Okay great! So yeah, so let's please settle down. I'm going to get started. Great. So just a couple of quick announcements here. You have your quiz due on Friday just as usual. You have your review on Friday as well during the class time as usual. we will not be able to finish all the problems that are posted for this week today. Okay just because Monday's class was cancelled so that does impact how many problems we can do today. So, what I will do is leave the rest of the problems and we continue with them next week because lecture 15 there are not so many problems that we will have. Okay? So, I'll post all the answers together, I'll actually post this even now just so you have them I'll post the answers even if we don't complete them this week. And then next week, we go over everything we didn't finish and do the lecture 15 problems. There is no review next week. So just to make it clear next Monday, Tuesday, Wednesday, there are no classes, right? Yes. I feel the same way. And but next Thursday is a Monday schedule, so you have a class in Leacock-132 at 10:30. So I'll remind you again but just please make sure that you are aware that that Thursday is a Monday schedule, Okay? There is no review next week so there's nothing graded next week. All the schedule dates are already up so you should know when each review is. There is no review next week, Okay? There is only a review this week, we take a little bit of a break. I think we have two weeks break and then we have another one. Great, so that's it for the announcements. Okay, let's settle down.

So, I don't know if you've seen the news, but in the morning today the announcement of the Nobel Prize for chemistry so of course I must make an announcement in class. So, the Nobel Prize is shared this year between three chemists, there's Carolyn R. Bertozzi at Standford, there Morten Meidal and Barry Sharpless. And each of them is one third, one third, one third. Also, it's it's clearly unique because I think Caroline is professor Bertozzi one of the I think eight women to win the Nobel Prize. But professor Sharpless, this is his second Nobel Prize, so he won the Nobel Prize in 2001 for another really and also in chemistry. So, this is his second Nobel Prize in chemistry for this and what I do want to point out here is that the the reaction right so there is this description that comes from the Nobel committee which explains why they won the Nobel Prize, and it is for the development of click chemistry. And in very simple terms it is it really is just you know two compounds clicking to form something which is super useful. Now, what is interesting about professor Bertozzi's research is the fact that she utilized that click chemistry to look at mammalian cells and you see the word "bioorthogonal" mentioned. What that really means is they've done chemistry on the surfaces of the cells in these biological samples without actually disturbing the implications in terms of what they are trying to study. So, you are doing indeed chemistry where you're able to label say a cell with a certain fluorophore or certain compound that lights up right under conditions that you would like, but you're not actually disturbing the functioning of that particular cell which is you know it is mind blowing. It was unthinkable even you know 15-20 years ago, until this chemistry was made possible. We, you know, if you continue learning chemistry these cycloaddition reactions that are common to this click chemistry are actually very common to learn in higher level organic chemistry. Clearly, right now is not exactly what we are learning okay, so we don't we don't really learn these reactions. But when you learn about shapes of molecules and when you learn intro organic chemistry in CHEM-110, you'll learn something about… So, this, this group right here is an alkyne group. You won't usually see it within our ring okay, and there is a reason for that: It's very constrained, it's extremely unstable. So, you will learn why that would be unstable and by adding these electron withdrawing floorings right so we talk about how flooring is so electronegative. So, these electronegative right here meet this even more reactive and the purpose of these reactions is to make it more reactive so you can get this particular reaction to occur. So that was pioneered in in professor Bertozzi's lab and there are lots of, so these species are aside species, a lot of these cells contain these aside species. You are able to, then, do these reactions with these particular molecule and attach say some kind of a label to it and by doing this chemistry on the top of the cells, you get you get some kind of labeling on the cell that you're interested in. So, you know, sometimes we can't always think about why this would be relevant, why something so complicated would be so applicable right? It's extremely applicable because it tells us so much about these mammalian cells, about any sort of cells as well, which is made possible by this amazing chemistry by all these three amazing scientists okay? So great, well congratulations to all three of them! Yeah, that's pretty cool. Okay. So, so coming back to what we are trying to build upon, much much you know much less complicated. So, so we will start to talk about this week, we're going to build upon something we've