Chem1311Ch10Ep3 Transcript

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Hello and welcome to the third episode of molecular geometry and orbital hybridization.

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Previously in molecular geometry and orbital hybridization.

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We practiced determining a molecule geometry from its Lewis dot structure.

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And we learned how the bond dipoles and molecular geometry interact to give rise to the molecule's resultant dipole moment.

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In today's episode.

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We will consider hybridization of orbitals in the central atom.

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We will consider the different possible types of hybridization in bond formation.

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We will determine the type of hybridization taking place in a molecule central atom based on the Lewis dot structure.

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Is the mixing of two or more atomic orbitals.

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To form a new set of hybrid orbitals.

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The salient points of atomic orbital hybridization are.

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That it results from mixing 2 or more non-equivalent orbitals.

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The hybrid orbitals will have very different shapes from that of the original orbitals.

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The total number of hybrid orbitals will be equal to the number of the original orbitals.

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And covalent bonds are formed by the overlap of two hybrid orbitals or a hybrid orbital and an unhybridized orbital.

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"SP³" hybrid orbitals are the result of mixing 1one "s" and three "p" outer shell orbitals to form 4 "sp³" hybrid orbitals.

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This is the type of hybridization that a carbon atom undergoes to form the four bonds with hydrogen atoms in methane.

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This is also the type of hybridization that nitrogen undergoes when forming the ammonia molecule.

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Notice that in.

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Both cases, four hybrid orbitals were formed.

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A different type of hybridization is what we see when One "s" and one "p" orbital are mixed, resulting in two "sp" orbitals.

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As you have probably noticed by now, the name given to the hybrid orbital is derived from the original orbitals that went into the hybridization process.

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Beryllium undergoes SP hybridization to form the two orbitals when it reacts with chlorine to form beryllium chloride.

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"sp²" hybrid orbitals are the result of the mixing of one "s" and two "p" orbitals.

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This is the type of hybridization that boron undergoes when it forms boron trifluoride.

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How to predict the hybridization of a molecule's.

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Central atom?

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It turns out that if you know what the Lewis structure is, then it's just a matter of counting the number.

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Of electron pairs in that central atom.

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Treating bonding pairs and electron pairs the same way, the total will give you the number of hybridized orbitals.

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And the number of hybridized orbitals is the key to the type of hybridization.

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Here are three examples of molecules.

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From their Lewis dot structures, determine both the hybridization of the central atom and the molecular shape.

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Pause the video.

00:07:00

Write down your answers, then come right back.

00:07:14

Welcome back.

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The Lewis structure of the first molecule.

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Has 2 bonding pair of electrons and no lone pairs.

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This will result in a linear molecule.

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And because there are only two pair of electrons on the central atom, the orbitals in beryllium are "sp" hybrid orbitals.

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This is the Lewis dot structure of the second molecule.

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The central aluminum atom has three bonding electron pairs and no lone pairs.

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Therefore, the molecule will have a trigonal planar geometry.

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And the orbitals in the aluminum atom are "sp²" hybridized.

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This is the Lewis dot structure of the third molecule.

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The central phosphorus atom has four electron pair around it, giving it a tetrahedral arrangement.

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And because one of them is a lone pair, the molecule will have a trigonal pyramidal geometry.

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The orbitals in phosphorus are "sp³" hybridized.

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Because there are four total.

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Here's another fun molecule.

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Pause the video

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Write down your answers and come right back.

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Welcome back.

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The Lewis dot structure of phosphorus pentabromide shows 5 electron pair around the central atom.

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Because there are no lone pair, the shape.

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Of the molecule is trigonal bipyramidal.

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And the orbitals on the phosphorus are "sp³d" hybridized.

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Now carbon.

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Carbon is a great example of what happens when double and triple bonds exist.

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Carbon is capable of forming more compounds than any other element.

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It is very talented that way.

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The presence of double and triple bonds affects the type of hybridization.

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Because a double bond requires an unhybridized P orbital.

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And a triple bond requires 2.

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In each of the atoms involved.

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I believe this is what is called a plot twist.

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If a double bond is present, an unhybridized P orbital perpendicular to the "sp²" hybrid orbitals must exist.

