who, in 1865, dared to suggest that the chemical structure of benzine was a ring structure. So Kekule suggested a carbon ring. Up until this time, people didn't have the common understanding that carbons could link to themselves, let alone form a ring.

And so now, if we go through this, we will put, from each of these junctions there's the carbon, so we have to have four struts. We need hydrogen, so let's put hydrogens 1, 2, 3, 4, 5, 6, So there's the 6 hydrogens. And now, each carbon needs 4 struts. So 1, 2, 3, and I'll put a double bond. That will make 4. 1, 2, 3, 4, 1, 2, 3, 4, 1, 2, 3, 4, and now this one here. So you have an alternating structure of double bond, single bond, double bond, single bond on the carbons. OK?

So that's the way things stood until the twentieth century, when on the basis of new data, it was discovered that all carbon-carbon bonds-- and I don't want to make this a single bond, I'll just say all carbon-carbon bonds in benzine, C6H6, were measured to be the same length.

Well, that's a problem, isn't it? That's a problem, because we would expect that, you know, we know that carbon-carbon single bond is about 1.47 angstroms, and the carbon-carbon double bond has to be shorter, because it's double-bond, it's tighter, so it pulls things in. So that's 1.33 angstroms. And the measurement here was found to be 1.39 angstroms. And 1.39 angstroms is midway between 1.47 and 1.33.

So people were frustrated with this new information. And it was Linus Pauling who made sense of this for us. Pauling, who in 1931 came along and said, what we have a mix? What if we have a mix of two structures? And he said, suppose you had something that looked like so. 1, 1, 1. So here's the set of double bonds off of the ring. And what if there is a resonance with an opposite hybrid, where this double bond, instead of being here, goes over to here? So now we have this?

And he said, if the system resonated between the structure on the left and the structure on the right, the time average value of the carbon-carbon bond length would be around 1.39 angstroms. So this is a resonant hybrid structure, to account for the mix of single and double bonds, which gives rise to the current representation of benzine as simply this. So we're not saying this is single or double. It's a resonant between one and two. OK.

The other thing is, if all of these bonds are the same length, and that means that between the 12:00 position here and the 2:00 position, there is pi bonding, and then along this double bond, there's pi bonding, and then along this double bond, there's pi bonding. But if all of the bonds are the same length, I can't say that the pi bonds are

confined to where the double bonds are indicated, because resonance indicates that it could be between adjacents.

So that means that the electrons, in fact, can go all around. So resonance, then, means that the pi electrons are delocalized. And now here's some better drawings. It's hard to draw that. So here they are. OK. So this is a nicer drawing of benzine. So you can see, when all of these bonds are resonant, it doesn't matter whether I choose this front one as the double bond, or the one next to it. They're all the same length, and so the electrons can go all around. And you can imagine, if I did this with graphite, and I kept going, well, if the electron can go here, then it means it can go here, which means it can go here, which means it can go everywhere. Which is why graphite is an electronic conductor, because of delocalized pi bonds.

OK. This is 1-3-butadiene. Double bond, single bond, double bond. Again, all the same length, and you can hybridize, and-- This is how you start making electronically conducted polymers. You just go double bond, single bond. Double bond, single bond. Smear. That's it.

OK. Let's jump over that. Let's talk about Kekule So he studied architecture, switched to chemistry. And then he got a job in London, and he used to fall asleep on the bus to his apartment. And one day he was sleeping on the bus. He woke up, and he'd been dreaming about carbon forming chains in 1855. And he proposed carbon chains. And then he got a job as a professor in Belgium, at the University of Ghent. And one night he fell asleep at the fireplace, and he dreamt of a benzine molecule as a snake biting its tail while spinning. And that's where he got the idea of the ring molecule. All he knew was that this thing had a formula, C6H6. And for this, he's dubbed the founder of structural chemistry.

So what's this formula for success? Well, he moved into chemistry from another field. But to dream, you've got to get some sleep. And I'm probably talking to one of the most sleep-deprived populations on this campus. So I urge you to get some sleep, and then maybe you'll be able to dream.

Hey, don't make noise. We've got two minutes left. nobody's going anywhere.

All right. So the next thing is, I want to talk a little bit about octane ratings and automobiles, and how we go from straight chains to chains of different lengths. If you're looking at combustion, the idea is, you admit gasoline, which vaporizes, and then the piston comes up, and under high compression, the spark plug fires. When the spark plug fires, this causes an explosion which then pushes the piston down, and then you've got the camshaft that takes the vertical motion and converts it to rotary motion.

Now, if you're running an automobile after a certain period of time, the combustion chamber becomes so hot that just admitting the fuel, having it vaporized, before the piston has risen fully to get maximum compression, this

thing could just explode. And when it explodes, it'll send the piston down, and it's out of sequence with the other pistons. And that's when you get the knocking. Sometimes you're driving up hill, you hear this [GRINDING SOUND]. That's the knocking going on. You need to get a tune-up, or switch to higher octane gasoline. So what's going on in here? What you're doing, is you're changing the chemical composition, the constitutional isomerization. And you do that in the process by which you synthesize the gasoline in the first place.

So the figure of merit is called octane number, which was first instituted here in the United States. And they started with 2-2-4-trimethyl-pentane. It's an octane, but it's an octane that's 2-2-4. Three methyl branches. So that means there's only five left. So it's a pentane with three methyl branches. And that's called 100. And then is 0 is heptane. If you put heptane into the combustion chamber, you'll knock yourself silly. And the engine will just shake and shake and shake. So gasolines have to be somewhere on that interval.

So there's the-- you know, if you tested and you had something that was trimetylpentane, 90%, versus 10% heptane, you'd call it a 90% octane. And you can have octane numbers higher than 100, by the way. It just depends on what the repression is.

And oddly enough, when you when you increase the octane, you make the fuel more difficult to burn. Because you want it to burn only on demand. Fuel that just burns whenever it wants leads to random events in the engine, and that's not good, if you want to get good thrust. So oddly enough, high octane gasoline requires more ignition than low octane. And the additives are tetraethyl lead or ethyl alcohol. And from that, you know now, we get E10, which is gasohol, 10% alcohol. E85 is 85% ethyl alcohol.

And all of this you get by synthesis. You know, playing with the catalysts, temperatures, and so on, to get the right mix to get good fuel, and get good emissions. And you can decide amongst yourselves whether going to E85 high levels of ethanol, ethanol derived from corn, is that smart, is that food for fuel? You know, what's the sensible use of agriculture, algae, switchgrass? All of that stuff. I mean, it's all loaded into here. All loaded into here.

So yeah. With that, I think we'll adjourn. I'll see you on Wednesday.