

## MITOCW | MIT3\_091F18\_lec13\_300k

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But back to MO Theory, which is where what we talked about last Friday.

And there are sort of two last things I wanted to mention now related to MO Theory before moving on to the topic of hybridization.

Now, before I do that, you know, I'm going to use a couple of videos today and I'm-- I'm not going to show you this one, but I just want to mention there's lots of videos that help, you know, animate this stuff, right?

So like this is what-- so like-- Oh, OK.

[VIDEO PLAYBACK]

Bonding molecular orbital.

[END PLAYBACK]

Antibonding, we talked about this last Friday.

But look at this animation.

I couldn't do that on the board.

Watch this.

[VIDEO PLAYBACK]

Antibonding molecular orbital formed by the combination of two 1s orbitals are called sigma 1s and sigma star 1s respectively.

[END PLAYBACK]

Kind of remembering stuff now, right, from Friday?

[VIDEO PLAYBACK]

The bonding molecular-- Oh, look at that.

-- is lower in energy than the original [INAUDIBLE]..

[END PLAYBACK]

That's that kind of animation that you just can't-- you know, I can't animate the bonding and antibonding.

Well, I kind of can with my hands, but there it is.

And there's lots and lots of those kinds of videos.

I'll show you a couple today related to hybridization, because it really does help see these 3D orbitals, right, in their full shape and what's happening.

So I highly recommend it, especially if you're having some conceptual questions related to stuff we talked about Friday and today.

Take a look at some of those videos.

Now, where we left off on Friday, before I get to nitrous nitric oxide, which has all those different uses in the world, and many more.

We left off with the homonuclear dimers.

That's what we were doing on Friday, right?

How do the atomic orbitals come together to make orbitals that are molecular MOs, right?

And we did it for dimers that have the same atom, right?

And so here-- so we did nitrogen and 2 and we did O<sub>2</sub>, right?

And so let's do that on the board and that-- because now we're going to start with one that has N and O, heteronuclear dimer, right?

Two different atoms.

But if I had the same one, remember, this is what it would look like, right?

So here-- here would be, say, N<sub>2</sub>.

Would-- you know, I'm going to just stick with the valence.

So the 2s for our nitrogen atom, and the 2p would have, remember, Hund's rule.

All right.

OK.

So that's what the 2p would look like.

And then over here, we would have the other atom.

This is how we roll when we talk about MOs.

And we'd have the 2s over here.

And this is how we start.

These are my AOs for the 2 nitrogen atoms.

Right.

This is what we did, we're just getting back in the mood because we had the long weekend.

And when we make the MOs, remember, we have all those sort of things that we talked about in terms of how these-- this difference between the bonding and the antibonding orbital, right, is related to the overlap in the same symmetry of the two AOs that come together.

[INAUDIBLE] we talked about that, the more overlap, the biggest this delta.

Gesundheit.

All right.

And then we also-- so that would be filled, this would be filled, right?

Sigma 2s, sigma a.

Sigma star 2s, right?

And then for nitrogen, we also talked about how-- oh, boy, these are reversed, right?

See, I'm now already...

...backwards twice, which is not what I want.

OK.

And this was-- but this is the weird thing about nitrogen, because if we go to oxygen, then the way it goes is, you know, you've got your 2s in oxygen. OK.

And in oxygen, you've got your 6 valence electrons in the 2p.

And over here, you've also got the same.

OK.

And you've also got that 2s.

And-- but for oxygen-- OK.

So those are the sigma, sigma star.

But for oxygen, we go to the ordering that we usually have, which is that-- remember, those pi orbitals have-- and this isn't necessarily on a scale.

But those pi orbitals-- all right.

So this was sigma, sigma star, and these are the pi orbitals.

Pi and pi star, right?

Sigma p.

Sigma p.

All right.

2p, 2s, sigma.

This would be sigma s, sigma 2s, sigma 2s star, sigma 2p, if you want to be exact about it.

Sigma 2p star.

Pi x, pi y, pi x star, pi y star.

We talked about all this on Friday, I'm just getting us back in the mood.

But the way we left off was that these are-- OK.

So for-- if I fill these all up now-- right.

Remember, we do the MO filling, then we'd have this, and this, and-- ooh.

Now, hold on.

Yep.

