CATHERINE DRENNAN: All right. So moving to today's handout, this is one of my favorite parts of the course. Honestly, when I first started teaching 5.111, I said transition metals are rarely covered in the intro chemistry courses. Is it really necessary to cover it here? And I was told it absolutely was.

It's one of the reasons that a 5 on the AP exam is not good enough, that you have to take the Advanced Standing exam. Because the people who teach inorganic chemistry found that people who placed out of 5.111 didn't do as well in their course as people who took 5:111 here. So this is one of the reasons.

And then I started teaching it, and I realized this is one of the-- even though people haven't seen it before and sometimes will get a little scare-- it's actually one of the most fun units. So I absolutely love it, and hopefully you will love it by the end. People are like, we never covered it. Why are you covering it? We're in chapter 16. It's fun. OK.

So transition metals, d-block metals, because they have those d orbitals. Yes, we're going back to talking about orbitals again. And they're called transition metals because you transition from this part of the periodic table with your, what kind of orbitals?

AUDIENCE: s.

CATHERINE DRENNAN: s. To this part of your periodic table with your, what kind of orbitals?

AUDIENCE: p.

CATHERINE DRENNAN: p. So they are the transition metals, and they're often really reactive and very cool. And many of them, since we're on a biological theme, many of them are super important in biology. And I have some of these written down in your notes, but here they are up here.

In the transition metals, you have a lot of metals that we could not live without. Iron carries oxygen to our blood, very important, hemoglobin. We talked about cobalt just now. That is the metal in vitamin B12. So we know why that's important. We have zinc everywhere. Nickel's important for bacteria, not so much for us. But bacteria is important for us, so therefore nickel's important to us. So all of these are really important.

Also, this part of the periodic table is a part of the table that people love that want to make pharmaceuticals or want to make new kinds of electrodes or batteries or all sorts of things. There's a bunch that are used as probes. We talked about imaging agents, detecting cancer, and all sorts of different things like that. Many of these transition metals are used in those probes and also in pharmaceuticals.

And so this is sort of a very rich part of the periodic table, where those d orbitals allow for properties that are incredibly useful for our health and for doing all sorts of stuff. So I love this part. Again, some of the biological-- global cycling of nitrogen. We talked about nitrogen fixation, that triple bond. It's really hard to break nitrogen apart, but bacteria can do it. It does it using transition metals.

Fixing carbon, hydrogenase, if you want to make hydrogen fuel cells that are more biological. Again, biology uses transition metals in this. Making vitamins, making deoxynucleotides, respiration, photosynthesis, it's all due to transition metals.

All right. So we'll start with just one example, or one of In Our Own Words segment. And this focuses on nickel, which is something very important in bacteria. And this is actually an example from a collaborative project between my lab and course 6. And I know a lot of you are thinking about being course 6 majors, so I thought I would tell you about some research of Collin Stultz, a course 6 professor.

So he was doing some computational analysis on these proteins that we're studying. So many of you at this point in the semester probably feel like you might be getting an ulcer. But unless you have H. pylori in your gut, you probably are not actually getting an ulcer. And you just take a little B12, you'll feel a lot better. OK. So here's the video.

[VIDEO PLAYBACK]

- My name is Sarah Bowman, and I am a post-doctoral fellow at MIT. I am working on studying a protein from Helicobacter pylori, which is pathogenic bacteria. Its kind of ecological niche is in mammalian stomachs. It's actually very difficult to treat using antibiotics, because a lot of times when you're given antibiotics they're going to actually be broken apart by the acidity of the stomach before they actually ever get to killing the Helicobacter pylori.

Transition metals in biological systems are actually really important. They increase the range of reactivity that proteins and enzymes are able to access. Nickel is a transition metal. I mean, it's a transition metal that's actually fairly rare in biological systems. So one of the big things that H. pylori uses nickel for is an enzyme called urease. Urease requires something like 24 nickel ions, which is a lot.

