PROFESSOR: OK. All right, let us take ten more seconds.

All right, 92%. I like it.

That's great, I told you there'd be a formal charge question, and there was, and you learned it. Awesome, that's what I like. All right, that was carbon had four, no lone pairs. So, 4 minus 0 minus 1/2 of eight bonding electrons, which is four. So, 4 minus 4 is 0. OK, so today we're having a clicker competition. And problem set 4 will be posted later today. And today's goal is for all recitations, but recitation 12 to try to unseat recitation 12. And the goal of recitation 12 is, of course, to win an unprecedented third week in a row. So, that's the goal for everyone today. Now, remember recitations that win a multiple weeks will be in the playoffs for the clicker competition at the end of the semester for a specially designed t-shirt.

OK, more on clicker competitions actually later today. So, if we can get settled in. I know it's exciting. And you are allowed to talk to your recitations, other recitation members during the clicker questions, that is allowed. But I'll need to cover a little material first.

So today, we're going to talk about shapes of molecules, and I brought some molecules with me today to help me out. And we're going to be talking about VSEPR theory. So, why is shapes of molecules important? So shape, which we can also call geometry. So, you're going to ask about the geometry of the molecule, you are asking what is it's shape. It's particularly important in chemistry because shape can dictate properties of a particular molecule. It can tell you about-- or dictate melting points, or boiling points, or reactivity. And I'm a biological chemist, so I care a lot about shapes of molecules because in biology shape is really important. So, you have enzymes in your body catalyzing reactions. And for those enzymes to work, they are specially designed to react with one particular kind of molecule and not any molecule in the cell. So, they're designed to recognize the shape of that molecule.

So, biochemistry really works by shape recognition, so shape is very important. So, there's a lot of ways to get information about shape. But there's one very simple theory that does exceedingly well in predicting the shapes of small molecules. And that is called the Valence Shell Electron Pair Repulsion theory, which is known as VSEPR. And is also known as the V-S-E-P-R theory. I will call it VSEPR because it is really hard to say V-S-E-P-R theory. So, VSEPR is the topic today and this is based on Lewis structures. So, this is highly exciting because you just finished a problem set that had you draw lots of Lewis structures. And now on the next problem set, you can draw more Lewis structures and then tell us about his shape. So, that's very exciting. So, you're going to retain all of the knowledge that you've gained in the last problem set and continue on with that problems.

All right, so this is a very simple theory. And it's based on the idea that valence shell-- valence electron pairs repel each other. Electrons repel each other. They're negatively charged, they repel each other. That is the heart of this theory, very simple. And as many of you may have gathered, I love it one very simple theories explain a lot of stuff. I enjoy that. Gets it, it doesn't get it right 100% of the time, I'm OK with that, I'm good with about 90%. So, this works really pretty well. So, again we're talking about the geometry around a central atom. And the idea is that the atoms are lone pairs around that central atom are going to move in such a way, attain such a shape, such that the repulsion is minimized. So, this is all about minimizing repulsion, minimizing stress. Again, I'm a big fan of minimizing stress, so I also like VSEPR for that reason as well.

AUDIENCE: [INAUDIBLE]

PROFESSOR: Lewis structures? What causes stress? Problem sets. Oh, but they build character. All right, VSEPR Very simple nomenclature, again I like simple nomenclature. A is the central atom. Maybe should be C, but atom is there, so central atom is A. X is the bonding atom, and E is the lone pair of electrons. So, E lone pair electrons. X is whatever bonding atom, and A, A is in the center. A is the central atom. One more term you need to know for VSEPR and that's steric number, . And that's used to predict geometries. So, what is steric number equal?

Steric number is the number of atoms bonded to the central atom, plus the number of lone pair electrons. And you count one pair as one. And so, you want to note, when considering VSEPR double triple bonds are all-- and single bonds all the same. So, you don't have to worry about double triple bonds right now. They're all counts the same. It's only the number of bonded atoms and the number of loan pairs. So, let's look at some examples of this. So, we have our central atom A, bonded to two bonding atoms X, with one lone pair of electrons. So, if you were asked what is the formula for this, the VSEPR formula, it would be AX2E, as shown there. And then you might be asked to steric number. And the steric number in this case would be what? Three.

We have two bonding atoms. One lone pair of electrons. Again, the lone pair just counts as one. So, I could also draw this structure with a double bond to one of the Xs. And this would have the equation AX2E, exactly the same formula. Because we don't care about the double bond in the formula, only the number of bonding atoms. Only the number of lone pairs. And this would have an SN number of what?