This and there and there for oxygen. And then over here we're filling them up.

We got-- those are filled, those are filled, and this is filled for nitrogen. Remember, I showed the movie pouring liquid oxygen, pouring liquid nitrogen. Oxygen has unpaired electrons, which means it's paramagnetic.

It has unpaired.

Electrons, the O<sub>2</sub> dimer.

This is O<sub>2</sub>.

These are the MOs of O<sub>2</sub>.

These are the atoms.

These are the MOs of N<sub>2</sub>.

N<sub>2</sub> doesn't have any unpaired electrons so it's not paramagnetic.

But the thing that we left off on-- that's what I'm saying is, that these two switch order for those dimers, for lithium through nitrogen. Those two switch.

All right.

We talked about that on Friday.

The reason is because for those atoms, these sigma-- these sigma p orbitals can mix in with the sigma s orbitals.

And so you get more.

Because, remember, the more overlap, the greater the energy, right?

The greater that delta between bonding and antibonding, right?

So the more mixing you have, the more spread you get.

And it's so much that for those dimers you push the orbital up.

OK.

That's all what we talked about Friday.

Now, I got on of each.

I got one of these on the board so that you can see them as we now take one atom from one side and one from the other.

It's actually kind of what you expect.

You just have to have one thing in mind when you think about  $\text{NO}$ -- so if this is N, OK, over here, and we're going to put 2s there and 2p there.

Here, here, here, here.

There's my N atom.

And now I've got O over here.

So I've got my-- but, now, see, the thing is always more electronegative.

And the way this works for a heteronuclear dimer is-- let's write this on the board.

So for a heteronuclear, that means that the two atoms are different.

OK.

But for a heteronuclear dimer, the bonding MOs are closer in energy to the more electronegative atom.

Electroneg atom.

Atom.

So what that means is that if I draw-- so if I draw my oxygen levels-- but you know this, right?

You know this, also, from all the work we did on atoms, right?

You know a little something about-- remember, this axis is energy.

This axis is energy.

And so if this is 2s for oxygen, and then the 2p's are going to be sort of like here.

OK.

And we'll fill them up thinking about Hund's rule.

Well, now, what we're going to have is that those-- the MOs for the  $\text{NO}$  molecule are going to come closer, they're going to come further down from that side, right?

Then they come from that side.

Do you see that?

There's an asymmetry.

I made it kind of subtle.

I made it kind-- and why is this curved?

That's not good.

What's going on here?

There's a little bit of an asymmetry here, right?

So-- so this is lower down-- these are lower down, right?

So when you draw your orbitals now-- oh, man.

I'm going to draw this a little bit lower.

Right.

So now when you draw then, you see that bonding orbital is going to be closer.

You see, this is a bonding orbital still.

It's still bonding, right?

It's going to be a sigma p, sigma 2p.

But it's closer in energy to the more electronegative levels.

You see that?

Right.

So it's not symmetric anymore.

OK.

And then the same what happened with these guys, right, they'd be closer in energy to the-- yeah.

They would be closer.

And I'm not leaving myself enough room here.

That's why I've got some videos.

They do it better.

All right.

OK.

We'll leave that one there.

It's a little bit better.

Not much, but a little bit.

So the bite-- now-- oh, by the way, for No-- let's talk about MO for a second, because there's another thing that's important that we talk about, which is bond order, and, you know, the bond-- oh, electron configuration is another thing, right?

So like the configuration, we can write this out, right?

So the configuration of, say, O<sub>2</sub> would be sigma 2s, right, you're just filling them up with two electrons.

Sigma star 2s with two electrons.

Sigma 2p with two electrons.

Pi, 2p with four electrons, and then pi 2p star with two electrons.

That's how that looks.

It's a little bit messy, but I hope you can read it.

That's just reading up, right.

It's the electron configuration of the MOs, just like we would write the s, p, d electron configuration of that, but this is a molecule.

Now, sometimes you might, you know-- say there's a convention we use, which is that the pz is the sigma.

Nobody can agree on it, so different textbooks say px but we're saying p-- if pz is sigma, then this would be like a sigma 2pz, right, and you might, if you really want to write this, you know, you might write that that's pi 2px if you really want to keep track of it.

There's 2 orbitals in there.

You didn't fill 1 orbital with 4 electrons, right?