Urease is one of the proteins that allows for a lot of buffering capacity of the organism, of the H. pylori organism. The stomach pH is very low, so pH 2-something. And this bacteria has to swim through the stomach and then colonize it.

And you'd think that the stomach would just break it apart like it breaks apart your food. But in fact, the bacteria itself has mechanisms in place that allow it to create buffers that allow it to move through the stomach and live in the stomach. And one of those enzymes, and it's really important for that, is urease.

In humans, nickel, as far as we can tell, is not essential for any enzymes, whereas in Helicobacter pylori, for instance, nickel is an essential transition metal. And so a really intriguing thing to kind of think about is just whether we could somehow target the nickel requirement in this organism and in other bacteria that would allow us to kill these pathogenic bacteria while not doing anything that would be harmful to humans.

[END PLAYBACK]

CATHERINE DRENNAN: So I like that video partly because it brings back acid-base and buffers, as well as talking about transition metals. And I love the bacteria being attacked by the acid and then making a buffer and saving itself. It's awesome. OK.

So one of the reasons why these transition metals are so powerful, they can do so many things, is that they like to form complexes, and they like to form complexes with small molecules or ions. And those ions often will have a lone pair of electrons, and the metal wants that electron density. It wants the benefit of that lone pair. So when you have this lone pair, the metal will come in contact with that lone pair, and it'll make a very happy, very happy metal.

And we can think about this interaction here as the donor atoms are called ligands. And now let's review something we learned before about whether this is a Lewis acid or a Lewis base then. OK, 10 more seconds. That's right. So it's a Lewis base.

So if we put this up here, donor atoms are called ligands, which are Lewis bases, and the Lewis bases donate the lone pair of electrons. And again, we can think about the definition that we've been more used to, where a base is taking H . It's accepting the proton from the acid. But there, when it's taking H , it's taking H without its electrons. So it's actually donating its loan pairs to form a bond.

And then we can think about Lewis acids. So the acceptor atoms, which are our transition metals, are Lewis acids. They accept the lone pair. And when an acid that has a proton on it loses H , it is taking the electrons with it, because H is leaving without its electrons. So these definitions work, but these are sort of more broad definitions.

So here, our metals, any of our transition metals, are going to be our acceptors, our Lewis acids. And here are a bunch of ligands. We have water. We have NH3. We have CO. They have lone pairs. They can be donor atoms.

And the ligands form complexes with the metals. And the kind of complexes-- they're often called coordination complexes, and that's a metal that's surrounded by ligands. And here's a little example, a metal in the middle, and it has the ligands around it.

So let's consider this coordination complex now and think about what this picture is telling us. So we have our coordination complex. We have cobalt in the middle, and we have NH3 groups as our donor ligands. And here this bracket indicates the overall charge is plus 3. Again, the transition metal is going to be the Lewis acid. It's going to be accepting the lone pairs from the Lewis bases, which are the ligands, or the donor atoms.

Now, we can think about a new term called "coordination number." And that's simply the number of ligands that are bound to the metal. So a CN number of 6 would indicate six ligands make up what's called the primary coordination sphere, which is the things that are bound directly to the metal. So CN numbers for transition metals range from 2 to 12, but 6 is probably the most common.

So before we think about the shapes of these molecules, let's just look at the notation for this, so coordination complex notation. So I would write this structure up here within brackets-- cobalt bracket NH3. You have parentheses around NH3. There's six of those. Another bracket here with a plus 3 charge, indicating the charge on everything, this whole structure.

But often, coordination complexes with a plus charge will have counterions around. So there might be, say, three chlorine minus ions around, and so that could be written like this, or it could be written like this. If you see Cl3 outside of those brackets, it means that they're counterions.

So if I looked at this, I'd say NH3 is within the brackets. That means it's bound to the cobalt. So that would tell you there are six things bound to the cobalt. The Cl is outside. That indicates it's a counterion. There are three of them. So there are three counterions, which then tells you the charge must be plus 3. All right. So there is our notation.