Three, right. Because it only matters about the lone pairs and the atoms. That's all we care about here. We don't care about the double bonds, triple bonds, whatever, we'll care about them later. But for right now for geometry, we're only caring about bonded atoms and lone pairs.

Some of you will be very happy after doing the last problem set to know, that you can apply VSEPR theory to all the resonance structures that you may come up with when you're doing Lewis structures. So, if a molecule has one or more resonance structures that's OK. VSEPR can be applied to any one of them.

Also, if there's more than one central atom, you need to consider those atoms separately. So, in this case, you would be asked about the geometry around the carbon and the geometry around the oxygen. But a lot of the examples we have today just have one central atom.

All right, so that's just a little introduction to VSEPR Now, there's two cases that we're going to consider today, one are molecules without lone pairs and one are molecules with lone pairs. Without lone pairs is a little bit easier than with lone pairs. So, let's start there. And we have a nice table for you. I'll tell you that these lecture notes are of high value. I've had people who've taken this course want to come back, it's a nice summary of all these shapes. So, you want to keep this in a nice, secure location. A highly desirable notes.

OK, so let's look at a formula type. The simplest we have AX2 here. So, that has a SN number of 2, because it has two bonded atoms. And it has this molecular shape. And I brought an example here. And this is, of course, a linear molecule and therefore the angle-- and here we're talking about the angle around the central atom. So, from this black atom to that black atom, we'd have 180. So, that's the simplest linear molecule.

So, it's got a little more complicated and add three bonding atoms. So, here we have a case of AX3. It has a SN number of 3. There are three things bonded to the central atom. And its shape is trigonal planar. Trigonal should be remembered, it looks like a triangle. And you should also remember that it's planar. You can hold it this way and see all of the atoms are in one plane. And I'm making some of this point because later on the exam, when you're asked to name geometries of things. People come up with all sorts of crazy things. So, if I spend a little time doing my demonstration here of the shapes of molecules, it will pay off later in the exam. It'll sear in your brain. It will be hard to forget. Trigonal planar. What are the angles?

120.

Right. OK, let's move on. We'll have four atoms. SN number of 4, AX4. So, this is our tetrahedral geometry. And I'm going to just note, as it says down here, that when you have a thick arrow coming out at you, like this bond and a thin one going back, that means the one coming out is coming straight out towards you. The dashed line is going back into the screen. And the two that are not thickened or dashed are in the plane. So, if I hold it like this, we have two atoms in the plane, one coming out, one going back. And so, that's how-- if you see that drawn that way, and you will, you can think about it in three dimensions.

All right, so what's the angle of tetrahedral? 109.5. So, a lot of people have already familiar with this is great, if you're not, you will learn it quickly when you do the problems.

All right, we'll keep going. We have five. And you move over here. So, a SN number of 5. We have AX5. And here is our shape. So, this is trigonal bipyramidal. So, if you think about the shape of your trigonal planar, you see along here, we have our trigonal planar. So, we have our trigonal shape, but we also now have an atom above and an atom below. And that forms a kind of a pyramid on top and a pyramid on the bottom. So, it's trigonal bipyramidal. And there are two sets of angles now. What are the angles in the equatorial? 120.

I love when people yell things out, it's awesome. Angle from the axial to the equatorial? 90. Awesome. All right. And one more basic shape group, and that's this one here. We have six atoms bonded to the central atom, AX6. And we have this shape. So, now I'm holding it, two axial atoms. Two atoms coming out towards you, two going back. Octahedral geometry. And what are all the angles here? 90, awesome.

All right.

So, let's look at some examples. Hopefully, everyone got the 90 and can write it down. If not, I'm sure someone will yell it out again for you. All right, so let's look at some examples. We were talking about CO2, and you calculated the formal charge on it for me. So, we have AX2. SN number of 2. This is linear and our angle is 180. And now we can explain why this is non-polar.

We learned before that we had polar bonds where we had difference of electronegativity between the carbon and the oxygen, which would make a polar bonds. You have two polar bonds, so why is this not a polar molecule? And it's not a polar molecule because this oxygen is pulling this way, this one's pulling that way, and that makes it non-planar because there's no net dipole. So, shape is important for this.

All right, let's look at the next example. We've seen a lot of these examples before. We're talking about Lewis structure. What was boron an example of?