So this would be pi 2px and pi 2py 4, right?

That's what's happening in there.

OK.

But the other thing you can see from the bot-- from this filling is, also, the bond order.

And so for the bond order-- bond order, here you would have 1/2 times, and it's all of the-- remember, the bond order is all of the bonding electrons minus antibonding electrons divided by 2, right.

And if it's zero, it's not a stable molecule.

If it's greater than zero, it's stable.

And the higher the bond order, the higher the bond strength.

So that's also from Friday.

And so this is 2 minus 2 plus 4 minus 2, and so you get 2.

Nope.

I missed a 2.

Plus 2, plus 4 minus 2.

$2 - 2$ , that kind of cancels out, right, plus 2, plus 4, minus 2, all divided by 2.

So the bond order for  $\text{O}_2$ -- now, the bond order for  $\text{NO}$ -- bond order for  $\text{NO}$  is going to be 2.5.

Well, that's kind of interesting, now, you can see-- OK.

So  $\text{NO}$  is stronger-- more strongly bound than  $\text{O}_2$ , less strongly bound than  $\text{N}_2$ .

And it's kind of interesting.

How can I make  $\text{NO}$  a stronger-- a more strongly bound molecule?

Well, I can take an electron away.

It's a little bit counterintuitive, right?

Because for  $\text{NO}^+$ , the bond order goes up because the electron that I'd taken away was in an antibonding orbital.

Right.

So  $\text{NO}^+$  actually has a stronger bond than  $\text{NO}$ , even though it has one less electron.

You get this all from just looking at which electrons are occupying which orbitals.

In this case, you had in this very much not drawn to scale-- oh, boy-- figure, right.

What you had was one up there.

So if I take that out, my bond order increases.

Right.

OK.

All right.

And the last point I wanted to make about MO theory before we talk about hybridization is with one other example, and that's  $\text{H}_2^+$ .

OK.

So for  $\text{H}_2^+$ , you got something else happening that's important.

So we just talked about electronegativity driving where the-- you know, that this is tilted.

OK.

Fine.

And now we got the fact that chlorine comes into this situation.

You know, it's like chlorine is like, hey, I got 7 electrons.

And hydrogen is like, hey, I got 1.

So how's this work, right.

So this gets to what was brought up Friday, which is, there's another type of MO.

And it's not bonding, and it's not antibonding, it's non-bonding.

That is not the same as antibonding.

Non-bonding doesn't take sides, it just is.

Because the way this works is hydrogen comes at this-- OK.

So hydrogen's up here with this 1s.

Chlorine is, you know, let's say over here.

It's got 3s filled, and it's got-- let's draw these here.

OK.

3p.

And chlorine's got 7 electrons in the 3p.

Do they-- do they all come together and form MOs?

No, they can't, because there's only 1 on this side, right.

So this 1 orbital will form-- let me make myself enough space so I'm going to exaggerate this a little bit.

So this 1 orbital will form a sigma bond.

Sigma 3p, sigma 3p bond.

This would be a sigma 3p star.

Now, that's exactly what we've already done, right, the sigma.

But-- but-- oh, there's something a little interesting about this one, right, because it's an s.

It's an s.

But that's OK because the s can talk to a sigma p.

This is -- we talked about this.

The s can talk to this.

That's OK.

Because it's aligned, right.

Now, but these guys don't have anyone to talk to.

They don't.

So they can't make their pi orbitals.

They can't do that.

And so they just stay there.

These stay there exactly at the same place.

They're not doing anything different.

They're just hanging out.

Those electrons are non-bonding.

So these are not pi orbitals.

Sigma and pi are special names.

They refer to these orbitals of the molecule that involve the LCAO, the linear combination of the atomic orbitals.

These do not.

Nb.

Nb.

That's what we'll call them.

Just to be very clear.

Just to be very clear.

Now, they're filled.

They're filled, and they do not count in the bond order.

By the way, this is, also, a non-bonding orbital.

That's a non-bonding orbital.

Non-bonding orbitals.

These are not going to count.

So in the bond order, you only have-- you have  $\frac{1}{2} \times 2 - 0$  equals one.

These do not participate in bonding.

I got to fill this.

2 electrons are in that one.

OK.

So that's another example of something that happens in these MO diagrams.

One of them brought too many electrons to the party.

