Introductory Organic Chemistry
Dr. Neeraja Dashaputre
Department of Chemistry
Indian Institute of Science Education and Research, Pune

Lecture - 04
Tutorial - 01

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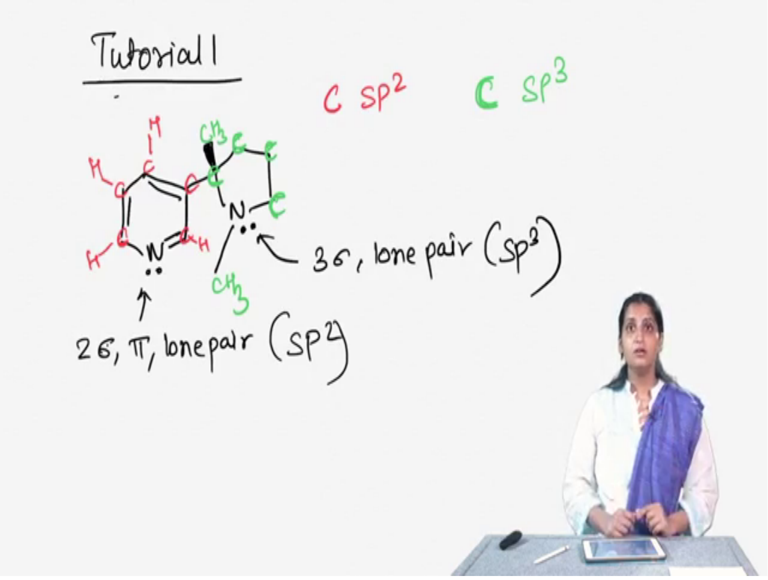

Welcome to the first tutorial of the course. If you have really seen the questions and try to solve them, you might actually see them as a revision of what we have learnt so far or even revision of what you have learnt until now; until the organic chemistry course which is the very basics of molecular bonding, the very basics of atomic orbitals and structures of different molecules. But my understanding is that our foundation should be strong so that we can build up really good as we go ahead and look in the actual course material.

So, for now let us really get quick and really comfortable with such questions. The first question here talks about identifying the hybridization of each atom in the molecule, it also talks about identifying the geometry around each atom and it asks you to identify in which orbitals are the lone pair situated, right?

So, now let's start with the first molecule that is Nicotine here. And what we have here is; let's start with the Carbon. So, all of these sp^2 hybridized Carbons, why I am saying sp^2 hybridized, because they are double bonded, right, they are forming three sigma bonds and pi bond. Now, the third sigma bond might not be really visible, but that's right there; we do not draw Hydrogens as we already saw, right. So, it is forming 3 sigma and a pi bond and that is sp2 hybridized right; so all of these carbons are sp2 hybridized.

Are there any sp3 hybridized carbons that are forming 4 sigma bonds? The answer is yes. I have these 3 carbons here this fourth one, and then this methyl group. So, all of these carbons here, the green ones, are forming, are sp^3 hybridization because they are forming 4 sigma bonds, right.

Now, the question arises which other element here is actually sp^3 hybridized. So, now there are two Nitrogens in the molecule; one of the Nitrogen's here, is forming 2 sigma bonds, has a pi bond and has a lone pair, whereas this particular Nitrogen here is forming 3 sigma bonds and a lone pair. As you can see this particular Nitrogen, the other Nitrogen, which is forming 3 sigma bonds and a lone pair, the situation, is very similar to that of having 4 sigma bonds.

So, you can see that it is sp^3 hybridized, but do not always fall for the trick. Every time you make 3 sigma bonds and a lone pair, the Nitrogen need not be just sp^3 hybridization. We can look at some of the examples, in the next one I am gonna go over an example, where that is not the case. But right now if you really see this Nitrogen; that lone pair is kind of locked on that Nitrogen, it cannot move around because it is bonded to another two sp^3 hybridized atoms. When that is the case, you can be assured that the Nitrogen is sp^3 hybridized. So, this one here is sp^3 hybridized, where as the one that is double bonded is actually sp^2 hybridized, okay.

So, now we have kind of jotted down the hybridization of each and every atom in the molecule. Now, the second part of the question talks about, identifying the geometries around each atom. And sp^2 hybridized centers as we know are Trigonal Planar; they have 120 degrees bond angle and they have the pi bond going up and down, that is, above and below the plane of these three atoms, right; whereas sp3 hybridization is Tetrahedral in nature; 109.5 degrees bond angle, so we are kind of sorted on the geometry, right.

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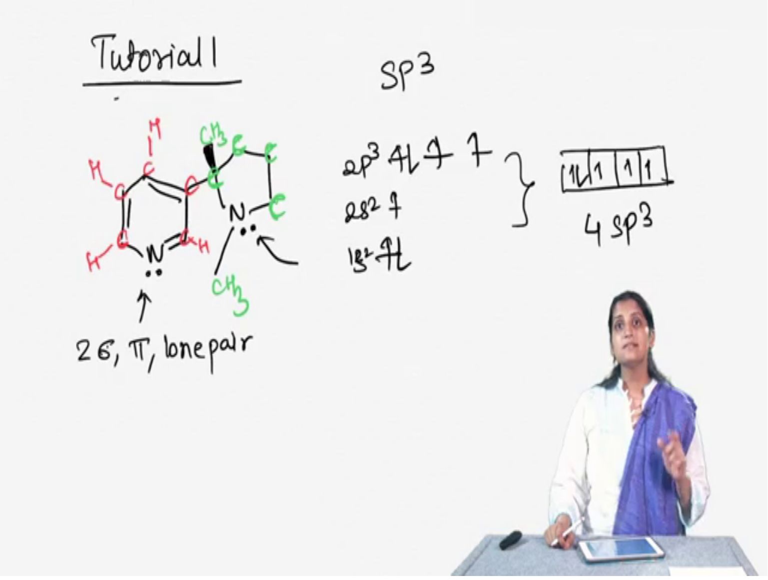

Now, let us talk about the orbital in which the lone pair is residing, right. So, let us look at the sp^3 hybridized Nitrogen, now its ground state electronic configuration is $1s^2$, $2s^2$, $2p^3$ right. In order to show that it does sp^3 hybridization; remove one of these electrons up, right, so it goes to the excited state electronic configuration. And then we combined these three p orbitals and one of the s orbitals to really form 4 sp^3 hybridized orbitals right; so that's what the theory says.