Incomplete octet. That's right. So, here's an example, it has three things bound to it. And it has the shape of trigonal planar. And the middle atom is incomplete octet. But it's OK. Boron and which other one are OK being incomplete octets? Aluminum, right. So, the angles are 120 here.

Now, we move on. We have this molecule CH4. Can someone tell me what this is? Methane. So, we have AX4, SN number of 4. And we have our tetrahedral shape and our angles of? 109.5. Right. So, methane we should all care about. Greenhouse gas, but also some people believe will be the salivation to our biofuels and energy problem. We will see if that's true or not.

OK, another example over here. So we have phosphorus in the middle and five chlorines. So, what is this an example of, in terms of Lewis structures? Yep, so this is trigonal bipyramidal or bipyramidal. I think both are right, I don't know. And so we have our 120 and our 90. And this was an example we had in Lewis structures or something very similar of an expanded octet. So, we have more. We have five bonds around the phosphorus and that's OK because it's N equals three or greater. So, we have an example again of an expanded octet. And if we keep going, we'll have another expanded octet. So, here we have AX6. So, we have six bonds around the sulfur. Sulfur is OK with that. And what is the geometry here? Octahedral angles? 90. Yep, all right, so we have several examples here of things with expanded octets and also deficient octets.

All right, so this is pretty straightforward. People get a lot of points on exams, except when they come up with all sorts of weird kind of shapes that don't exist. But most of the time you can learn this and it's great. With lone pairs it's a little more complicated, but also much more fun.

All right, so let's talk now about what happens when we have lone pairs.

So, electrons in bonds are hanging out in their bond and they're not really doing much, but being in their bond. So, they have less spatial distribution than loan pairs. Meaning that, electrons in bonds take up less space. And electrons in lone pairs, they can be anywhere. They're not restricted to their bond, so they take up more space and therefore, cause more repulsion.

So, the whole idea of when you're talking about VSEPR with lone pairs is that you're thinking about electron pair repulsion. And if you have a lone pair, that's going to give you more repulsion than bonded electrons because lone pair electrons can take up more space. So, it's a very simple idea, but it actually works to explain a lot of stuff and again, the geometry re-arranges to minimize that repulsion.

So, when we're talking about repulsive forces then, we go from the most repulsion, lone pair, lone pair. That's like two messy roommates living together, that's very repulsive situation. Lone pair, bonding pair. And then bonding pair, bonding pair is the least repulsive. To neat roommates that usually works out quite well. So, if we keep this in mind, we can now predict shapes of molecules based on this repulsion. So, we can rationalize shapes based on VSEPR theory.

So, now we can think about AX4E has a seesaw shape. Which of these two shapes is seesaw? So, we have one case-- I'll make this one-- where we have the lone pair in an axial position. And we have another case where we have the lone pair in an equatorial position. So, when you have it in the axial position, here, you have three bonding pairs of bonding electrons pretty close. They're 90 degrees away from each other. So, there are three sets of bonding electrons that will be repelled pretty strongly. Now, clicker question. Think about what's going to be true here. How many things with an equatorial lone pair will be repelled strongly?

Very repulsive, those lone pairs.

OK. So, that was the answer we're going for two. Because there are two sets of bonding electrons here that are 90 degrees away. There are two that are 120, but 120 is bigger than 90. So, there are two that are close whereas, with this geometry there three sets of bonding electrons that are 90 degrees away. So, it turns out that the equatorial lone pair is more favorable. It has a little bit more room in the equatorial area to spread out. And so, this is the shape that we find. We never see this shape, does not exist. We'll just take that away, it doesn't exist. And this is called seesaw. And let me demonstrate to you why that is the case.

May I'll do it this way.

So, how many of you had seesaws in a playground? Quite a number. Seesaws are not considered that safe. You know, the heavy kid sits on it and keeps you up in the air for days or whatever, until your mom comes and gets you from the playground. Or gets off suddenly, and the seesaw flips, and you go flying in the air. But studies show actually, that dangerous playground equipment builds neural networks and is good for cognitive development. So, I think we could do a survey here and see if there's a correlation between how many of you played with seesaws when you were a kid. And whether you ended up as an MIT student or not. That could be interesting, but now none of you will ever forget that this shape is seesaw, right? You will remember this forever? OK.