already learned last week. We are going to start now talking about each covalent bond okay? So yeah, I'm not yet going into Lewis structure. I'm just going to show you molecules, we're not going to talk about the shapes of these molecules either but what we are going to talk about is what determines if a molecule will have polarity okay? So right now, we're going to start small. Last week, we talked about polarity of individual bonds okay, so we will talk about okay, if I give you a molecule, first we also learn right, later on we learn how to do, draw Lewis structures but if I gave you the Lewis structures, could you tell me if those bonds, not the molecule but just the bonds are polar or not okay? But what we rely on is this electronegativity chart right here and you will have this. So for any exam will be given of course for the midterm it's open book, so you have access to your notes but even for the final exam, if you need it, you'll be given the important electronegativity values okay? So, you don't need to memorize these values. Always good to know that nitrogen, oxygen, fluorine, really really highly electronegative elements okay. That's always good thing to know. Okay so in each of these cases, what you really are looking for is whether there is a difference in electronegativity which is greater than a certain threshold. We're looking at something that's below so that's above 0.5 okay. So, you want to at least hit that threshold and I'll just give you an example so for instance oxygen is 3.5 and hydrogen is 2.1. So, this bond, the difference in electronegativity is 1.4 which is you know quite high. This is indeed a polar bond. So, I'm just going to circle the polar bonds. Maybe it's better if I just circle it. Okay, so this is indeed a polar bond. Carbon is 2.5. So again, carbon is here, the electronegativities 2.5, oxygen electronegativity is 3.5. This is also a polar bond. The single and double bonds, when you're calculating the delta chi, will not matter if you're just calculating the difference between those atoms okay. It does affect what the shape of the molecule is going to be which in turn will affect the polarity of the entire molecule, but here we are just looking at bonds. Yes. No, so the question is: "Is there a unit for electronegativity?". No. So each of those, we when we talk about it, we're just giving an absolute number value which are calculated from energy calculations but there is no unit associated okay. The carbon-carbon bond, both of them are the same so this is not going to be electronegative. The carbon-hydrogen bond has a delta chi of about 0.4 okay. It's not… No carbon, if you just have a molecule which only has carbon and hydrogen bonds, there's no it will not be polar, the bonds are not going to be polar. Focus, so just remember that. Carbon-nitrogen however… So, nitrogen is 3.2, carbon is 2.5 so this is just under the task of 0.5 so it is indeed polar. And then you have these nitrogen-hydrogen and nitrogen-hydrogen bonds. Okay? And the nitrogen-hydrogen bonds, the delta chi's about 0.9 and this is about 0.5. And we have talked about this, this is about 1.4. Yes. No, so I just mentioned that it does not. It's exactly the same because you're just looking at the two atoms, okay? Great, so I I'm going to take a couple of other questions, yes and then I'll take yours. Go ahead. Nitrogen is three and hydrogen is 2.1. So, the difference between them, that's what I'm showing is 0.9 okay. Yes, go ahead and I'll come to you after that. Okay so the question is "you have something wich is semi polar or…". So experimentally yes, there are some that are in between but we're not going to use that terminology. We'll either call the bond a polar bond or a nonpolar bond and you can use the definition if it's more than 0.5. So, 0.5 and above delta chi, you call that polar bond. If it's 0.4 and below, it's going to be a nonpolar bond okay and that's it. And I know you had a question, yes. Because it's 0.4 okay. Great! Any questions from the back? Okay so in this, the most polar bond, the highest delta chi that's the OH bond. Okay now for the second molecule, so for the second molecule that we have, again carbon-hydrogen nonpolar, you have carbon, oxygen, you have carbon, oxygen and you have OH. Okay and I think I forgot to put the delta. So here the question, on the second part of question is "indicate delta positive and delta negative for each of those bonds". So, in this bond… So, your less electronegative atom would be delta positive. So, we call, we use the term the small Greek symbol delta so small delta to indicate that it's not, it's only a partial positive charge okay, it's not a complete… So, nitrogen is not withdrawing all the electrons. It's still sharing the electron; it just has a higher share of those electrons okay. So will indicate that with the delta negative. So, delta negative will go to the more electronegative atom, delta positive will go to the less electronegative atom. And we do the same thing here: the oxygen is going to be delta negative, delta negative. Yes. Using examples just like this, so I'm not trying to, you know, give you one but it could be delta positive one bond and delta negative no, you'll see examples like this. Yes. Okay, any other questions? Okay, so so nothing really. Here, look at the