Hi. I'm Jocelyn, and we're going to go over fall 2009, exam 1, problem number 3. With every question, we want to make sure we read the full problem first. So answer the following questions about the difluoro-iodate ion. Draw a three-dimensional representation of the molecular geometry around the central atom, not simply the Lewis structure. Show all atoms and bonds between them.

Because this part seems pretty separate from the rest of them, we're going to start with that. So we have a difluoro-iodate ion. It has a positive charge. And it asks us to draw the molecular geometry in a 3D representation, and not just Lewis structure. However, to do the molecular geometry in 3D, we need to know the Lewis structure. So we're going to start with that.

To draw the Lewis structure, we first need to figure out which atom is the central atom. And it's always a good idea to do that by looking at the relative electronegativities. As we know, fluorine is the most electronegative element in the periodic table, and so that's probably not going to be in the middle. Therefore, iodine, being the less electronegative, is going to be in the middle, and it will have fluorines on either side.

This is our rough sketch of the molecular geometry. Not too important right now.

Next we need to know how many electrons we have. Iodine has seven valence electrons, fluoride has seven valence electrons, and we have a positive charge, so we know that we're missing one. And that gives us 20 electrons.

Now that we know how many electrons we have to work with, we need start filling them in. I always start with the outer electrons. Those are usually, hopefully, the most electronegative, and therefore, the most likely to have the octet rule satisfied. Also, if you remember, fluorine is in the second row, and therefore needs to have the octet rule satisfied. Iodine is a little bit lower in the periodic table, so the octet rule isn't as important. It can have an extended octet if we need it to. So we will work with iodine second.

Let's start out with satisfying the octet rule for the fluorine, putting first the seven valence electrons that the fluorines brought themselves, and then noticing that each needs to share one from iodine, so that we have a full octet. So each has one bond with iodine. We've used 8 electrons, or 8 per fluorine. So we have a total of 16. We have 4 left, and we're going to put those on the iodine.

Counting up how many electronics are around iodine, 1, 2, 3, 4, 5, 6, 7, 8, we see that the octet rule is also satisfied for iodine. This is looking like a pretty good Lewis structure.

Now, just to make sure it's really good, we're going to look at the formal charge. So if you recall, the formal charge is the number of valence electrons that the element usually has, and minus the number of unshared electrons, and the number of dots.

So this sum subtracted from the valence electrons will give us the formal charge. Looking at the fluorine, we have seven valence electrons. We have 1, 2, 3, 4, 5, 6 unpaired, and one bond to the fluorine. So that gives us a formal charge of 0.

The same goes for the other fluorine. So I guess its formal charge is 0. Now for the iodine, we have-- I'm going to do the formal charge down here-- it again has 7 valence electrons, usually. We have 4 unshared electrons, and 2 bonds. That gives us a formal charge of plus 1.

Because we have a charged species here, the full molecule has a plus 1 charge. There's no way we cannot have a net formal charge. Right? We need to have a net formal charge of plus 1. And it makes sense that that formal charge is on the least electronegative atom. So this looks like the right Lewis structure for our purposes.

One more thing that you might want to write down. I'm not sure if points were taken off from this. But technically, if you have a charged species, put brackets around your Lewis structure, and put the net charge on the outside.

This is not, however, our final answer, right? This is the Lewis structure. But the question asks for the molecular geometry to be represented. So if we remember how we determine molecular geometry, we need to look at the number of electron domains around the essential atom.

So we're going to move over here, and I'll redraw our Lewis structure. And remember, an electron domain can be either an electron pair, or a bond. So we see that iodine has 4 electron domains. And if we remember our skeletal geometry, we know that that's tetrahedral.

Thus, to draw this in a more 3D fashion, we might want to say that one of the electron pairs is here, one of the electron pairs is over here, and then we have a fluorine coming out of the board, and a fluorine going into the board.

We didn't really go over how to do three-dimensional drawings. So as long as you show something like, you know the fluorines will be an angle, not straight from the iodine, and that the electrons be on the other side, something to that effect. A little bit more detailed than just a Lewis structure, was what Professor Sadoway was looking for in this problem.

And that would be your answer. And for this one especially, I'd say it's important to box your final answer. Because you hopefully would have figured out the Lewis structure, but this is not the correct answer. So you always want to signify what answer you want to be graded.

Part B says, name the type of hybrid orbitals that the central atom forms. So we are almost all the way to answering this question, so I'll put it down here. Remember, hybrid orbitals are-- we have hybrid orbitals, because we can't really say what orbitals exactly are bonding to which atoms. We want to have 4 equivalent orbitals for this system, right? We have 4 electron domains, so we want 4 equivalent bonding orbitals. And from our talks earlier in the lectures about hybridization, we know that if we want 4, we take 1S and 3P, and that gives us the hybridization of the SP3. if we only had 3 electron domains, it would be SP2, because we only use 3 orbitals. And the number of orbitals that go into the hybrid is the number of orbitals you will get out.

So we know that our hybridization is SP3. And that goes along with the fact that our shape is tetrahedral. Now we need to name the molecular geometry of this molecule. so we already have the molecular shape and the kind of spacial arrangement is tetrahedral. However, when we're naming the molecular geometry, electron pairs no longer matter.

So for part C, tetrahedral is our skeletal geometry, right? It shows us our special arrangement. The molecular geometry is ignoring those electron pairs. So it is called bent when we only have 2 bonds. And I think about that because in spectroscopy and ways that are used to study molecules, you can't see the electron clouds very well. So they could only see the iodine and the two fluorines, thus looking like a bent structure.