So, in this case for this sp^3 hybridized Nitrogen, the lone pair resides in the sp^3 orbital right. That's really easy. Now, let us go to the tough part; now let us talk about the double bonded Nitrogen and let us figure out where does that lone pair reside. So, for that, gonna get rid of this part and we are going to talk about the sp^2 hybridized Nitrogen.

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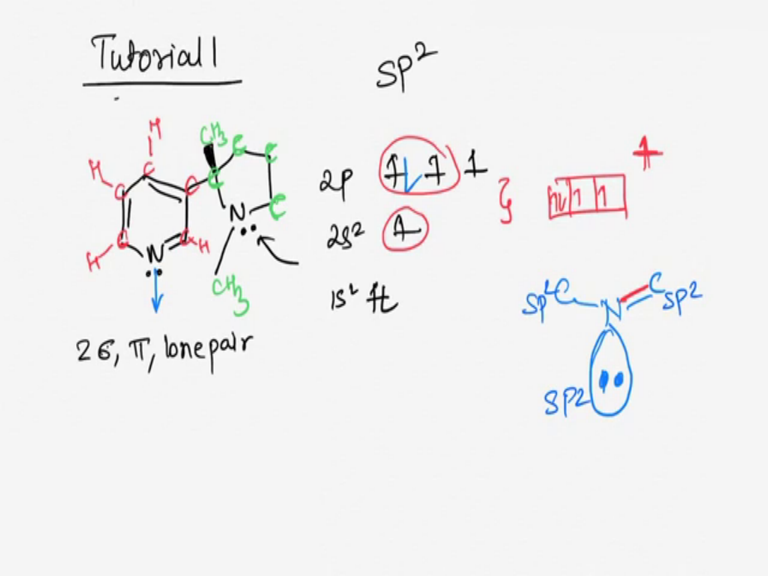

We have $1s^2$, $2s^2$, $2p^3$ that is our ground state electronic configuration; then as we talked about we go to the excited state electronic configuration by moving one of these electrons up. And then in order to form sp^2 hybridized orbitals what you need to do is, you need to combine one of the s and two of the p orbitals together to form three sp^2 hybridized orbitals. If you look at this particular Nitrogen, it is forming 2 sigma bonds. Now, sigma bonds are typically formed when you have a higher s character, because there is a better overlap happening.

So, s character gives rise to sigma bonds. So, when you have these 2 sigma bonds, you can say that the 2 orbitals are sp^2 hybridized. The question lies whether the lone pair is sp2 hybridized or whether the lone pair really resides in the p orbital, which is, which remains unhybridized throughout the process. If you really look at this particular molecule and if you look at the geometry of the molecule, the lone pair is such that it is away from the molecule, it is pointed in the direction that is in the same plane of the molecule right, but it is also going away from the molecule.

So, if this is the molecule here, the lone pair is kind of directed outside of the plane of the molecule. So, the lone pair also really resides in the sp^2 hybridized orbital; okay; right. So, the Nitrogen has three of the sp^2 hybridized orbitals, with two it forms the Nitrogen- Carbon bonds. And the third one is used to occupy the lone pair; whereas the one electron that does not take part in hybridization really forms the pi bond with the other carbon, so that's how it forms that Carbon-Nitrogen overlap to form that pi bond. So, this is sp^2; this is sp^2 hybridized; sp^2 hybridized, right and where as the lone pair is now in sp^2 hybridization and the Carbon-

Nitrogen pi bond takes place because of the electron in the p orbital that does not hybridized from each of the Nitrogen and Carbon.

So, when this molecule has to hybridize it's gonna take one of the s and two of the p's to form that sp^2 hybridization; 1 electron is gonna remain unhybridized, right. So, not necessarily the lone pair is always in the p orbital; that is not always the case. In fact, you have to look at the molecules geometry in order to figure that out.

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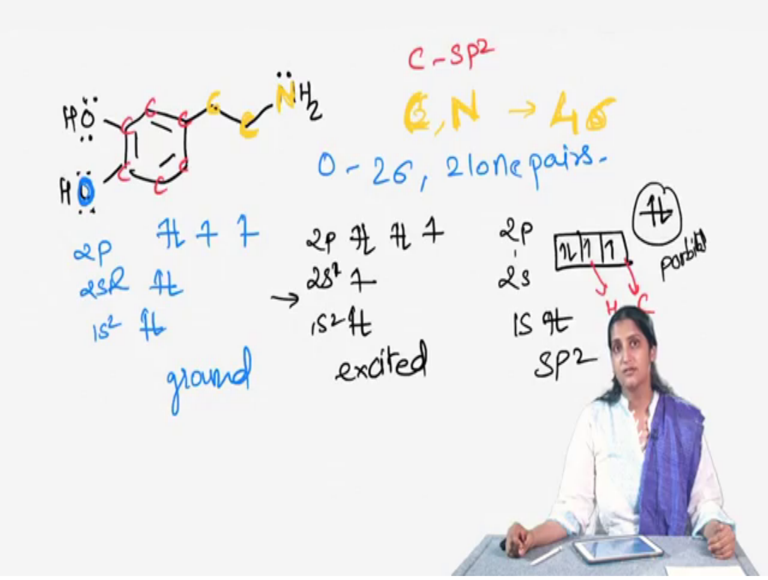

Now, let us look at this particular example, this is Dopamine. Dopamine is the molecule that really is the key molecule that makes us happy, so this is the happiness molecule we have here. And if you look at the hybridization of each atom first let's go over that; all of these carbons which are double bonded and a part of the benzene ring are sp^2 hybridized, right. Bond angles of 120 degrees, they also are forming three sigma bonds and a pi bond; so we are good with that. These other 2 Carbons and this Nitrogen here, these three atoms are sp^3 hybridized right, because they are forming 4 sigma bonds.

So, Carbon and Nitrogen they are forming 4 sigma bonds, Nitrogen again is kind of having either 3 sigma bonds and a lone pair right, so that is very similar to having 4 sigma bonds. Again the Nitrogen's lone pair is kind of locked on that Nitrogen, it cannot give it to the Carbon next to it; so we are safe to assume that it is sp^3 hybridized.

Now, let us look at the Oxygen's here, this is the tricky part. If you look at this particular Oxygen, it may seem like it is forming 2 sigma bonds and has 2 lone pairs, right. So, the Oxygen here has 2 sigma bonds and 2 lone pairs; if you really think about it this is very similar to that of water; water also has 2 sigma bonds on that Oxygen and also has a lone pair, but in the case of water it was sp3 hybridized.