electronegativity chart. Always, anything with nitrogen, oxygen, fluorine, you very quickly would be able to determine that it's going to be electronegative since those are extremely there like the higher electronegative species okay. In the final one, there is only one polar bond, it's carbon and fluorine. So that single bond is going to be the most electro, the most polar bond as well. The difference is about 1.5 and any bond with chlorine really is most likely going to be a polar bond okay. Not even most likely, let me see the most yeah, if it's going to be a polar bond. It will fit our definition as well. Right, so any atom of course fluorine, fluorine F2 not even a polar bond exactly the same but if it's with another atom it would be. Okay? Okay, great! So, this is the next question where we are now getting into… So now that we now know what are covalent bond is, we can also determine what it means for an atom to be electronegative. It's going to pull electrons towards itself if it's more electronegative than the other atom. So that's unequal sharing. But really, what we also want to be able to do is, if you're given a molecule, you are able to draw what the bonds are going to be okay. The one thing I want to point out, like examples like this, this is a little more challenging okay to begin with. So, I'm telling you this, the reason why is when you practice more and more of these, you'll become familiar with certain compounds okay and as you become familiar, this will become easier. So, you know, I can very quickly like when I look at the molecular formula, I can very quickly say "Okay, on the left, there will be a nitrogen with two hydrogens bonded to carbon". Okay, I I know it because I've seen this enough that I just know it now. It might not come so naturally to you just yet. The whole point is to keep practicing and then it will come, you know it will come to you and it will, it will start to look familiar okay. For examples like the right, very often I'm not going to give you just the molecular formula okay. I will very likely give you a structural formula and what I mean by that is… And you know, we haven't covered this just yet… And we'll come to that… So, this is exactly the molecular formula still is C2H402, but I am also telling you what the structure so this will help you determine the structure. So, once I draw the Lewis structure, I will explain to you what I mean by that notation okay. So so just you know just be patient and this will, it will come to you okay. And the second thing that I want to point out is only when we have tried a few, so the first few times you do this, there will be a few different structures that you might draw which may be acceptable Lewis structures OK and that's okay. But chemically, we are really looking at molecules that have a very specific Lewis structure okay. So, until the time that you're giving me an answer which is a valid Lewis structure octets are complete, you have now organize and make sure that all atoms have their octets and are, you've accounted for all the valence electrons. Even if your organization is not exactly this, it is going to be acceptable okay. So, I'll come to that in just a second. So, the first example is CH5N okay. Now, what you want to keep in mind, the hydrogens are going to be terminal okay. Carbon, at most, can have four bonds, always. When the molecule is neutral overall, you want to give your carbon 4 bonds. Could be doubled with two single, all four single, all of those options. And you want to give you a nitrogen 3 bonds okay. So, so we see that as well, so we have the carbon. Carbon is going to be central. Carbon or nitrogen, either of them, the main idea being hydrogen can never be a central atom, right. So, your carbon has four valence electrons, so it's going to form four bonds. Your nitrogen has 5 valence electrons okay. So, you can see there are five hydrogens. All of them can't go on the carbon right, there'll be too many hydrogens for carbon okay. Then, the carbon will be forming a bond with the nitrogen, leaving only three other bonds to be formed. So, you can just… Okay, now that leaves behind two more hydrogens, they will form the bond with your nitrogen. You do not have to show the dots if you don't want to, except the lone pair okay. I do expect you to show me any lone pairs when I ask for Lewis structures. You don't have to form… You can just show the bonds here, not necessarily not necessary to show dot and dot and forming the bond between okay. Yes. You can draw in the bottom there is nothing wrong with it. So, this is the same, you can draw the hydrogen on the bottom, on the right. We are not yet talking about geometry at all okay. That's good, just make sure when you show this, you can also show it like this… Okay, that's totally fine, just the bonds only with the lone pair on that nitrogen. Now, when I say right, when the molecule is overall neutral, your carbon will have four bonds as it does its octet is complete. The nitrogen, when the molecule is overall neutral, most often will have three bonds with a lone pair okay. Those three bonds can be two single, one double or triple bond, any of that. Can somebody tell me why that would be the case? Yes, go ahead. Okay, almost there, almost there. I think I, there I also saw a hand in the back yes, go