So hopefully you're sticking with me here, because this problem has a lot of parts. And we're going to move on to part D now.

Part D asks, is difluoro-iodate polar or nonpolar, and explain. If you put down polar or nonpolar and did not explain, we assume you guessed, and you didn't get any points. Explaining shows that you know why, and you understand the concepts behind this, and thus deserved points on this problem.

So there's two different parts of this problem. First, to have a polar molecule, you need to have polar bonds. So I would first ask myself, are the bonds polar? We have a difference of electronegativity that you can look up on your periodic table and know is not negligible. And so yes. A difference in electronegativity signifies that our bonds are polar.

However, not all molecules that have polar bonds are themselves polar, right? if we have spacial symmetry, those polar bonds could cancel out. So now we need to make sure, is the molecule polar? And for that, we need to have answered the first questions correctly. Right? We need to know that this is not a linear structure. If this was a

linear structure, those polar bonds would be geometrically symmetric, and cancel each other out. But we all are very smart, and we got this part C right. We know that we have a bent structure. Thus, I'll rewrite this over here. We know that if we draw in our polarity vectors, we can realize that we'll have a net dipole pointing down. So the answer to this is yes, because of the difference in the electronegativity, and because of the spacial asymmetry.

If you said both of those things, you would get full credit. If you said one or the other, you'd probably get part credit. But knowing to look at these two things shows that you fully understand what it means to have molecular polarity.

All right. So now we're going to move to part E. This part is changing gears a little bit, so if you didn't quite understand the first part, you can still try this one. Part E, again, what is the question asking us? Determine the maximum wavelength of electromagnetic radiation capable of breaking the IS bond.

So we need to find the maximum wavelengths which corresponds to the minimum energy. Right? Because we have e equals hc over over lambda for electromagnetic radiation. We're trying to figure out the lowest energy of photon that will break the IF bond.

Now, what are we given to answer this question? We're given the energy of the homogeneous fluorine-fluorine bond, is 160 kilojoules per mole. And the energy of the homogeneous iodine-iodine bond is 150 kilojoules per mole. And in order to find the wavelength that will break the IS bond, we need to know the energy of the IS bond. Right? That's what we're working towards in this problem.

So given the homogeneous energies, do we have a relationship that will help us determine the energy of the IS bond? If you recall, in lecture 9, Professor Sadoway went over the Pauling formula for determining a heterogeneous bond energy from the homogeneous bond energies, and the electronegativity. So I'll just write that up here. The Pauling formula is, in a generic form, the energy of an AB bond equals the square root of the product of the homogeneous bond energies plus a constant times the difference between the electronegativities squared.

And I just said what all of these terms mean. Because even if you had this on your equation sheet in the test or something, a lot of people wrote this down, and then used it incorrectly. So it's really important to know not only what equations you have, but especially what each of the terms in the equations mean.

So we're given the homogeneous bond energies, and we have a periodic table, or some other resource, that has the elecronegativities. Next step is to plug in the numbers and find the energy of the IF bond. So we have the square root of 150 kilojoules plus. And the exact values you get for the electronegativity may vary, depending on the source. But they'll be close enough, and relatively, probably very-- the difference between them will be very

similar. So don't worry if your numbers were a little different than mine. So we wrote all that out, and plugging it into the calculator, you get that the IF bond has an energy of 323 kilojoules per mole.

Now that we've found the energy of the IS bond, are we done with the problem? You have an answer. You might be ready to move on to the next question. However, this isn't what the question is asking us, right? We are asked for the maximum wavelength that can break a bond of that energy. So we need to go back to our energy relation to our wavelengths. And we know that the wavelength equals hc over the energy.

Now here's where some people tripped up a little bit. You have a value for the energy. h and c are constants. And you find the lambda. However, if you plugged that all in, your answer would be 6.15 times 10 to the negative 31 meters.

Some people, that was the answer that they found. And it makes sense. You have an energy, you have some constants, and you can get your lambda. However, if you think about it, does this wavelength actually make sense? The gamma ray radiation has a wavelength of about 10 to the negative 17. That's very, very, very energetic electromagnetic radiation. So this is 14 magnitudes smaller than gamma ray radiation, and it just doesn't make sense for a bond energy of this much. Plus it doesn't-- it just-- you should look at the answer you get, and try to be comfortable with if it makes sense or not.

So where people got tripped up, is in the units. We found Pauling's formula gives you the bond energy in kilojoules per mole. However, if you look at Planck's constant, it has units of joules time seconds. The speed of light is in meters per second. And you want lambda in meters.

To make everything or cancel out, that means we want the energy to be in joules. Not joules per mole, because we're looking at the wavelength of one photon, right? We're looking at the energy of one photon to break this bond. So we want the energy of the bond, not per mole of bonds, but just of one bond.

So instead of just having this equation, we want to have put in h times c over 323, change it to joules, and then multiply by-- sorry, moles are on the bottom, 1 mole and Avogadro's number. And Planck's constant of the speed of light, you should have on your equation sheet, or somewhere else, some other resource that you have. You don't need to memorize those in most cases.

If you remembered to divide by Avogadro's number, you got 3.7 times 10 to the negative 7 meters, which equals 370 nanometers. And if we recall, the light we can see is in the 400 to 700 nanometer range. So this puts us in ultraviolet radiation, which makes sense that it would take that much energy to break a bond.

The other problem some people had was forgetting to convert from kilojoules to joules. That would still give you a physical answer. But again, I can't stress unit cancel out enough. So make sure you know that since your Planck's

constant is in joules, you want to make sure you have joules on the bottom.