All right. So now we're back to thinking about geometries. So this is one of the things I love about this part. I feel like some people in the course are just like, new topic, new topic. Oh, man, when is the new material going to end? Well, you find you get enough into chemistry, and you start revisiting topics you've already seen before. So this is great.

All right. So coordination number 6. We haven't maybe heard coordination number 6, but that's pretty easy to remember. It's the number of atoms bound. What type of geometry is this? You can just yell it out. Right. So that's octahedral geometry. Again, the solid triangles coming out indicate they're coming out at you. Back dashes are going back, and we have our axial.

All right. So let's see how well you remember CN 5 structures. And you can keep this up here, and you can tell me what the name of those two geometries are. All right. Why don't you take 10 more seconds.

And here are our structures in real life down here. People are just like, I want to put see-saw. No, no. That is the parent geometry of see-saw, but not see-saw itself. OK.

So we have the trigonal bipyramidal and the square pyramidal. So I'm holding up the square pyramidal right now. And then we have the trigonal, because it's trigonal along here, bipyramidal. So it's sort of like one pyramid here, one pyramid there, so bipyramidal. And if I took off one and we had a lone pair, then we would get our friend the see-saw. OK. Next we have this. What's that one called?

AUDIENCE: Square.

CATHERINE DRENNAN: Square--

AUDIENCE: Planar.

CATHERINE DRENNAN: Planar, yep. And this one? Tetrahedral. And now CN number of 3. What is this one?

AUDIENCE: Trigonal--

CATHERINE DRENNAN: Trigonal planar. It's in a plane kind of, if I hold the bonds and they don't fall off, and it's kind of trigonal. And then what about the last one?

AUDIENCE: Linear.

CATHERINE DRENNAN: Linear. OK. And let's just run through and think about the angles as well. With octahedral, what are our angles?

AUDIENCE: 90.

CATHERINE DRENNAN: 90. Trigonal bipyramidal?

AUDIENCE: 90 and 120.

CATHERINE DRENNAN: 90 and 120, that's right. So we have one 120 around here, and then the top parts were 90. OK. We have the square pyramidal. 90. Square planar?

AUDIENCE: 90.

CATHERINE DRENNAN: 90. Tetrahedral?

AUDIENCE: 109.7?

CATHERINE DRENNAN: 109.5. Give credit for 0.7 too. That's quite close. Trigonal planar?

AUDIENCE: 120.

CATHERINE DRENNAN: 120. And linear?

AUDIENCE: 180.

CATHERINE DRENNAN: 180, right. So you're going to need to remember these for this unit, but that's OK because you need to remember them for the final anyway. So it gives you a nice review. All right. So we got every one. We got them down. Can look up your old notes. Just review.

All right. So coordination complexes also have another name. They can be called chelates. Just another name for coordination complex. So chelates can be the thing. But you can also say that the ligand will chelate as another way of saying that it will bind to a metal. And it can bind more than once with one or more sites of attachment.

And the word "chelate" comes from claws, and I like that picture. I feel like, yes, these ligands coming in like claws and binding that metal. They're chelating that metal.

So there are different names depending on how many points of attachment they have. And we have what's known as monodentate-- "dent" for dentist or tooth. So that's one point of attachment. And I bet that without having seen this material ever before you can tell me what the rest of these are. What do you think bidentate means?

AUDIENCE: Two.

CATHERINE DRENNAN: Two. Tridentate?

AUDIENCE: Three.

CATHERINE DRENNAN: Tetradentate?

AUDIENCE: Four.

CATHERINE DRENNAN: Hexadentate?

AUDIENCE: Six.

CATHERINE DRENNAN: Six, right. There's not one for five. But this is good. So don't lose a point on this. I feel like sometimes people lose a point on this on the exam. You knew it in class before I taught it. You don't want to like somehow work backwards. So this is easy points right here. Just remember on the exam, wait a minute, maybe I already know this.