ahead. The formal charge on each of them, so I think you're describing what a formal charge would be. You're, the formal charge on each of the atoms here is zero. We learn how to calculate formal charge as well, but the formal charge on carbon here is zero, on nitrogen here is zero, on hydrogen with one bond is always zero okay. So, when you're looking at neutral molecules which don't have an overall charge, you're not… You know, so we use the term minimize formal charge but really what you're looking for is zero formal charge on each atom. That's most preferable. You want to complete the octets and then look at the formal charge, okay? So, we look at the calculation and then you come back and calculate the formal charge in each of these atoms. Now the second example, I will confess, it's a little more complicated okay. So, there are, there are few correct ways to draw this okay, so I would need you need to give you more information so that you are able to draw correctly but I'll show you the one that I have drawn. So, our carbon has four, there's another carbon which also has four and then we have one oxygen and then there's another oxygen, so I'll just come to that in just a second okay. Now, if we form the bond between these two carbons, right the carbon on the left, it either needs, it can bond with another oxygen or it can bond to three other hydrogens. It can form 3 total bonds. So, I am going to form the bonds with the hydrogen. So that carbon and its octet is complete, I've accounted for CH3. I've accounted for one carbon and three hydrogens. Now, this oxygen we had spoken about, oxygen to keep its formal charge zero, that oxygen needs a double bond. It needs 2 total bonds okay, so one will be with this carbon. But the carbon still has one electron, and it hasn't found its four total bonds, its only got three. The only other atom remaining is an oxygen, so I'm going to show it at the bottom, I could've shown it at the top… The oxygen has six, but I haven't still given you the hydrogen so that's the final hydrogen okay. If you want to draw this molecule, you will have at least one double bond. There are other different ways to draw this, which will still be correct, but you need at least one double bond. You have to do it, otherwise either your atoms will be off, or your number of electrons will be off okay. If you are say trying to put the oxygen in between the carbons, also you can try to do it, but you just have to make sure that you've accounted for all the bonds okay. Yes, I see a hand in the back. I think they should be, see OOH, sorry about that. That's okay, that was a mistake on my part okay. So, if you see this formula right here… Sorry, so initial formula was incorrect. So, if you look at it because I think I didn't see 2H502 so that's not correct. So, you see that this is CH3COOH, so very often when you see… When I give you a question like this on the exam, I'll be giving you this right here okay. So that's an indication that the first carbon is bonded to three hydrogens, the next carbon is bonded to two oxygens and then their oxygen is also bonded to a hydrogen okay. So, you can use this to draw this okay. And I thought I saw a hand… You're good? Okay. So just keep in mind, I know I gave you a sort of very abstract way to start the Lewis structure but there may be other options but if the molecular formula is given like this, you want to make sure you draw it the way it's written out. That's an indication to you how to draw this out okay. Just building upon this and I think that might be one of the questions later as well, what you can try to do is practice the molecular forms, practice the Lewis structure for CH3COO minus. It is directly derivable from the top right structure okay, but just try that out on your own. Make sure all the electrons are accounted for and then you want to calculate the formal charge in each of the atoms okay. When you add up the formal charge on each of your atoms, it should add up to the overall charge on the molecule. So, in this particular case, formal charge in the carbon is 0,0,0, 0, and 0 on all the hydrogens. So overall it's zero as well and we calculate formal charge not in the next question but very very soon so we will go over that as well okay. Yes. So if I say this is the molecular formula, there is only one structure possible right? If I just give you a random like this molecular formula, if there are multiple possible, then you can give me based on the structure you drew which was correct okay. So, in the next question, we are going to calculate the formal charge on carbon in three different states. So, we are starting out you know, this is kind of the same atom but in three different states and how in organic chemistry this is what you will notice from time to time for these molecules okay. So, I'll do the formula for formal charge and then we can we can see from there. So, for this particular case, so I'll label this A, B and C, okay? So, we'll start with A first. The formula that we use is… So, you want to calculate the number of valence electrons, minus how many lone pair electrons there are, minus half of how many bonded electrons there are or you can call it shared electrons as well. Okay, so for the first one, the number of valence electrons is 4, so for that for an atom stays constant always. Carbon will always have 4 valence electrons okay. The number of lone pairs; there are no lone pairs on this carbon and either you can count the shared electrons as number of bonds that's correct too or the numbers… Sorry, half of the shared electrons is just equal to the number of bonds. So, in the case of this, there are three bonds or 6 shared electrons. So, the formal charge on this carbon is actually +1. It is electron deficient okay. So, in this particular case, if I was looking at this particular model, this particular ion, this carbon cation, the overall charge would be plus one in this case. So, if you do the same thing for carbon