And so there's no one to talk to.

But it's OK.

They're just going to stay there and wait.

Maybe something else will come along.

Right.

OK.

Now, that is, right there, what I just drew, look at that.

2p 2s, and you get this thing, right.

Oh, it's a beautiful molecular orbital between-- OK.

Fine.

It was a 1s and a 3p, but you get the point.

It was a p oriented along the axis with sigma symmetry.

That means if I look along the axis-- sorry, cylindrical symmetry, sigma means if I look around the axis, everything is the same, in a circle, right?

Right.

So-- but-- and so that's a nice sigma orbital between two different atoms.

MO theory.

But see the other thing that can happen that is something that I want you to know today, is that orbitals on the exact same atom can also do this.

They can also do this.

They can combine to form a new orbital.

sp, same atom forms an sp orbital.

It's called a hybrid orbital.

So we're going to give a couple of examples.

This is hybridization.

Hybridization.

And I can give a couple of examples, then my [INAUDIBLE] matters.

So this is when two or more ao's that are similar in energy.

They're similar, they're not the same but they're close.

You know, you can't hybridize if one level's way up here and the other level's way down there.

But if they're close, 2s 2p, then those orbitals can combine.

And what they do is they combine to form, and this is critical, sets of equivalent orbitals properly oriented.

This is really important.

That's why I'm writing it all out.

Properly, not property.

Properly oriented to form bonds.

Oh, this is such a big deal, to form bonds.

It's the last word there, but I did it all caps to emphasize.

Because that's what this is all about.

You see, it-- now, let's take an example.

I'm going to take an example of methane.

Who brings what to the party?

OK.

I got H-- I've got 4 H's.

H1s, H1s, H1s, OK.

This is definitely getting repetitive.

H1s.

But I did it.

There they are.

And they come, and they're like, hey, carbon, what you got?

And the carbon is like, OK, hold up.

I got my 2s, that's here, and I got my-- all my 2p's, and what are we to do?

What are we to do?

How can these four hydrogens come and bond with the carbon given that it looks, at least, like the carbon isn't bringing the kind-- you know, how's it going to do this?

And, also-- and this is critical.

This is why I wrote it here.

Properly oriented to form bonds.

You see, because this taps into what we learned with VSEPR. It does, because VSEPR says you got to not repel.

Or, no, you're repelling, but try to minimize it.

I mean, that's what's at the core of VSEPR.

So what this system sees is a way out.

And by way out, I mean a way to happiness.

And by a way to happiness, I mean a way to lower energy where all 4 of these can bond to the carbon atom if it has 4 same bonds.

If it had 4 same bonds, then it can orient them tetrahedrally and make VSEPR happy, minimize repulsions, right.

And you get your methane molecule.

So what ends up happening is this carbon atoms says, hold up.

You 4 hydrogens, you wait there, I'm just going to go into the back and do a little reorganization, and I'm going to come in it with-- I'm going to say, well, if I had-- if my carbon atom comes at-- comes at it with 4 equal orbitals that look like this, those are called  $sp^3$ , because I took all 3 p orbitals and I mix them up with my s orbital and I formed equivalent orbitals.

Now, carbon comes out of the back room, and it's like, hydrogens, I'm ready, I can make a minimum energy, and I can follow VSEPR and we can have-- all have very, very low energy, compared to whatever else I could've done, if I stuck like that.

What I'm really talking about when I'm talk-- OK.

You might know carbon, you know, it doesn't go into the back room.

It doesn't.

But what is it doing?

It's nothing more than solving its Schrodinger equation.

That's what this is all about.

You just change the boundary conditions.

And so what happened-- that's all hybridization is.

You're solving the Schrodinger equation all the time.

We are always solving the Schrodinger equation all the time.

And in this case, carbon said, well, how can I solve it and minimize my energy with the 4 hydrogens?

This gives you the lowest energy.

Because, now-- oh, I got a movie.

Now, I can make those, see.

So I have them-- I have a picture here.

s and those three.

And it says, how can I do it and minimize repulsions?

This is how.

Three equivalent-- 4 equivalent orbitals to make it  $sp^3$ .

And this is where I think a movie helps.

So I'm going to play this.

It's like a minute long.

[VIDEO PLAYBACK]

Each 2s orbital is a two lobe shape converging at the nucleus.