All right. So let's look at some examples of chelating ligands that bind with multiple points of attachment. And the first one-- we're kind of on a theme today-- is vitamin B12 that we're going to look at. So this is called the corrin ring. Cobalt is in the middle, and that ring binds with four points of attachment. So it is a tetradentate ligand, this corrin ring.

There is also an upper ligand, which is 5 prime-deoxyadenosine, and a lower ligand that's called dimethylbenzimidazole. You don't need to know their names. Overall, it has six ligands in octahedral geometry. But the corrin ring is a very nice biological example of a multidentate ligand. Heme would be the same.

I thought I would show you this rotating around so you get a better sense and tell you that this structure of this vitamin was determined by Dorothy Hodgkin, who won the Nobel Prize in 1964 for determining the structure by crystallography and also solving the structure of penicillin.

This was the most complicated molecule to be solved by crystallography, and a lot of people said that technique could never be used to do something that big. She showed that they were wrong.

In terms of determining the structure of penicillin, it was during the war. And people wanted to make more penicillin, but they had no idea what the structure was so they didn't know what to make. And she figured out the structure. And it's a weird-looking molecule, so no one would have guessed it without knowing the structure. So for her pioneering work in crystallography she won the Nobel Prize.

All right. So vitamin B12 is one example of a chelate. Another that's probably more that you probably hear about the most-- it's almost synonymous with the word "chelate"-- is EDTA. Here is the EDTA molecule, and you see that it has lots of lone pairs that are just dying to grab onto a metal. So we have six. We have 1, 2, 3, 4, 5, 6, six things that are capable of chelating that metal.

And so here is what the complex looks like. So the red oxygen can chelate. The green oxygen here can chelate, the nitrogen here in dark blue, the other nitrogen in dark blue here, light blue oxygen here, and also the purple oxygen. So now why don't you tell me what the geometry of this would be as a clicker question. You ready?

AUDIENCE: Yeah.

CATHERINE DRENNAN: Yeah. 10 more seconds. Should be fast hopefully. Yeah, great, 86%. It is octahedral. And sometimes it's a little bit hard to see that, but I helped you out by drawing those bonds in black that you needed to look at.

So we have four that are in the plane here, one above and one below here. So that is octahedral geometry. Also, how many points of attachment? What kind of dentate ligand is this?

It's hexadentate as well. So it has six points of attachment here. All right. So EDTA is a really good metal chelator. And part of the reason that it is such an awesome metal chelator is because of entropy. So we're back to entropy again.

So the binding of EDTA to the metal is entropically favored. And the reason for this is that metals that are, say, in your body, like if you happen to eat some lead paint, and that lead is hanging out. It's not just by itself.

It's coordinating hopefully just to water and not to proteins in your body. But when you take some EDTA to prevent your lead poisoning, one molecule of EDTA will bind to metal, and all of these waters are going to be released.

So I have over here some lead with a whole bunch of little waters. This is quite an ordered system. But if I take out all of those waters here, that's a lot more entropy going on than what we had. And then you have one chelating ligand here, and that's a pretty simple system. So this is ordered. This is disordered.

So the binding of EDTA, one EDTA releases six water molecules, and that makes it very favorable to do this. And because of that, chelating molecules, or the chelate effect, molecules that are chelates, like metal bound to EDTA, are unusually stable because of this favorable entropic effect, this release of water.

So the release of water, the release of increasing entropy, drives that metal chelation, and you sequester your metal, which is really good if you're trying to avoid lead poisoning. So I think this is a nice example of our friend entropy driving a reaction.

So a lot of you did really well on the exam talking about factors of delta H and entropy and when you'd have favorable delta G's. Here's another nice example where the chelate effect explains why metal chelates are so unusually stable.

All right. So uses of EDTA. I already just told you one. Lead poisoning-- all ambulances have EDTA in case someone is eating some lead paint. Another thing that EDTA is used for, which I think is fun, you should all go check if you buy little packaged goods, and they have a long list of chemical ingredients. Look for EDTA. It's often there.