B, the only thing that's changed is, now it has four bonds, so 8 bonded electrons. So, I had said previously, if you have, a carbon with four bonds always has a formal charge of zero okay, every single time. There is no, there is no lone pair shown. Yes. Okay so, if you're giving something a formal charge, it does imply that its octet is not complete, that's why it has a formal charge. So, the first one, the object is incomplete, that's why it has a formal charge of plus one okay. It only has three bonds, no lone pairs. If I ask you to calculate the formal charge, either I will say something very specific "lone pairs are not shown", okay. But if I'm asking you to calculate the formal charge for something like this, I'm showing you everything you need to know to calculate that formal charge. Now for carbon C, I'll show that on the top. So, we can see carbon C. So, it has 4 valence electron. Here you have lone pairs, so that's two lone pairs. And there are 6 shared electrons. So, this carbon has a formal charge of minus one okay. So just make sure this is clear when you try to do this. You should be able to do this not just for carbon, but you can go back this Lewis structure and you should be able to do it for all nine non hydrogen atoms. Okay, so just make sure you are able to do that. Yes. I'm sorry? No, I don't really… You're saying the octet is not complete for the formal charge is zero, boron is an example. Yeah, you can have that. It just depends on the number of valence electrons, but you will always be able to calculate it. Not with carbon, no. Yeah, yeah. Okay. Okay, so the question was will you… Okay, so I'll ask you all to just, when somebody is asking a question, please I know I know you're really you know, you want to the person next to you, but you have to be mindful that we are in a classroom okay. So please, just when somebody asked me a question, please let me hear it okay? I really, you're not kids and I don't want to you know shush anybody but please don't make me okay? So yeah, so the question was "if, will we ever have a case where the formal charge is 0 and you have something where the octet is you know it has its maximum number of bonds that it can form?". Right and if that case ever arise, and there are certain exceptions to that right so there are cases where that does happen and one is… boron is an example okay so boron can only form three bonds okay so its octet is never… Not never, its octet is not complete even though its formal charge is zero. Okay so here clearly when the formal charge of carbon is plus one, its octet is also not complete. But in the case of boron, it has no formal charge in BH3 for instance and it's often is not complete okay. So, the, my sort of main idea to you would be just draw out the molecule and just calculator formal charge. Formal charge of each atom should add up to the total charge on that particular molecule or on that ion okay? You just have to wait okay; I just need to go to next question. Okay great, so the other question that is there where is to draw the Lewis structure for this particular formula, so this is H3PO4 okay. Just along the same lines and I write this down on the side. You should also try to practice all of these okay, these are all similar compounds okay. And can somebody tell me why there are similar bounds? They're all assets okay, and they're also very specific kind of assets will talk a little bit about this. Does anybody know why they are specific kinds of assets? Well, H2CO3 is not such a strong answer but the others the others yes are strong assets. There's something else, there's something to do with the way we are bonded, which makes them unique to themselves. Yes. Polyatomic ions, yes. Something else, something else about the acidity. Anybody? Something about the way the bonds are there. Yes? Oxy acids. The oxyacid, what that really means is that is right so that the phosphorus in this case is the central atom, sulfur is a central atom, nitrogen is a central atom, and we look at that. In all of them, there is at least one bond with an oxygen which in turn is attached to the hydrogen okay. And you won't learn about acids and bases in 110 but in 120, we'll spend quite a bit of time if you do end up taking it, talking about these kinds of assets okay. So, I wanted to say that because you, just put it out there, that when you're seeing these… Again, you will be able to draw cases where you might show oxygen-oxygen bond okay. So, you might show something where you have this. In these oxyacids, there is no oxygen- oxygen bond okay so that's not there. Okay so that's one. The second is these oxygen-oxygen bonds are for very specific