Now, here is the tricky part think about the resonance structure. Is the lone pair locked on that Oxygen? If the answer is yes, then okay, then it is really sp^3 hybridized. But if it is not; if you can move that lone pair from that Oxygen to the next atom and really create a new resonance structure, then that lone pair has to be in an orbital that can be moved to the next atom right, such that let us now look at the hybridization of this Oxygen. We have the ground state configuration as $1s^2$, $2s^2$, $2p^4$, right, that is the ground state electronic configuration.

Now, I will draw the excited state electronic configuration and that is $1s^2$, $2s^1$ right, something like this, okay. Now, the real question arises that when the Oxygen exhibits hybridization, it wants to have that lone pair such that it can put in those electrons into the benzene ring, because we can clearly draw a resonance structure such that, that OH can donate inside the ring.

So, here in this becomes sp^2 hybridized, because that lone pair has to be in the p orbital which is kind of perpendicular to the plane of the ring such that it can easily be donated into the ring. So, you have $1s^2$, now I am going to do 2s and 2p, but I am going to combine one of the s and two of the p's to form the sp^2 hybridized orbitals and one of the lone pair does not take part in hybridization, so it remains in the p orbital, right, the unhybridized p orbital, okay.

When you have these sp^2 hybridized orbitals, one of them will form a bond with Hydrogen, the other one will form a bond with Carbon and one of the lone pairs will actually stay in the sp^2 hybridized orbitals, okay. But in all if you really see the hybridization of Oxygen, it is not like sp^3; even though the Oxygen is forming two sigma and two pi bonds, right; it is not like sp^3, but it is more like sp^2. So, just don't be tricked by just remembering the number of the sigma and pi bonds.

But rather look at whether that particular Oxygen or Nitrogen, can it really form a resonance structure because remember resonance hybrid always exist, right. It is not like Oxygen is sp^3 at a time, then once it donates electrons into the ring, then it becomes sp^2 hybridized that's not the case, because the molecule is gonna exists as a resonance hybrid. So, in order to exist as a resonance hybrid the right way for that Oxygen to exhibit is having a sp^2 hybridization, okay. So, be careful when you are solving questions like this.

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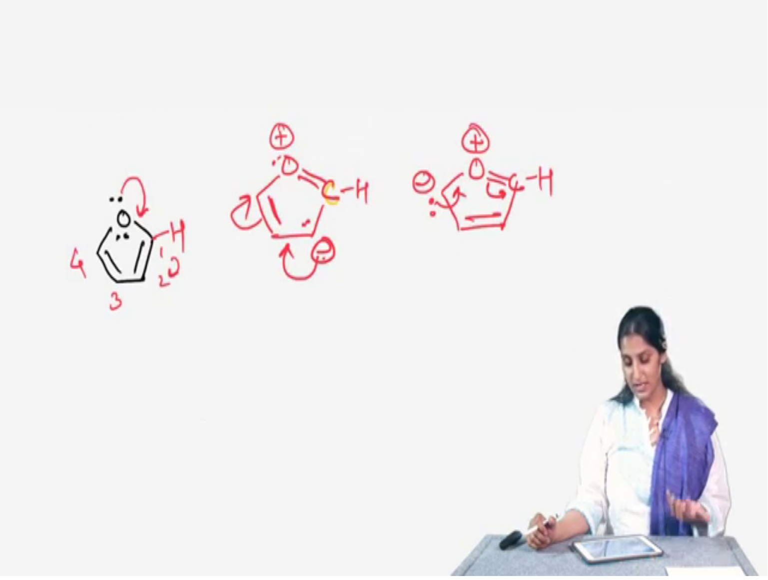

The next question really talks about drawing resonance structures and resonance hybrids. So, now let us draw a couple of molecules and their resonance structures. Again when you are drawing resonance structures remember, that you are allowed to move only the pi electrons or the lone pair or the charges, right, charges also the negative charges, right.

So, in this case I am going to draw a resonance structure of this particular molecule, when Oxygen with its lone pair is gonna give electrons to this particular the first Carbon. But remember that as this Oxygen shifts the lone pair on that Carbon, that particular Carbon is already forming a double bond with the other Carbon and it cannot just accept a pair of electrons; what it does is, it breaks the Carbon-Carbon double bond and gives the electrons to the next Carbon. So, what you have here is a negative charge on this particular Carbon, okay.

We can continue doing this; this particular lone pair can now form a bond between Carbon numbers 2 and 3, right. So if I go 1, 2, 3, 4 the lone pair can now go between Carbon numbers 2 and 3, but again Carbon number 3 is forming, already forming a double bond in order to accept a pair of electrons, it cannot really form more than 4 bonds, so it has to break one of the bonds and the bond it chooses to break is the pi bond.

So, what you have here is this. So, what I have is this particular resonance structure, I can again put this one back here and I can continue doing this kind of rotation, but when we draw resonance structures please remember that in order for the carbon to accept, it cannot just accept it; it has to break a bond in order to accept the new pair of electrons.

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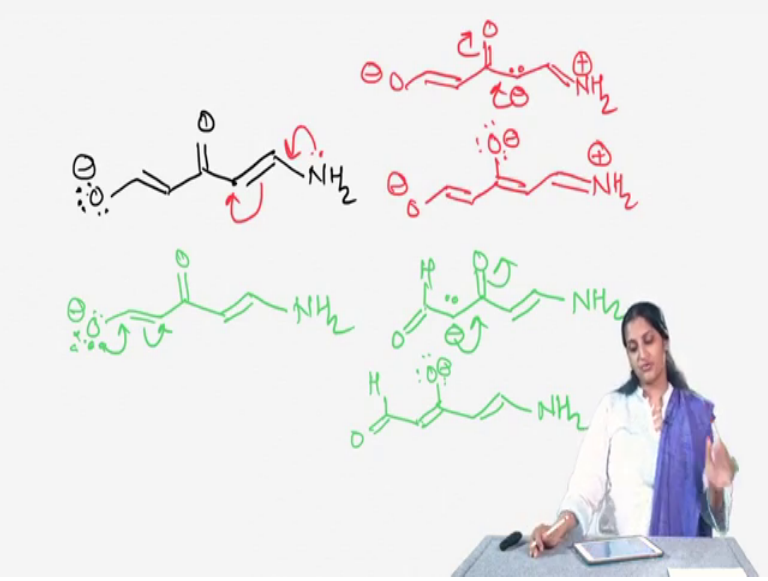

Now, let us do this one example here. In this case, again you have the resonance structures for this particular molecule. Now, the Nitrogen can donate its electrons here and that will make this particular Carbon move its pi bond from that double bond in between, right. This will give rise to this particular structure, right; I can continue doing this; this particular electron pair can go here and that can go up, and you can form this structure.