And it says it's "added for freshness," which means that bacteria need metals. You have EDTA. EDTA sequesters the metals. The bacteria can't live on the food that you're eating. So instead of saying, food additive added to kill the bacteria that were otherwise growing on your food, they say added for freshness. And I do think that is an improvement.

All right. Another thing, we've already talked about the importance of cleaning bathtubs. To chelate calcium out of bathtub scum, you have EDTA or other metal chelates. And then I have my favorite other example of the use of EDTA. This favorite example is in Hollywood, the movie *Blade.* How do you kill a vampire? Vampires drink what?

AUDIENCE: Blood.

CATHERINE DRENNAN: Blood. Blood has?

AUDIENCE: Iron.

CATHERINE DRENNAN: Iron. EDTA chelates?

AUDIENCE: Iron.

CATHERINE DRENNAN: Iron. So you get a little dart, and you have-- you can kind of see them maybe up here-- they're filled with liquid. That's EDTA. You shoot the vampire with EDTA, and the vampire just disappears, just kind of turns to sort of dust.

[LAUGHTER]

Because it's like mostly iron, and the iron gets chelated. But it happens right away. But anyway, I think that's cool. Yes, what a good way to kill a vampire. EDTA, it's brilliant. Excellent use on Hollywood's part for EDTA. OK. Metal chelates, all sorts of potential values that they have. OK.

So when we're talking about coordination complexes, we're talking about geometries. Sometimes the atoms can be arranged in different ways. And when you have these geometric isomers, they can have very different properties. So just look at an example here. It's a platinum compound, a platinum compound who has two NH2 groups and two chlorine groups.

And you could arrange those in two different ways. You could put the NH3 groups on one side and the chlorine groups on the other side, and that would be cis. These are cis to each other. Or you could put a transconfiguration, where chlorine is here, and then another chlorine is trans on the other side. And the same with this.

So cisplatinum here is a potent anti-cancer drug. And it has to be cisplatinum because it binds to DNA, and the two bases of DNA displace these chlorines. So if they're not on the same side, it can't bind to the DNA. And so this prevents the cancer cells from being repaired from damaging agents.

Transplatinum does absolutely nothing that anyone knows. So it's exactly the same composition, but because they are different isomers from each other-- and I have, let's see, ah, over here-- different isomers of each other-- and so chlorines on the same side, cis versus the trans-- have completely different properties.

So cisplatinum got a lot of fame because it cured Lance Armstrong of cancer. Lance Armstrong now, of course, is a much more controversial figure than he was at the time. But still he created an amazing charity that hopefully is still doing well despite some of his fall from fame. OK.

So another type of isomer are called optical isomers, also called enantiomers or chiral molecules. And these are one, again, you have the same composition, but they are non-superimposable. They are, in fact, mirror images of each other. So if my head was a mirror, these would be mirror images of each other.

And I could try very hard to superimpose them, bringing the blue molecules over here, but then the green and the red don't match. You can come and try. These are, in fact, non-superimposable mirror images from each other.

And sometimes they can have very similar properties. It depends. But if you put molecules like that that are known as chiral, chiral molecules, i.e. enantiomers-- non-superimposable mirror images. The human body is very much of a chiral environment. You have enzymes designed to bind things in a particular way. So they can have very, very different properties. OK.

So we have to do some d-electron counting before we end today. And I love this because it's really pretty simple to count d-electrons. And so we're going to just take a look at some examples. And for doing this part, we're going to start using our friend the periodic table again. And we need to find oxidation numbers, which we just talked about in the last unit.

So if we have a coordination complex with cobalt, and this cobalt has those six NH3 groups and our plus 3 charge-- so this is the complex that we have been talking about-- let's now figure out what the oxidation number of this is.

And so this NH3 is neutral, so that's given as a hint. Many of our ligands are going to be neutral ligands. So if that is 0, what is the charge on the cobalt?

