>> We can use the VSEPR system to look at four electron group molecules. We have three of them here. The first one is CF4, the second is NH3, and the third is H2O. Each one of them we're looking at individually because each one actually ends up forming a different shape molecule. So first off, let's look at what we have up here. When we have four electron groups, the bond angles between the electron groups will automatically be 109 degrees. Actually it's 109.5, but we're going to go ahead and just call it 109 degrees. This forms a tetrahedral. Here is a tetrahedral with the ball and stick models. You can see that no matter what I have at the top, which of these yellow balls, they're all going to look the same because they are. All those angles are the same measurement. If you look at this also with, this is a different kind of a model. It's a space filling model where now you've got the black ball being the central atom and the white ones being, in this case, the fluorine. Ok, this, now you don't see the bonds themselves, but this really is more of a representation of what the molecule looks like because you're actually seeing that the electrons clouds will overlap with each other. So you really don't see, per se, a stick. It's just an easier way of representing what we're looking at. Ok, so we know the electron groups will be 109 degrees apart from each other, forming a tetrahedral. Let's look at our first molecule, CF4. That's carbon tetrafluoride. We're going to put carbon in the center, and we have our electron dot structure here. Ok? And of course that's very two dimensional, so we need to make something three dimensional. Let's go ahead and put in our lines. This is actually how you form your tetrahedral. It's very important that you remember to use dotted lines for something that's behind the plane of the board and a wedge to show something sticking out of the plane of the board. That indicates that it is three dimensional. Ok, so here's your carbon, your four fluorines. Two of the lines are in the plane of the board. That's how you depict a tetrahedron. Since each one of these electron groups ends with an atom, then the shape of this molecule will also be tetrahedral. And so any time you have electron groups that form a specific shape, as long as each of the groups ends with an atom, then it's going to be the same answer. The geometric shape, the molecular shape will be the same as the electronic shape. Ok, on this one now, let's decide whether it's going to be a polar molecule or a non-polar. Looking on your table, you'll notice it's a 4.0. When you subtract those, you get 1.5, and 1.5 definitely is in that polar range. So we have a polarity where there's, electron density is being shifted up to, there we go, up to each of these fluorines. Up to that one, down to this one, toward us over here, and behind the plane of the board there. Ok? So you might say that's a very polar molecule, in which case, well, individually those are very polar bonds. But the molecule itself, overall, is a non-polar molecule. Any time you have equivalent atoms range 109 degrees apart from each other, they cancel each other out. Ok? So this would be a non-polar molecule. Now let's look at our second one. Our second one is NH3. That's ammonia. When you do your electron dot structure, you'll notice that there's a non-bonding pair of electrons here. There's only three atoms even though there's four electron groups. So when you draw your tetrahedron, one of those lines is going to result in having just the two electrons at the end of it. So now when we look at this, if you were to look at that with

the space filling model, this is what you would see. Ok? Here's your nitrogen and the three hydrogens coming out at angles. This actually forms a pyramid, which is probably a little bit easier to see if we use this model here. Now this was our tetrahedron. If I take out one of these sticks, this would represent what the molecule looks like. Here's nitrogen. Here's each of the hydrogens. And up here, we would have the electron cloud, which you cannot really see because the electrons are so very, very tiny. All right? And so this would form a trigonal pyramid. We would say the shape of the molecule is trigonal pyramidal. If you want to just call it pyramidal, that's fine too. Ok? So this is a pyramidal molecule. Notice it was tetrahedral arrangement of the electron groups, but the shape is pyramidal. Why? Because one of these is a non-bonding pair. We only have three atoms that are in part, as part of the four electron group. Ok? Now the next question we have to ask ourself is, is this a polar molecule? Ok? Here again are our electronegativities. Nitrogen has electronegativity of 3.0, hydrogen of 2.1, so there's a difference there of .9. Remember a polar molecule starts at .5 and goes up to about 1.6. So that's definitely going to be a polar molecule or at least polar bonds. Let's check and see if it is a polar molecule. The nitrogen is the more electronegative, so we're going to see electrons being shifted from the hydrogens toward nitrogen. In other words, we have these electrons being share between nitrogen and hydrogen. It's not shared equally, so nitrogen is kind of pulling them more toward itself, leaving each one of these hydrogens with a partial positive charge. What happens now is you have a partial negative charge in this whole area, not only because you're pulling the electrons toward the nitrogen, but you also have that non-bonding pair. So if I took all three of these vectors and averaged them, they would actually form a resulting vector that was going in that direction. And what this tell us now, if we take these out of the picture and just look at this, this would be our dipole moment. And what this is telling us is the dipole moment is, this whole area down here is positively charged, and up on this end it's negatively, partial negative charge. Ok? So that's our dipole on that one. It's a polar molecule. If you look at the third one, this is water. Again, looking at this, we might, you know, if we don't want to do the work we're going to say, oh, oxygen with two hydrogens, that's linear. And that would be wrong because oxygen with hydrogen actually has four electron groups. Yes, it does only have two atoms, but it also has two non-bonding pair, which means that putting that for it three dimensionally, again, we're going back to something that looks like this. To begin with, this would be the four groups, but now look what happens. We've got non-bonding pair there and a non-bonding pair here. So I'm going to go ahead and take off two of the four sticks, and this is what we're left with when I take off those two extra sticks. We've got a bent molecule. Water is bent because coming up from this way is electron cloud. Coming up from this way is electron cloud. Here's another way of looking at this. Your – – well, the other one. Ok. Now let's go ahead and take a look at this. We said it's bent. Let's take a look and see if it's polar. Oxygen has the electronegativity of 3.5 and hydrogen of 2.1. So when we subtract those, we get a difference of 1.4. Definitely a polar molecule. And we actually know that about water because water is one of our best solvents.

Anything that's a polar that goes into water will usually dissolve really well. And it's because of water's polarity that it allows it to dissolve. So let's take a look at what we've got here. We've got oxygen is electronegative, so we're going to have electrons shifting from the hydrodens toward the oxygen. We also have an electron cloud here, electron cloud here. This whole thing now will average out to give us electronegativity overall for the –