All right. So, this idea also is responsible for a T-shaped molecule. So, we'll add another one. And I think this is a little harder to rationalize why it wouldn't go in a different place, but it does. This is how it goes. Both the lone pairs are an equatorial positions, and we get something that looks like a T. So, if you have

AX3E2 SN 5 number, it's a T-shaped molecule.

We can also think about our core octahedral geometry. And if you have two lone pairs, so here's our octahedral with six bonding atoms. And if you have two lone pairs and four bonding atoms. If you have AX4E2, it has this shape. The lone pairs go on opposite sides. And now they have their repelling the bonding electrons here, but they're far away from each other, which is favorable. And this is called square planar. So, it's square and it's planar. Square planar.

OK, the two lone pairs are far apart on the opposite sides of the bond. So, in addition to predicting these shapes, which work pretty well, we can also think about the geometries. And when you get lone pairs, you find deviations from your standard geometries. So, let's look at some examples. So, with molecules with a lone pairs, such as NH3, the angles tend to be smaller. So, we saw methane before, it was 109.5. You yelled that nicely out for me. These are both SN 4. This is AX4. This is AX3E, but it's still an SN 4 system.

So, it looks like this, we have-- here we have methane, here we have NH3. And now, we want to think about which would be more repulsive, the bonded electrons or the lone pair electrons in terms of the geometries around. So, if we have hydrogen carbon hydrogen angle of 109.5 here, then here with nitrogen we're thinking about hydrogen nitrogen hydrogen. And this lone pair is pushing on those bonded electrons, and it's taken up a lot of space.

This is the messy roommate. The messy roommate has a lot of stuff that's spreading out all over your room. And so instead of having 109.5 amount of space, you now have 106.7 amount of space in your room. Because it's just-- the messy roommates just spreading out all over the place, pushing down on those bonds and the bond's contract.

All right. Now, let's go back to trends in the periodic table for a minute. We learned that atomic size increases as we go down the periodic table because what increases? N. Principal quantum number N increases, so atomic size increases as we go down. Think of this as a messy roommate having more stuff.

So, the lone pairs now occupy a larger volume. Still messy, but now messy with a lot more stuff. And this stuff impedes in your space. So, the angles tend to be even smaller between the bonded atoms. So, we had NH3 before at 106.7. But nitrogen is up here. Phosphorus is below it, so it has a bigger atomic size and it has lone pairs that occupy larger volumes. So, this is what happens. [EXPLOSION SOUND EFFECT]

Some of you may have experienced that. That is a messy roommate with a lot of stuff. Now, if you have a roommate that's very neat, they're putting their clothes in drawers. When the clothes are in drawers, just like electrons and bonds, they don't really go anyplace. But when the messy roommate does not put their clothes-- when they're lone pair clothes-- then they go and this is what happens.

OK, so we can predict now by looking at periodic trends, and thinking about how much room those lone pair electrons are taking up, we can think about and predict the angle between the bonded atoms.

All right. So, this is pretty cool, but can it save the world?

What do you think, can VSEPR save the world? Do something important? Of course it can, it's part of chemistry. So let's think about a pressing problem in the world right now and how VSEPR can address it. And if you listen to NPR or open a newspaper these days, you're hearing about a lot of car bombs, explosions, things are not good in the Middle East. We're hearing about all sorts of-- I heard on the news about bombs going off at a school and then when parents rushed toward the school more bombs went off. I mean just really horrible stories.

And even when the conflict is over, a lot of times those improvised explosives are still there. They still exist in the countries. And it's estimated by the UN, by the United Nations, that landmines kill 15,000 to 20,000 people, mostly children, women, elderly who are out in the farm fields trying to grow food for their family and they step on something and blow up.

So, how do you find these explosive devices? How do you find explosive devices that are actively being used now in dangerous parts of the world? And how do you find the explosive devices left behind when the war is over? So, if you are Stephanie, you use VSEPR That's what you do to find those explosives. So, in her own words now I'm going to tell you-- or she's going to tell you, why VSEPR what she calls V-S-E-P-R, because she can say that better than I can, to find explosives.

STEPHANIE SYDLIK: My name is Stephanie Sydlik and I am a graduate student in Tim Swager's research group at MIT. The research that were perhaps the most well-known for is for sensing explosives, such as TNT. [EXPLOSION SOUND EFFECT] New, bigger, better explosives have been developed. And two of these are RDX and PETN. And these kind of have more bang for your buck, if you will. Unfortunately, they have an even lower vapor pressure than TNT. Which means there's even less molecules of the explosive in the air and it makes them even harder to detect.