00:12:06 The way it goes down. 00:12:08 Is like this. 00:12:16 Carbon has one empty P orbital when it is in its ground state. 00:12:25 In order to form bonds. 00:12:28 One of its 2s electrons is promoted or moved to occupy that empty orbital. 00:12:39 Then the "s" and two of the "p" orbitals hybridize to form three "sp²" orbitals. 00:12:48 And the third "p" orbital remains unhybridized. 00:12:59 This is the orbital which will form the π bond. 00:13:10 Ethylene has a double bond between its two carbon atoms. 00:13:18 This is the Lewis dot structure. 00:13:21 Notice that the number of bonds indicate trigonal planar shape and "sp²" hybridization. 00:13:31 Whenever you see a double bond. 00:13:35 "sp²" hybridization is a must. 00:13:38 Just because an unhybridized "p" orbital is required, not suggested, required. 00:13:55

One of the "sp²" hybrid orbitals forms the bond with the other carbon.

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The "p" orbital is perpendicular to the hybridized orbitals as picture to the right.

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The overlap of the "sp²" orbitals between the carbon and the "sp²" and S orbitals between the carbon and the hydrogens are called sigma bonds. Here drawn in green.

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The overlap of the orbitals between the carbon atoms.

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That's the unhybridized.

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2p orbitals.

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Is called a π bond.

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 Π bonds have two lobes and therefore made the connected atom unable to rotate.

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Sigma bonds are able to rotate.

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Single bonds are Sigma bonds.

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Double and triple bonds are made of 1 Sigma and one or two Pi bonds respectively.

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As I mentioned earlier, atoms can rotate freely around Sigma bonds, but not around Pi bonds.

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Pi bonds freeze the molecule in place.

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The process of SP hybridization of a carbon atom is precisely similar to the SP2 hybridization.

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It begins with the promotion of an electron from the 2s to the empty 2p orbital.

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And it's followed by the hybridization of the "s" and one of the "p" orbitals.

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The remaining two "p" orbitals will be perpendicular to each other and to the hybrid orbitals.

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Orbitals, yes.

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Acetylene can be used as an example of this type of hybridization.

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This is the Lewis structure of Acetylene.

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The triple bond will consist of two Pi bonds formed by the overlap of the unhybridized "p" orbitals in the carbon atoms. Besides the Sigma bond.

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As you can see here, the Acetylene molecule is linear.

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Just that VSEPR would predict.

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Here we have.

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Another lovely molecule for us to describe.

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I can't believe they seriously told us that.

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We could have figured that out.

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Notice that the oxygen has three Sigma bonds and one Pi bond, so it is impossible for it to be anything other than "sp2" hybridized.

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See those right there?

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And of course, the carbon will also be.

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"sp2" hybridized because it has three Sigma bonds and one Pi bond and the Pi bonds require.

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Unhybridized "p" orbitals.

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The shape of the molecule.

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Will be.

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Trigonal planar because the carbon atom, which is the central atom, has three.

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Bonds attached to it.

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The carbon and oxygen will both show "sp²" hybridization leaving one unhybridized "p" orbital in each.

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The "sp²" orbitals in carbon form 1 Sigma bond to the oxygen and one Sigma bond to each hydrogen.

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The Unhybridized 2p orbital forms a π bond to the oxygen.

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One of the "sp²" orbitals in oxygen forms a Sigma bond to carbon.

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And the other two hold 2 lone pairs on the oxygen atom.

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The unhybridized "p" orbital forms the π bond to the.

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To the carbon atom.

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Delocalized molecular orbitals extend beyond the two adjacent atoms, and are the result of Pi bonds in close proximity.

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Benzene is the best-known example of delocalized molecular orbitals.

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Benzene is drawn as two resident structures which alternating single and double bonds.

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Here are the Sigma bonds of each carbon atom showing trigonal planar arrangement for every atom in benzene.

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Here are the delocalized "p"orbitals of the molecule.

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These delocalized molecule or orbitals give benzene a great deal of stability and are the main reason why benzene lasts and lasts.

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That is not a great thing if we consider it is a carcinogen.

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It's one of those so-called "forever molecules."

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And that's all there is.

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There isn't any more.