You can again bring back the lone pair of electrons and go in the reverse direction, so that's when you start from the Nitrogen or I can start from this Oxygen from the other end. So, the Oxygen, let me just redraw the molecule here; the Oxygen here has a pair of electrons that can be moved here and this goes here so, right, that is one of the possibilities.

Again I can draw one more structure such that this can go here and that goes up. So, what I have here is these are all the possible resonance structures for this particular molecule. And when you are drawing resonance hybrid, make sure that you are looking at where the double bonds are whether they are present in all the resonance structures or whether they are presenting only some of the resonance structures, then you have to draw a partial dashed line right or a partial line.

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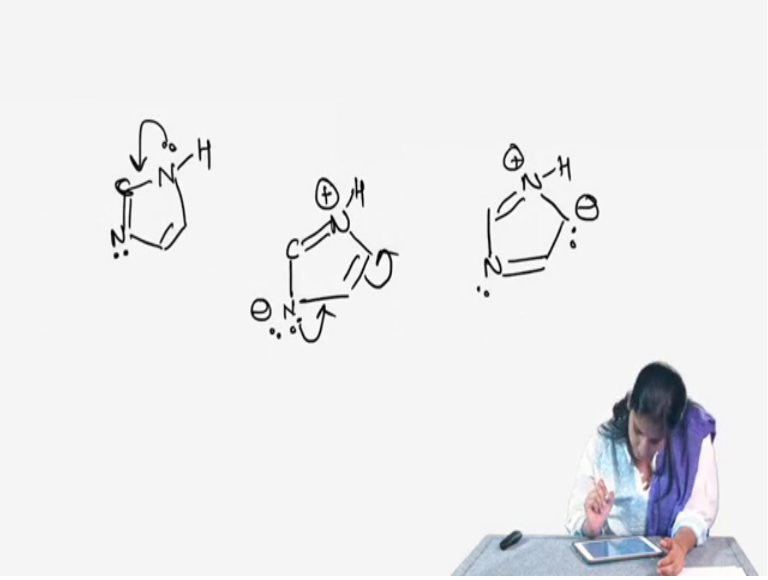

Okay, let us draw this Imidazole molecule here; now in the case of Imidazole remember that I can draw this particular lone pair giving into the ring that will cause this Nitrogen to be positively charged. So, always remember to put those formal charges, it is very important that we do not forget our formal charges, right. As a result of which this Carbon-Nitrogen bond is shifted on this Nitrogen and that is the case. I can continue moving it, I can put it here and this will move the electrons back here, so that will form this particular structure, okay.

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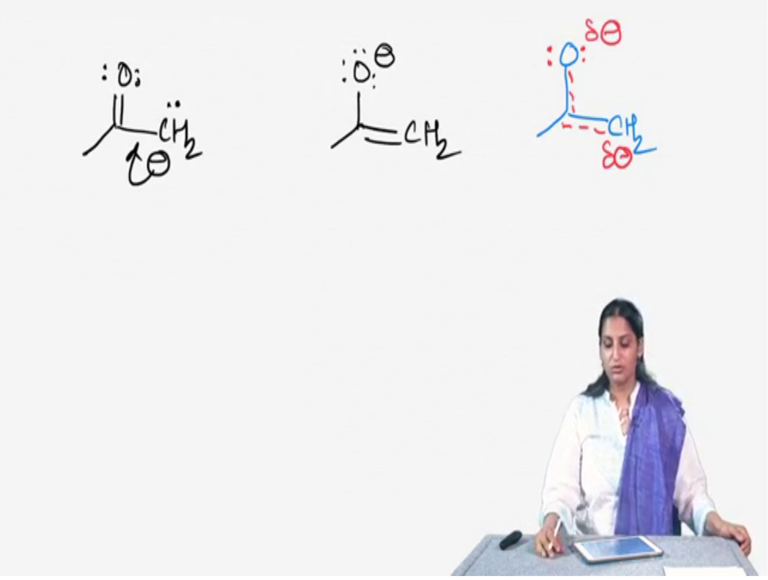

So, you can draw resonance structures of many molecules, here in you have an enolate ion. Enolate again you can draw it such that the Carbon lone pair can be pushed towards the carbonyl group, and you have this particular resonance structure, right. So, all of these are possible and you can go ahead and draw them. Now, let us quickly revise resonance hybrid. So, for example if I want to draw the resonance hybrid of this enolate ion, what I need to do is I need to look for the bonds that are constant that is the bonds that don't change. This is the sigma bond structure that did not change throughout the all the resonance structures.

Now, I will draw dotted lines wherever there are double bonds in only some of the structures and not all the structures, so that is that case. So, there is a partial line between Carbon-Oxygen and the other Carbon-Carbon. You also look at the lone pairs you give them right, and you also give a partial charge, because in one of the cases the Oxygen has a negative charge, but the other structure it does not have. Where as in the same case the first structure has a Carbon that is negatively charged, but the second structure does not have. So, what we are going to do is, we are gonna give delta negative on both of these atoms; that's how my resonance hybrid looks, right.

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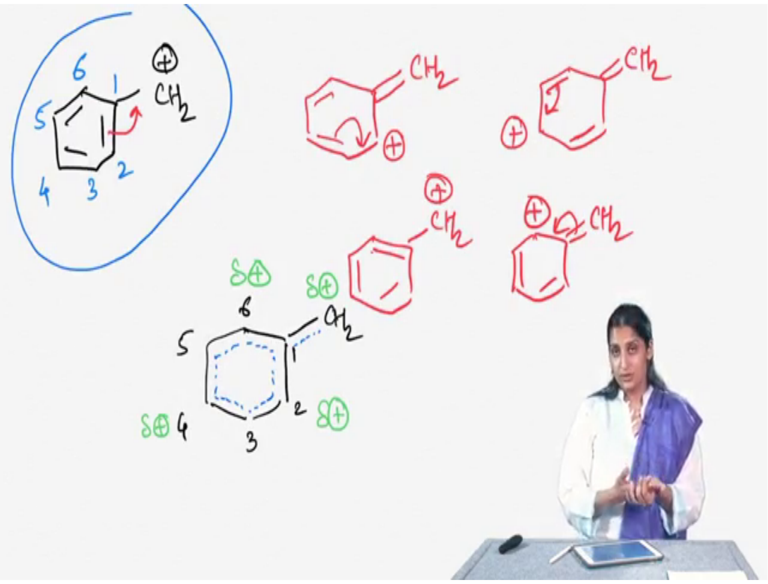

Now, let us draw it for this molecule; this is a Benzene ring, so we have a lot of charges to move around. Now, remember in this particular molecule the Carbon that is bonded to the Benzene ring is positively charged, it is a carbocation. In this case we have to move electrons from the ring back to the Carbon, right.