The dogs that they would send out are actually sensing cyclohexanone and acetone, which are molecules that are used in the purification of these two explosives. Both cyclohexanone and acetone have a carbonyl in them. And this carbonyl then can interact with a group known as the urea.

We have two nitrogens connected to a carbon that has a double bond to an oxygen between them. And these nitrogens have hydrogens on it, that can hydrogen bond with that acetone or the cyclohexanone that we're looking for. What happens is that the lone pair from the carbonyl reaches out and grabs those hydrogens and pulls them away from the nitrogens. And this makes the nitrogen hydrogen bond more lone pair-like. As it becomes more lone pair-like, we see more repulsion between the lone pair-like bond of electrons and the neighboring bonds.

And by the V-S-E-P-R theory, we know that this is going to cause the bond angle to become larger between the nitrogen hydrogen bond and the accompanying bonds and smaller between the other bonds around the nitrogen. And this causes large scale changes in the polymer. So, we can see differences in the way in the large scale the polymer interacts with light as fluorescents, so it will start to glow. Or, a refractive index change, which is also a different way the polymer interacts with light. We have instruments that can very easily measure both fluorescents and refractive index. And with these very easy signals, we now know that our acetone or cyclohexanone and therefore, the explosive is there.

For soldiers, this is a really big deal. In Iraq and Afghanistan, there are minefields and improvised explosive devices almost everywhere. And the soldiers over there really have to watch where they step. So, if we can come up with a handheld device, and we have in the past come up with some. And I'm hoping that my technology might in the future also go towards these types of devices that will be attached to a robot and sent out to sniff out the area before the soldiers go there. You can really save a lot of the soldiers lives as well. It's very cool to do the hands on work in the chemistry laboratory, and then know that what you've done at your bench will then one day be actually used by someone and potentially save their life.

PROFESSOR: OK. So, back to VSEPR and lone pairs. Let's look at some examples and think about the shapes we've seen some of these already, but let's look at some more. So, now we can have AX2E, SN number of 3. So, two bonded atoms one lone pair. This has bent geometry, but again think about this lone pair.

Now, we're going to talk about the angles. And we're talking about angles in this class, we're talking about the angles between the bonded atoms. So, the angle from the lone pair down is going to be bigger because the lone pair is really repulsive, but we're going to be thinking about the angle between the atoms you see. So, when we ask about angle, we're asking about this angle between one bonded atom, central atom, and the other bonded atom. So, this lone pair doesn't look that repulsive.

So, I just want you to sort of think more about this. So keep this in mind. This is really more what a messy roommate is all about. So, if this is your lone pair pressing down on those bonds, what do you expect the angle is going to be?

What's the angle in the normal case first?

[BALLOON POP]

So, we have 120.

And now think-- and actually it's a clicker question, think about what the answer is going to be. We talked about the normal.

All right, 10 more seconds.

All right, yes, less than 120. So, you don't know exactly, you can't say "Oh, that's going to be 118.5 or something." But you can say less than 120. That's how you would express it. And I'll just remind myself to say, if you like these model kits and want your own, some toothpicks and gum drops can create some awesome VSEPR model kits. And we'll try to bring some of these into recitation for people who want to have gum drops and toothpicks for making models.

OK,

let's keep going. So, now we have our tetrahedral based system

AX3E

and an SN number of 4 based on tetrahedral. And so here, now, we have trigonal pyramidal.

So, we have a bunch, this is why it's confusing. It's not bi-pyramidal, there's only one pyramid here and it looks like a triangle. So, trigonal pyramidal. And now what are the angles going to be? And you can just yell this one out. Yeah, 109.5. And now let's keep going. And we have another clicker question.

All right, 10 more seconds.

OK.

So the trick here was to think about the parent geometry of the system. And so this is the parent geometry is the tetrahedral system. And we know that because it has a SN number of 4. And so, when you have a SN number of 4, then it's going to be less than 109.5. And this is called a bent geometry. So, again you can think about it within terms of those-- if you have some whole cans, those lone pairs are pressing down on the bonds and compressing them. So, it's less than 109.5. I just teach chemistry because I like to buy whole cans and have a justification for it, is really the bottom line.