- There's the 2s.

So there are the three 2p orbitals.

However, when hybridization occurs, the s and p orbitals cease to exist, They don't cease to exist, they just didn't-- it's a new boundary condition.

You're solving the equation.

Right.

So just to keep that straight, right?

These are the orbitals-- and the two  $sp^3$  orbitals have an entirely different shape.

- That's true.

We can see that orbital hybridization explains the VSEPR placement of carbons for VLANs electrons.

Since all four  $2sp^3$  orbitals are equivalent, each  $2sp^3$  orbital repels the others with equal force, resulting in identical bond angles.

The carbon atom only hybridizes when it is in a bonding situation.

Here, four hydrogen atoms bond to carbon by overlapping their orbitals with carbon's hybrid orbitals.

So what would be the reason for this?

If we go back and see that both carbon and hydrogen have unpaired electrons, the overlap allows the electrons to pair, and thus go to a lower potential energy.

[END PLAYBACK]

Oh, that's such a beautiful place to stop.

Lower potential energy.

Happier.

That's what this is all about.

OK.

So I-- you know, like I said, you can see now the difference, right, if you just watch-- it's one minute long, and I think it really brings it to life.

So if you're, again, want to see this a little bit more conceptually, these movies are good.

I have one more movie I'll show you that's like a minute long.

OK.

Now, the thing is-- so what we just did is we did carbon hybridizes to  $sp^3$ .

It's called  $sp^3$  because you're bringing 3p orbitals and mixing them with s to form 4.

And that allowed it to do this.

But, you see, if I had ethane, which is  $C_2H_6$ , it's the same thing.

Now-- because, basically, what's happening?

A carbon atom comes into this system of ethane, and it says, I need 4 bonds that are kind of equivalent-ish.

All right.

I got to give one to each hydrogen and I've got to give one to that other carbon over there.

And by doing that, it can minimize its energy, it can minimize repulsions, it can be the happiest molecule it can be.

But to do that, carbon must have these four equivalent sigma like bonds.

It must be able to form those four equivalency.

So that's very similar to methane, except that in this-- so, also, ethane will be  $sp^3$  hybridized.

It will have  $sp^3$  hybridized orbitals.

But then, and you know it's coming, right, so ethane has  $C_2H_6$ .

But what if now I've got ethylene.

I think I have-- maybe I have a picture of it.

Do I have a picture of it?

Yeah.

Oh, there it is.

Ethylene.

Ethylene is used.

This tiny molecule is 2 carbons and 4 hydrogen instead of 6.

This is used in so many applications I can't possibly list them.

If polyethylene is a polymer made out of this as a base, we will polymerize this a little bit later in the semester when we talk about polymers.

It's also used in the back of your supermarket in case you didn't know that.

Very soon-- actually, kind of almost now-ish, most of the fruit we get will be super green when it gets here.

And then they use ethylene to actually induce, sadly, the brightness.

Doesn't taste as good.

But, anyway-- OK.

Now, but for ethylene, you've got a different situation.

Because for ethylene, each carbon atom needs 3 bonds.

Each carbon atom needs 3 equivalent bonds.

And it can do that, too.

So, for  $C_2H_4$ , each carbon atom needs 3 equivalent bonds, and it can do that if it hybridizes in the  $sp^2$  hybridization.

So let's draw that, right.

So these are the original orbitals of carbon.

And they will go to 3  $sp^2$  orbitals.

Notice, I can put one electron in each ready for action.

Ready for bonding.

One electron in each of these equivalent hybrid orbitals,  $sp^2$ .

And I got one left over.

Ho, ho, oh.

One left over.

But the one left over is a p orbital that didn't get in on the hybridization.

But that's OK.

That's OK.

These are p, right.

So this is s and this is p<sup>2</sup>s 2p.

And that's a p orbital right there.

Those are my hybrids, and that's a 2p orbital.

There's another 2p orbital on the other carbon.

Also, got left out.

Now, I only needed 3 because, you see, I formed this nice spatially maximize to minimize repulsion, right, separation, maximize to minimize repulsion so it's a planar-- oh, but it's a planar for another reason.

Because each of these carbon atoms has a p orbital, and we know that 2p orbitals can form a pi bond, right?

And I have a picture of that.

That's exactly what they do.