And what we are going to do is we are going to move these electrons here okay, you just say 1, 2, 3, 4, 5, 6. As the electrons between Carbons 1 and 2 move towards that outer Carbon, what happens is that Carbon number 2 is now short of two electrons, because what happened was, it lost the bond. So, as a result of which it gets the positive charge right. Where as in the next case, when this moves, the Carbon number 3, 4 double bond when it moves to Carbon number 2 and 3 that Carbon, Carbon number 4 is going to get a positive charge.

Always think about which is the Carbon that lost the bond, right. So, when Carbon number 3, 4 moves from that position 2 to 3; 3 is really just shifting one of its bonds right, but the fourth Carbon here, it really loses a bond, so that gets the positive charge. I can go on doing this, I can draw one more structure and now Carbon number 6 will get a positive charge right. So, all of these are possible structures, I can go back to the first structure by pushing electrons this way.

See one thing I always want everyone to remember is that when you are drawing resonance structures, please do not draw more than two lines at a time, or more than two curved arrows at a time, it gets highly confusing. So, it is much better to just do two steps at a time and that way you don’t miss out on a resonance structure. So, especially when we have to draw resonance hybrid, if you miss out on a particular resonance structure, your hybrid is gonna look incomplete. So, it is better to just push two arrows at a time and no more than that.

So, now I want to draw the hybrid of this particular atom; now I want to draw the hybrid of this particular molecule. And then let us now look at 1, 2, 3, 4, 5, 6 right, these are the bonds that do not really change; of course, I have been drawn the Carbon-Hydrogen bonds that are kind of there, but we have not drawn them anywhere. Now, let us look at the pi bonds. So, the pi bonds are there in between 1 and 6; it is there in some of the resonance structures, it is not there in some of the other resonance structures, so it gets a partial line.

Same goes for the pi bond between 1 and 2. So, it is present in this first structure, but it is absent in the second structure, so that gets a partial line. Same way if you see the pi bonds between Carbon and Carbon are kind of such that they get a dotted line that I have represented with this blue color; they get a dotted line between all the structures, so the system is really in conjugation. Meaning the double bonds are kind of getting interchanged in the resonance hybrid such that all of them coexist partially at the same time.

Now, let us look at the positive charges. So, in the first resonance structure the carbon that is outside the benzene ring has a positive charge. Whereas, in the second one the charge moves on the second Carbon; in the third one it moves on the fourth Carbon, and then it moves on the sixth Carbon and that's how, it is moving around, right. So, when I want to give a partial positive charge, I will look at the Carbons that have it. So, in the case here this CH_2 outside will have a delta positive; Carbon number 2 will have a delta positive, Carbon number 4 will have a delta positive and Carbon number 6 will have a delta positive, right; so all of these are the delta positive Carbons that are present, right.

If you see in any of the resonance structures, you do not have a positive charge on the third and the fifth Carbon or the first Carbon. So, in a resonance hybrid also we do not want to put a partial positive on those Carbons. So, 1, 3 and 5 do not have charges, the rest of the molecules will have a partial positive in the resonance hybrid. So, the analogy towards resonance hybrid is that you can imagine that it is something that exists right, but the corresponding contributing resonance structures do not exist as themselves.

So, it is like having a rhinoceros. So, rhinoceros are, have a horn, right and they have a very thick skin. So, you can imagine that rhinos are kind of a combination of dragons which have very thick skin and unicorns which have a horn, right. But both unicorns and dragons are kind of fictionary figures and they really do not exist, but rhino which is a combination of both of them really exists, so that's how you want think of a resonance hybrid is that, that is what it exists as and not as these individual resonance structures, okay.

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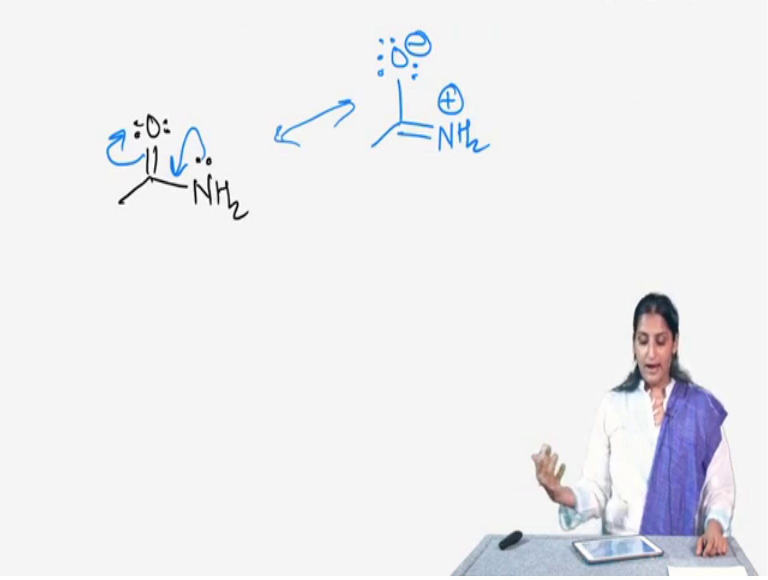

So, let us look at some more examples, here in we have an amide. In this amide, the Nitrogen can put electrons here and we can really have an overlap such that the Carbon and Nitrogen form a double bond and the Oxygen and Carbon form a single bond. So, infact, the amide bonds that exist in nature are such that the Carbon-Nitrogen bonds have much more double bond characters, because the particular resonance structure here is quite stable if you think about it, because the octet of each element is really complete and also the Carbon-Nitrogen overlap is kind of high, right.

So, what you have is that amides also show a good resonance contribution of this particular resonance contributor. And can really show a partial double bond characteristic between the Carbon and Nitrogen.