OK, so if we keep going now, we have our friend seesaw, which you're never going to forget. And I'll rebuild my-- oops wrong one, rebuild my seesaw over here. So, now what are the angles here? There's two of them. Think about the equatorial angles. Yep, I'm hearing it, less than 120 and the axial would be less than 90. So, the lone pair is pressing down both on the 90 and on that 120. Probably more repulsive for the 90, but all you have to do is say less than for both of them. So, the trick is just think about the parent geometry. What are the angles in the parent geometry? And then, it's less than.

And if we keep going with this, we had our T-shaped molecule, as well. So, when we added another lone pair to our SN 5 case. And what's the angle now going to be with those two lone pairs? Less than 90. Right. All right. So, there so many possibilities for lone pairs.

If we add yet another lone pair to the system, what's my geometry?

So, this is now going to be linear geometry. And what's the angle going to be?

Yeah, it's just going to be 180. So, we don't have a less then here because whatever way it would bend it would be just moving toward more repulsive loan pairs. So, there's no way you can minimize the repulsion in this case. So, it's just going to be a linear molecule.

So, now we're going to move into our SN six category. And we're going to talk about a shape that we haven't talked about yet. So, based on this what happens? We have our parent geometry of octahedral. And so this has six bonded atoms, SN 6. But now we're going to take off one of the bonds-- put it the same as the figure-- we're going to take off one of the bonds and put on the lone pair. And this is called square pyramidal because you have this square here in your axial. It looks like a square.

But when you consider you have an atom on top coming down to these four atoms on the side, that again looks like a pyramid. So, this is square pyramidal. And what angle do you think you're going to have here for these bonded atoms? Yeah, that will be less than 90.

All right, so we can keep going and we saw this one before. If we take off another bonded electron and put it in a second lone pair. As we saw before those lone parents want to be a far apart from each other as they can. So, one goes on top, one goes on the bottom. And this was our square planar geometry because it is so square and it's planar. Now, what do you think the angles are? 90. Right. There's no where to escape when you have a messy roommate on top and a messy roommate on the bottom. If you're in a triple between messy roommates, there's just nothing you can do. And you can't minimize the repulsion at all. You just have to live with it.

OK, so if we keep going again and now we're going to add another lone pair. And it really doesn't matter where we put that, it's all equivalent. And it comes up with the shape, that's T-shaped again that makes sense when you look at the structure. And now what do you think the angles are? Less than 90, and you would be correct in that. All right, so if we put on yet one more lone pair and take off a bonded electron or take off one more bonded atom and put on an electron. What's our geometry? All right, so we have this structure, it's linear 180 no place to go. So, we keep going far enough to come back to linear a lot.

Now, let's just look at a couple more-- some real life examples of molecules and think about their geometries and shapes. I have a couple more clicker questions, so let's start with our friend water. And I had water here. So, now let's think about what the formula type is going to be for water. And what's our formula type?

AX Yup. AX2E2, two bonded atoms, two lone pairs SN number of 4. And do you remember what this geometry is called? Bent, yeah. And this explains then what we talked about before that this is a polar molecule. So, these are polar bonds between the oxygen and the hydrogen. But in this case, it's not a linear molecule it's a bent molecule, so it creates a net dipole. And so that makes it a polar molecule.

And if water had any other shape and was not a polar molecule, then life would be entirely different because water is the solvent of life. So, this shape pretty important for anything. Actually, they researched on what medical doctors thought was the most important topic to learn as an undergraduate in premed education. And the number one topic that was most important was water. There, and you just learned about. It OK, so we keep going now and you have to answer a lot of things about this on a clicker question.

All right, just 10 more seconds. You got to finish your hand out.

All right, so let's go take a look at this one and fill it in. So, we have our AX4, which is a SN 5 system because we have one E. And then we have our seesaw geometry. All right, so let's just fill in the rest here and see who's won the clicker competition. So, for the next system we have Br. And now it's expanded, so we have AX3E2 SN 5. We've added another lone pair. And now this makes a T-shape, but it's kind of a little benty shape because of the repulsion. We come down. We have AX2E5 SN 5.

We've added three more with xenon. Xenon is expanded here. We have our linear shape. We also have xenon with four things bound. And if we did the Lewis structure of that, we'd realize there's lone pairs on the top and the bottom. AX4E2 SN 6. And this is our square planar geometry. So, you can predict a lot about just doing a Lewis structure or thinking about where the lone pairs. You can predict geometries. And let's see who has won today.

We have an upset.

All right, Sam.

All right, see everybody on Monday.