So this is the sigma.

The hybrid orbitals are sigma, right, because they're along the axis, and look at them.

If I look down this axis, I can draw a circle, and there's full cylindrical symmetry, right.

But I've got this left over p orbital here on each carbon atom.

That's that one there.

And I didn't need it to do this but it's still there.

And so it forms a bond.

It forms a bond.

Right.

It forms a pi bond.

So that p with the p from the other one forms a pi and a pi star, right, in the MOs, and you occupy the pi with the two electrons.

And that's a bond.

So in this you've got-- you've got a bond from this that's above or below-- in this orientation, it's above or below the plane.

And you've got a bond from this.

So you've got a double bond.

That's exactly what we know already, but we know so much more now.

We know so much more.

Because-- because, now, we know that it's not just two-- you know, two bonds and four electrons, we know that they have shape to them.

We know that they have structure to them, and we know that those bonds are different.

These are pi bonds, and that's a sigma bond.

Right.

And you, also, know why this molecule stays planar.

Because if I try to twist it, then those p orbitals won't align, right.

It's like a little spring.

If I-- if I try to twist it, it's going to go right back.

Because it's minimum in energy is when these things can maximize their overlap.

All right.

So that's going to stay planar.

OK.

Now-- and so for this case, you've got a double bond.

So this is double bond, and I think I have a little carbon, carbon, double bond.

And I think I've got a video to show this, because this is another case where, you know, again, I just want to make sure you guys get the intuition here, and I thought this is a pretty cool animation of  $sp^2$ .

So I'll play it for a minute.

[VIDEO PLAYBACK]

[INAUDIBLE] bonds.

The one 2s and two of the 2p orbitals hybridize.

Consequently-- - That's cool animation.

--this hybridization is termed as  $sp^2$  hybridization.

- Look at that.

The hybridization leads to the formation of 3 [INAUDIBLE]

sp<sup>2</sup> hybrid orbitals.

As you can see, each sp<sup>2</sup> hybrid orbital is bilobed.

One lobe bigger than the other.

The half-filled p orbital, which was not involved in hybridization, lies at right angles to the plane of the equilateral triangle.

Now, let us understand how this hybridized state results in the formation of a double bond.

For this, imagine a similar sp<sup>2</sup> hybridized carbon atom approaching this carbon atom.

As these atoms come closer, an orbital overlap takes place along the internuclear axes.

This bond is called a sigma bond.

At this stage, the unhybridized p orbitals which lie above and below the plane of the sigma bond, also come very close to each other and overlap laterally, resulting in the formation of a pi bond between the two carbon atoms.

[END PLAYBACK]

All right.

OK.

So I thought about it.

I watched that movie, of course, many times over the weekend.

And I thought about can I animate, you know, the electrons like that and do the-- and I-- I think I could.

But I don't know that it would be a good use of our time.

So I thought that they did a pretty good job of bringing it to life.

Now, this is a great time for me to tell you about my why this matters.

And, of course, the sp<sup>2</sup> carbon atom matters because of drinking water, of course.

And, you know, this is a well-- in a district in India where more than 50% of the wells exceed the WHO limits for arsenic by around a factor of 5.

1.8 billion people in the world drink fecally contaminated water on a daily basis.

600 million people boil their water in the world.

Boiling doesn't help with arsenic, though, arsenic just stays in the boiled water.

It kills bacteria, but it doesn't help with toxic-- toxic elements.

If you look at the world as a whole, and you look at sort of where there are water-- water crisis, where the water crisis is at this level, it's a little over 3 and 1/2 billion people.

And in those countries, almost 185 or so countries, where this is a serious problem, if you look at the cost of disease, and the cause of death in those countries, 70% to 80% of all disease and almost 30% of all death can be attributed to the water quality.

Right.

So when I talk about water as a problem, I really mean it's a serious problem.

It's a problem of life or death.

Now, this is-- to get fresh drinking water, if you look at the planet, and you say, well, where is the water on this planet?

Where is the water on this planet?

I like this picture because we consume water volumetrically, not in an area, right?

We use water as a volume, and this is the water we have on this precious planet as a volume.

This bubble here is all of the ocean water.

It's about 70-- it's about 97% of all the water.

That bubble there is freshwater that is inaccessible.

It's frozen.