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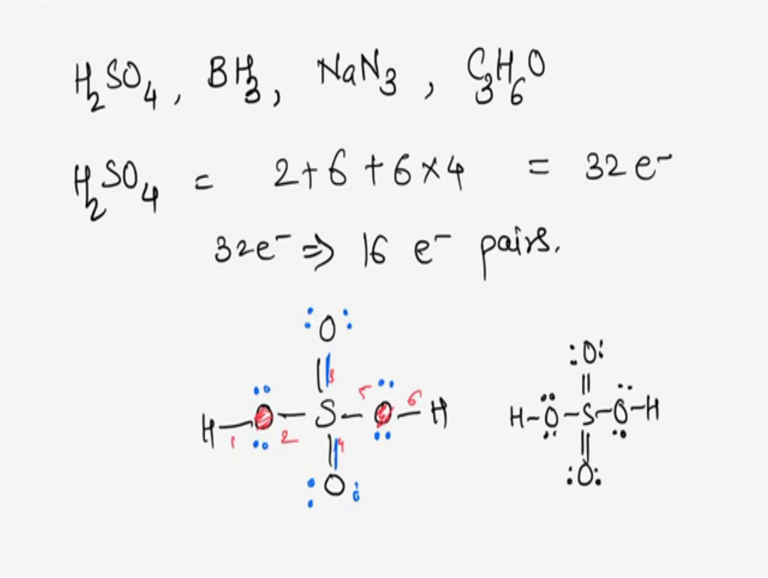

The next question, really talked about drawing the Lewis structures of these molecule. So, what I am gonna do is I am gonna do some of them, so let us do the first one which is H_2SO_4, which is a common mineral acid right. So, when we look at H_2SO_4, what do we have, first step while drawing Lewis structures is that we count the total number of electrons. So, the total number of valence electrons for this particular molecule is 2 for the Hydrogen's, 6 for the Sulphur and 6 for the Oxygen; but I have 4 of them, so 6 into 4 that is 24; this will give me a total of 32 electrons, right. 32 electrons meaning, I have 16 electron pairs, right.

So, now when I draw the structure, I have to fill in 16 electron pairs into the molecule. Again first things first, we have gotten to the number of electrons, now the next thing we want to do is we want to place a skeleton of the molecules such that the atom that can form the maximum number of bonds is at the centre of the skeleton. So, we want to put Sulphur in the middle, because Sulphur can really expand its octet.

And then I wanted to bond with the 2 Oxygen or rather 4 of the Oxygen's. And 2 of these other Oxygen's I want to bond with the Hydrogen's, right. So that is kind of the skeleton I want. The reason why we chose Hydrogen's bonding to the Oxygen, because I know that H_2SO_4 is an acid and you need a proton; so, right; so it needs to be attached to the Oxygen such that it can be easily removed in an acid-base reaction, but that is even if you didn't have that you could still draw one more structure and still try to come up with a possible scenario; but this is the one that I am gonna work with now; this is our skeleton structure.

Now, if you think about it, how many electron pairs have we attached? So, we have 1, 2, 3, 4, 5, 6; we have used so far. We have to use 10 more electron pairs. Both the Hydrogen's have satisfied their valencies; if you really think about the 2 Oxygen's that are here and here, which are attached to these Hydrogen's, have also satisfied their valencies; they are forming two bonds each. Then, these other two Oxygen we really need to fill in their valencies. So, what I am going to do is I am gonna give the two electron pairs here to form that bond, okay.

Now is it such that the octet of each atom is complete? The answer is still no, because we are yet to place the lone pairs. So, right now what I have used is 8 electron pairs, now I need to put two more. So, I am gonna put them here, here, here, and here. This really satisfies the valency and the octet rule of all elements. So, what you have here is the structure of H_2SO_4 becomes this, right. So, we have really taken care of the molecules such that we have used all 16 electron pairs, and we have also completed the octet for each atom in the molecules, that's the structure of H_2SO_4.

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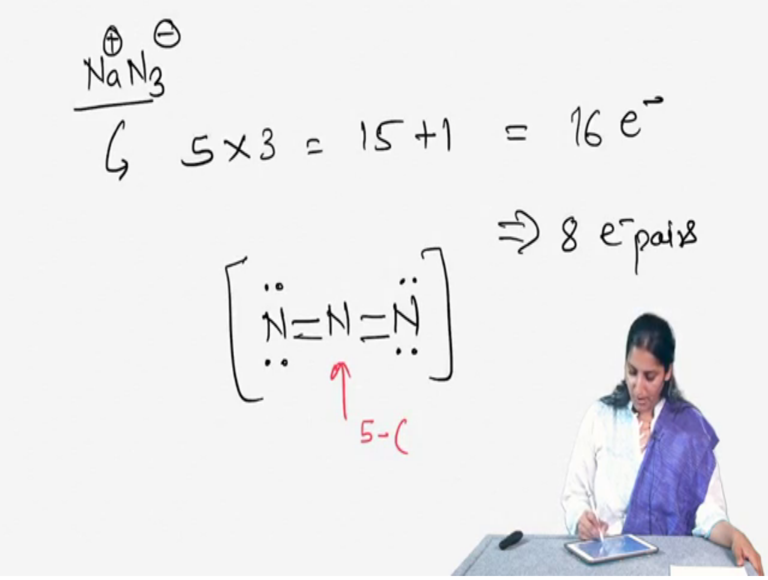

Now, let us do one more compound that is NaN_3. So, Sodium azide is also used in bio labs a lot. So, Sodium azide let us now look at the Lewis structure. So, whenever we have an ionic compound, we first give it the charges, so it will be Na^+ and N_3^- or azide ion, right. So, what we are going to do now is we are going to look at the azide ion and figure out the structure of that. Nitrogen has 5 electrons in the outermost shell or the valence shell, so the total number of valence electrons is 5 into 3 that is 15, right. But remember that negative charge that needs to be added for one more electron, so then you have 16 electrons present that I need to take care of, 16 electrons that will give rise to 8 electrons pairs, okay.

Now, when I want to draw the azide ion, I can look at it such that the 3 Nitrogen's are kind of, there is no other atoms so it's very easy to come up with the skeleton structure; it is the 3 Nitrogen's; but that really takes care of just 2 electron pairs. So, now let us put two more, so that is four electron pairs. Now, let us put fifth one, sixth one, seventh, and eighth. So, I can draw the skeletal structure like this. So, we have taken care of all the eight electron pairs, but if you really think about it what about formal charges; right. Remember we always want to think about formal charges.