molecules. We will talk about some of them, but not just yet okay. So, for instance, in H3PO4, what you want to remember is phosphorus where is going to be a central atom so usually our central atom is group 14,15 or 16 but that phosphorus would be a central atom. And the other thing you also want to keep in mind is that phosphorus is below nitrogen, and it is able to expand its octet. So, nitrogen, you could not, there's no compound which is H3NO4. Okay, even though nitrogen and phosphorus have the same number of valence electrons because nitrogen cannot expand its octet. It can only form enough bonds I guess 4 would be the maximum component okay. But in the case of phosphorus, you have 5 valence electrons so that's 1, 2, 3, 4, 5, okay and there are 4 oxygen so I'm going to show them around this as well. Okay so now, each of these can form one bond with each of the phos… Sorry, each of the oxygen. So, phosphorus can form 3 single bonds with each of those oxygens and it can form 1 double bond with one of the oxygens Okay. So, the reason why I'm giving it a double bond is because the phosphorus, to have a formal charge of 0, should have five bonds and it can expand its octet that stand electrons that is possible with phosphorus. There are still three hydrogens remaining. So, I'm just showing the electrons on this side. So, each of your oxygen has two bonds and two lone pairs. Each of your oxygen, whether it's this double bonded oxygen, this oxygen, this oxygen, or this oxygen. The formal charge on every atom is zero okay. So, when you are trying to draw these Lewis structures, your aim is going to be if the overall molecule has no formal charge, your aim is also going to be to show that each individual atom has a formal charge of zero okay. So, so if you have a phosphorus for instance, it can form up to five bonds. You could also have a phosphorus with three bonds and a lone pair that still has zero formal charge. But here, that doesn't really help us right. We need, there needs to be a bond between that phosphorus and oxygen, and phosphorus can donate its two electrons to form that double bond with the oxygen. Okay so that's that's just just so you know in this case. Yeah. So, in the case of nitrogen, the this is 8 electrons. If you want to form a double bond with this lone pair, that's 10 electrons, and nitrogen because it's in period two, it only has 2S and 2P orbitals. In case of phosphorus, you have 3S, 3P, 3D even though we start to fill 3D in the next period, 3D N is equal to 3 means three orders are available so it can bond using those orbitals. Yeah? Yes. Okay so the question is "In the case of this this molecule, does the phosphorus have to be bonded to each to every oxygen?". So, this molecule is phosphoric acid, we know what the molecular formula is, we know what the structure is, and indeed all four are bonded right. I I know this, I know there is another possibility where you might draw a POO, is that the idea? Like the action they're wanted to each other? How would you know? So, in an exam, if I don't give you more indication but you give me a structure where you've accounted for all the valence electrons and you've accounted for you know follow charge being zero in each atom, that would be acceptable but but just keep in mind it's very… I guess you've given the, you bonded the phosphorus to a hydrogen then, is that right? But then you wouldn't get, the phosphorus should have five bonds to have form of charge of 0. TO ox… Okay, got you, yeah. Okay, so usually, what you'll see an example, question like this, I'll also add, there are no O-O bonds in this molecule. So that will give you an indication that the phosphorus would need to be bonded to those, each of those oxygens. Okay, great so just along the same lines, along exactly the same lines, I would say, and I would recommend I think one of the practice problems is a rview problem as well but try to draw the structure for H2SO4, H2CO3 and HO3 okay? So, this is again, you're trying to build up on what you've learned. Yes. Anything in period three. You really are looking at group 14, 15 and 16. Yeah, so those can expand that okay. So just along the same line, I'll do one more practice because it's directly linked to this. Okay so if you get an anion, kind… The example here is a phosphate anion PO43-, so really it, the phosphorus is a central atom. You'll again do the same thing. Just keep in mind that three negative means you have to add three extra electrons when you're counting the valence electrons, that's what the charge is there for. And again, one of these will form a double bond with one of the oxygen. Okay so this is all the electrons of PO4 okay but what you see here is the octets of the oxygens are not complete. So that's why your 3 minus comes in, each of them gets one more electron in this particular case okay. Now, the next step is "calculate the formal charge on each atom", okay. This is 0, this is 0, this is -1, – 1, – 1. That's what results in your 3 minus overall charge on a phosphate: PO4 3- ion okay. I'm going to leave this up here. Questions? Yes. You should be able to look at any atom on the periodic table and calculate how many valence electrons, yeah? Okay, you have you on Friday, so I'll see you then and if you have questions, I'll be in the front of the class.