And this one here, you can't really see it, there is another dot there.

That's less than 1% of the world's water.

That is our drinking water ecosystem that is what we are destroying.

And-- so what-- OK.

So there's a lot of things we can try to do about this.

One of the things we can try to do is to see if we can tap into this in a way that's more affordable and more efficient.

Because there's a lot of water here, right, and use that.

But the problem is, if you look at the cost of desalination today as opposed to the cost-- that's desalination.

If you look at the cost as opposed to just digging water out of a well, like the one I just showed you in India, it's over a factor of 10.

Still.

And, in fact, one of the big issues with desalination isn't just the total cost or the operating costs, but the capital cost.

How do you build this plant in the first place?

And so it takes too much money.

If you look at the thing that's at the heart of desalination, it's a filter, right.

It's a filter.

And I talked about this in my last, why this matters, when I talk about separations, and I promised that in my next why does this matter, I'd talk about water.

Right.

Which is what I'm doing.

If you look at a desal plant, this is one of the world's largest, this is a plant called the Hadera plant in Israel, and you look at what their costs are that-- remember, the costs digits-- half of that cost-- almost half of the cost isn't just energy.

And almost all of that energy is in pushing salt water through a filter.

Through a membrane.

That's called a reverse osmosis membrane, because you're going against the osmotic pressure.

And if you look at the membrane itself-- this is a picture of it, it's actually a very small layer on top of this active layer that does all the work-- it's not a very good design.

In fact, membranes for desalination are pretty bad.

Right.

They kind of do everything worse than they should, except that they work, which is good, and they're cheap, \$1.00 a square foot.

But, see, they're very-- they take much more energy than you need.

They foul up very easily.

And then that means stuff gets kind of, you know, stuck in them.

And then you can't take it-- you can't clean them because these polyamide membranes, which are the same polymer used in these membranes for 50 years, haven't changed, the material.

Those polyamide membranes are destroyed by chlorine.

So even in a desal plant-- in a desal plant, if you have drinking water in your feed stream, drinking water has six parts per million chlorine.

That's not much.

They still will go through the cost to remove it.

They will remove it from the feed.

Why?

Because if you leave that little amount of chlorine in the feed, these membranes get destroyed.

By the way, there's 40,000 of these membranes in this plant.

Each one is 2 meters long and 40 square meters of area.

So-- so these are so delicate that you can't really clean them well.

And that's part of what the cost of a plant involves.

That's part of what the cost of a plant involves.

Now, this would be like a plant, right, you have sea water coming in and, you know-- and then you've got your membrane module that's taking the salt out of the water.

And then the product water.

But because this filter is so delicate-- there's the picture-- because that filter is so delicate, you actually have to add a whole lot more to the plant.

So much of a desal plant is built around essentially protecting this membrane.

So that's cost.

That's cost.

Right.

A better membrane could change this.

But like I said, the membrane-- oh, there it is.

You can't Snapchat on that.

Snapchat?

What is it that people do today?

Not Snapchat.

Yeah, Snapchat.

Yeah.

OK.

Thank you.

You can't Snapchat on that.

That's what I meant.

I know it's not My Space, but I don't know-- but, anyway, this is like even-- you know, when I look at-- at membranes today as a material scientist and materials chemists, all of your membranes, I think, that's what it looks like.

Why?

Doesn't have to be the case.

We can do so much better.

And so-- OK.

Energy costs goes into pretreatment, secondary treatment, and this is where this comes back.

This beautiful material that we have now understood more than we did before.

Because, remember, before I showed you graphene and I showed you-- no, I showed you benzene.

And we talked about it as a Lewis resonant structure, remember?

Right.

A Lewis resonant structure to help explain how it looked.

You know, gra-- graphene does not have alternating bonds.

Remember, that's what we talked about before.

It's a Lewis resonant structure.

So it lowers its energy by having two sets of structures that have sort of alternating bonds.

We talked about it in the context of benzene.

But, see, now you know so much more.

You know so much more.

Because, now, you know why graphene is so special.

You know why graphene is so special.

It actually is the hybridization.

It's the hybridization.

Because on top, and on the bottom, of a single atom of a single sheet of graphene, you've got pi bonding all the way.

Those pi bands are going across the whole surface.

And those electrons are critical.