The middle Nitrogen, it is forming four sigma bonds, right. So, if you really think about it, its valency is 5 minus the number of electrons that it is sharing, right. So, how many electrons really belong to the Nitrogen? Only four of them; so, we have it as 5 minus 4, right. So, the formal charge on this particular Nitrogen will be plus 1; the middle nitrogen will have a positive charge.

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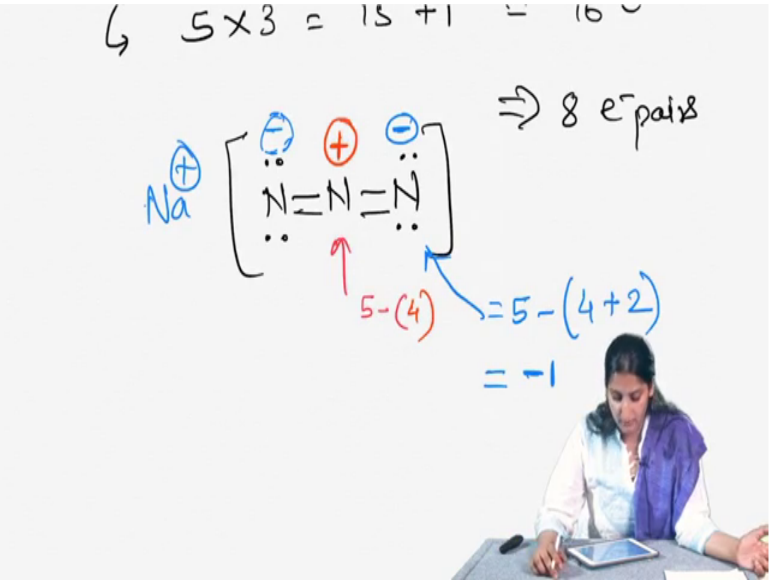

Now, if you look at this other Nitrogen here, five which is the number of valence electrons of the Nitrogen minus, how many electrons are in the lone pair, 4, plus the number of electrons that are in the bonding pairs divided by 2, so that will be 2, so that will be total of minus 1. So, the other Nitrogen's actually have a charge of minus 1 each, right. So, what you have here in fact is that the sodium azide has this particular Lewis structure okay; whereas both the

Nitrogen’s on the outside of the azide group are negatively charged; the internal Nitrogen is positively charged, and together if you really think about the molecule you have an ionic bond between Sodium and Azide.

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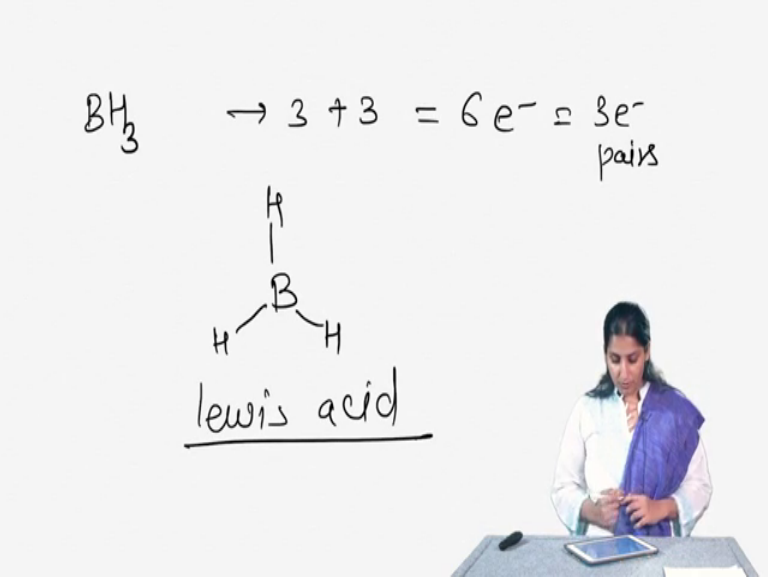

Let us take one more example, which is of BH_3. So, this is really really simple right because Boron is going to have 3 electrons in the outermost orbital, and then the 3 from the Hydrogen, so total it gets to 6 electrons that is 3 electron pairs, right. So, now, let us put the central atom as Boron, it is going to form 3 bonds with the 3 Hydrogens. So, everything looks good so far. And if you think about it, we are done with using our three electron pairs, but think about whether each and every atom in this molecule has the octet complete, the answer is no. Boron here has only 6 electrons around it right, so 6 from the 3 B-H bonds.

Boron here is electron deficient. So, when this particular molecule reacts in chemical reactions, what you really see is that the BH_3 is such that the Boron is kind of ready to accept electrons from other molecules, and Boron here really acts as a Lewis acid. So, this is also something you must have learnt before, Lewis acids and Lewis bases. So, Lewis acids are the molecule that accepts a pair of electrons in a chemical reaction, and such that they have the vacancy for that particular electron pair to be accepted. Boron here can accept a pair of electrons.

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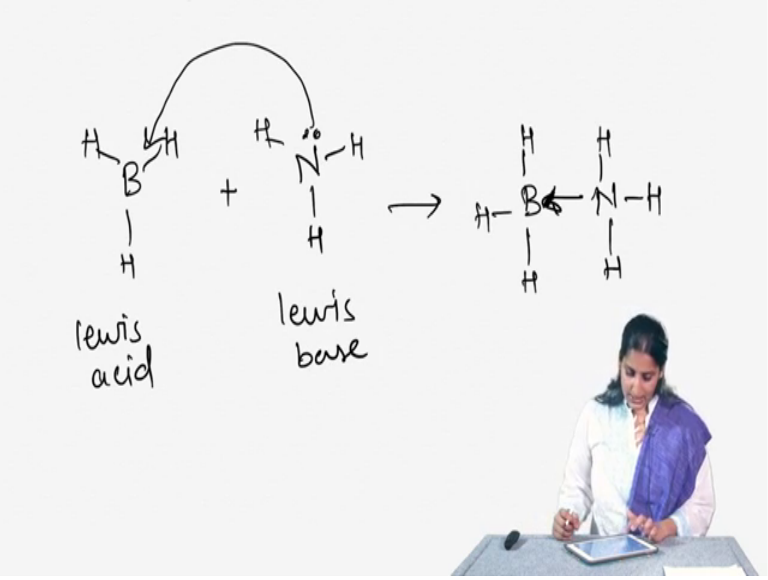

And in fact let us write down a reaction between BH_3 and NH_3. So, BH_3 which is our Lewis acid, and Ammonia which is our Lewis base, right. So, Lewis base here, the Ammonia is gonna put its electrons towards the Boron because there is a vacancy; Boron really wants 8 electrons around it. So, what gives rise to is a coordination bond between Nitrogen and Boron, because both the electrons are really coming from Nitrogen and Boron is not contributing any electrons towards the bond.