Those electrons that occupy those [INAUDIBLE]

are critical to the very special properties that graphene has.

So you now know the secret.

It's sp<sup>2</sup> hybridization.

And, you know, this has allowed us-- this has-- you might tell I'm kind of passionate about this problem.

Partly impassioned passion about a lot of problems but, also, we happen to work on this, and we have been developing this.

This is like the ultimate membrane.

I can soak this in chlorine overnight.

I can put it in negative pH solutions.

I can heat it up.

It doesn't degrade.

And we have figured out how to poke holes in it at just the right size, or to stitch it together so that maybe you can filter particles by the flow in between these sheets.

Right.

And this material, as a membrane, is so promising that we've actually started to commercialize it as of the last year.

And I think this is really going to make a big difference in a lot of areas.

Why?

Because of sp<sup>2</sup> hybridization.

That is why.

That is why this is such a special material.

OK.

That's my why this matters.

Now-- oh, we did see 2H<sub>6</sub>.

There it is.

We did see 2H<sub>4</sub>.

There it is.

Look at that extra pi bond, not animated this time.

And, of course, you can go on and you could do see C<sub>2</sub>H<sub>2</sub>, and you can imagine what happens in this case, is what you would expect.

Is what you would expect.

Because, now, I've only got-- see, each of these carbon atoms only needs two sigma bonds.

One to the hydrogen and one to the other carbon.

So it's only going to grab one-- it's only going to hybridize-- to make those two equivalent sigma bonds, it's only going to hybridize with one of the p electrons.

So that's sp.

That's sp hybridization.

But, notice, that in  $C_2H_2$ , it-- you know, now, I've got 2p orbitals left over on each carbon atom.

And so you can imagine that with these 2p orbitals, well, you know this already, right.

When-- when we talked about p orbitals, we talked about how, you know, you might-- you know, you've got this, right.

Awe.

OK.

That's supposed to come like out of the plane and into the plane.

They're orthogonal.

They're orthogonal.

Right.

And so-- so two of them are coming at the other atom.

One is like this, and the other's like that.

And then the other atom is one like this and one-- well, it might not at first, but then when it sees those, it's like, hey, wait a second.

If I do this, we can pi bond.

Right.

The other ones are like this, and they pi bond, and so you get a carbon carbon triple bond.

You get a carbon carbon triple bond for a  $C_2H_2$  with sp hybridization.

Now, it's not-- oh, there it is.

Right.

OK.

Forms pi bond network, forms sigma bond network.

It's not just carbon that hybridizes.

It's not-- carbon is a classic example.

But, you know, if you look at  $\text{BeH}_2$ , it's the same thing.

$\text{BeH}_2$ .

Well, Be, Be starts with an un-- can I raise this up?

Be starts with orbitals that don't look like they're going to give me the equivalent two sigma bonds.

Be starts with these as electrons filled.

Right.

But if Be had something more like this, if it had an sp electron in each of those orbitals, well, now, when I see the two h orbitals, the two 1 s orbitals coming at me, I can form those equivalent bonds.

So  $\text{BeH}_2$  will also hybridize [INAUDIBLE]

to make that molecule.

And all sorts of other things.

The deal orbitals can get in on the action.

All right.

So here's  $\text{SF}_6$ , which we talked about.

We drew this Lewis structure.

Took a long time to draw those lone pairs.

But, now, you know that the way this actually works because these are equivalent bonds.

So the only way this can work is if I form hybrid orbitals.

Right.

Right.

But they didn't have to be equivalent bonds until we learned about VSEPR.

And it said, no, I want to maximize my separation to minimize repulsions.

And so then we know that if they're equivalent bonds, the energy of the system will be lower.

Right.

And so that-- that's-- you know, that's going to have to take some d orbitals and create these hybrid orbitals.

Look, they're beautiful.

Because, now, this thing from the sulfur, this thing can say, hey, I can take six of you, and I can be equivalent bonding to all of you.

That's what hybridization is.

OK.

It's written right up there.

We're just keeping on saying the same thing.

OK.

Now, I think that I'm sorry-- I got very excited about the water so I have to throw some T-shirts out.

And, now, where do I-- where do I-- OK.

I don't go there.

I don't go there.

I got to go back there, and back there.

Oh, that's very loud.

OK.

And right there.

And right there.

See you guys on Friday.