So, what you now have is, is this particular compound, which is a Lewis acid-Lewis base complex that is formed, okay. We are also gonna look over Lewis acids and bases when we go over the chapter of Acids and bases, but I thought it would be good to include an example for a Lewis structure, wherein the central atom is actually short of the required number of electrons for the octet, okay.

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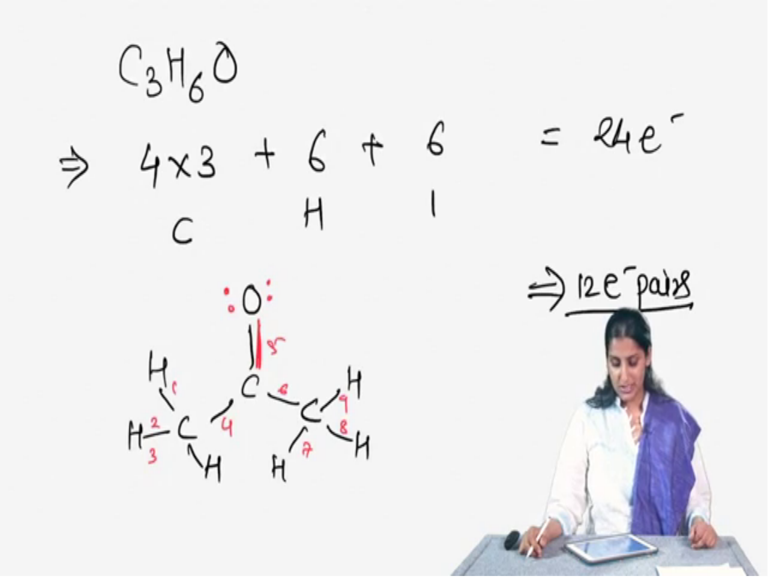

Let us take the last example, which is C_3H_6O, okay. Whenever we have a compound like this, let us first figure out the number of electron pairs that we have. In order to do that we first calculate the number of valence electrons that are present in the system. So, for Carbon, we have four electrons, and I have 3 of these Carbon plus Hydrogen; we have 6 Hydrogen's and Oxygen, there is one Oxygen with 6 electrons. So, what do you have here is 12 plus 6 plus 6 that is 24 electrons, right. Now, if you calculate 24 electrons, what you have is 12 electron pairs, okay. So, we have 12 electron pairs around it.

Now, if you think about it, let us draw a skeleton such that the 3 carbons are kind of in one line, okay. And you can draw the 6 Hydrogen such that they are forming a bond here; the middle Carbon is forming a bond with the Oxygen. If you really see this skeleton structure, we have taken care of 1, 2, 3, 4, 5, 6, 7, 8, 9 electron pairs, right. So, if you really see we have taken care of the 9 electron pairs and we have 3 more electron pairs to put and those 3 electron pairs can actually go as a double bond between the Oxygen-Carbon and the 2 lone pairs on that Oxygen. This really takes care of the octet rule for each element; this also fulfills our 12 electron pairs. But remember this is not the only structure that is possible.

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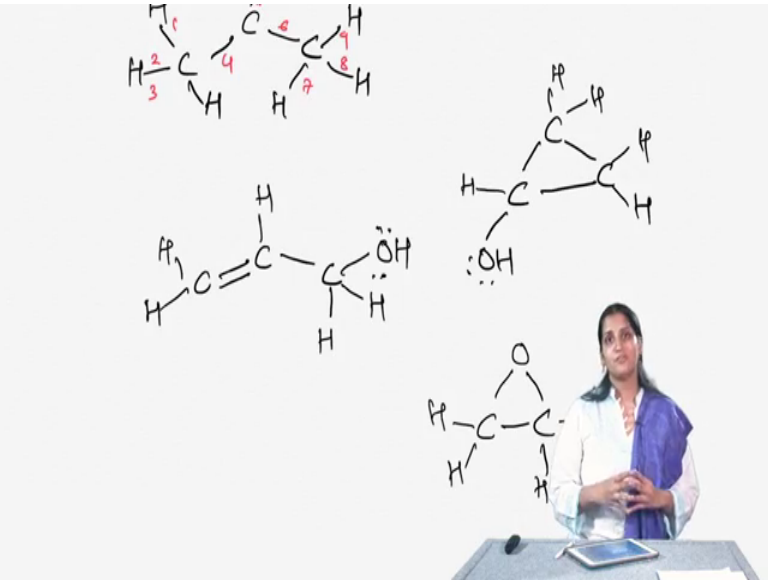

So, let us now draw one more structure, which will really satisfy the rules. So, now I can put electrons between the two, this is also a valid Lewis structure of the same, because I still have every atom completing an octet, and I still have the right number of valency for being fulfilled for all atoms; this is a valid structure. I will tell you one more. So, now, I can have, right, this also is a valid structure right. I can go on writing more structure, so I can even draw a Lewis structure for the same molecule wherein the Oxygen forms the ring.

So, there are so many possible ways in which you can really draw this particular molecule and its structure. The reason is that, why I want to go over all of these possibilities is that when you are looking at Lewis structures, you can arrive to a solution that seems very logical that fits all the criteria's and that is in fact it could be one of the right answers. All you need to check in order to identify if your answer is right or wrong is that are all the electron pairs satisfied; are all the atoms in the compound have fulfilling octet rule complete and also have you paid attention to the formal charges that may be there in the molecule. As long as you have done that you can really arrive at a number of possible solutions and each one of them would be right.

Now, in order to figure out which one is really the right molecule amongst all of these, I may have to do some chemical tests. So, for example, if it's ketone, like acetone that is shown there it will do the reactions of ketones or if it is an alcohol that we drew in the second time, it will do the reactions of alcohols; if it is an epoxide, it will do the reactions of epoxide. So, just based on the molecular formula you can still arrive to a number of possible outcomes and you will have to do chemical tests to identify which one is the right answer, okay. So, that was it, for the first week.

We will meet next week with our very first organic chemistry chapter that is the Chemistry of Alkanes.