AUDIENCE: Plus 3.

CATHERINE DRENNAN: Plus 3. Now we're going to use the rules of d-count. So we have a d-count. We need to look up the group number from the periodic table, which, in this case, is 9. Then we have minus the oxidation number, so we have 9 minus 3, or 6. And so this is a d6 system. And that is all there is to doing these counts.

So let's just try another one. So we heard about nickel. We'll do nickel. Nickel is coordinated by carbon monoxide, and there are four of those. So what is my charge on the nickel going to be, my oxidation number of the nickel? So what's my overall charge of this complex?

AUDIENCE: 0.

CATHERINE DRENNAN: 0. CO is also going to be 0. There's no charge on CO. So what is the oxidation number of nickel?

AUDIENCE: 0.

CATHERINE DRENNAN: 0. So then we can do our d-count. The d-count, what is the group number for nickel?

AUDIENCE: 10.

CATHERINE DRENNAN: What is it?

AUDIENCE: 10.

CATHERINE DRENNAN: 10. This is the kind of math that always makes me very happy. 10 minus 0 is 10. So that is a d10 system. All right. We'll do one more over here. And the next one is a clicker question. Gives me time to write

AUDIENCE: Whenever you're ready. We're out of time.

CATHERINE DRENNAN: Yep. All right. Let's just do 10 more seconds. Yep. So here our overall charge is minus 1. We have the chlorines are minus 1. NH3 is 0. Water is 0. So this has to be plus 2 because plus 2 minus 3 is minus 1. We have 9 minus 2 is 7, so it's a d7 system. All right. So Wednesday, d orbitals. I cannot wait.

Yes. All right, 10 seconds. OK. Does someone want to tell me why that's the right answer? Anybody? We got a nice dangly thing for your keys or ID. No? All right.

So here we're thinking about whether things are better reducing agents or better oxidizing agents. And here we're given two different redox potentials-- minus 600 and minus 300. So the one that is going to be the lower number is going to be better at reducing other things. It wants to be oxidized itself.

And then we can think about whether it's a favorable process in terms of whether the thing that likes to reduce is actually doing the reducing. That's going to make it a spontaneous process. All right.

So these are the kinds of questions for the oxidation-reduction unit that we just finished. And this will be on exam 4, which, amazingly, we just finished an exam, and there's another one. So exam 4 is two weeks from today. All right. From Friday, sorry, two weeks from Friday.

All right. So today we're going to continue with this unit on transition metals. The next exam is going to have oxidation-reduction and transition metals and a little bit of kinetics. Kinetics is our last unit. So we're getting very close to the end of the semester.

So we're finishing up the handout from last time. Again, we're back to the periodic table. We're thinking about transition metals. We're thinking about that middle part of the periodic table, and so we're thinking about d orbitals.

So there are five d orbitals. How many s orbitals are there?

AUDIENCE: One.

CATHERINE DRENNAN: One. How many p orbitals are there?

AUDIENCE: Three.

CATHERINE DRENNAN: Three. And so d orbitals have five. And we're not going to talk about really anything beyond d orbitals in this class. And frankly, not very many people do. But d orbitals are amazing, so we have to fit them in.

All right. So there are five d orbitals. And they're up here, and you need to be able to draw their shapes. And the bar for drawing the shapes is actually pretty low. So these are my drawings that I made. And so you can probably do just about as well.

All right. So the one that has the most unusual shape is the dz squared. And so it has its maximum amplitude along the z-axis. And for this unit, our z-axis is always going to be up and down here. y is in the plane of the screen, and x is coming out toward you and also going into the screen.

And so dz squared has its maximum amplitude along z, and it also has a doughnut in the xy-plane. And so I also brought a little model of this. So here's dz squared. We have maximum amplitude along the z-axis, up and down. And we have our little doughnut in our xy-plane.

So then we have dx squared minus y squared, which has maximum amplitude along x and along y. And that would look like this. So we have our maximum amplitudes that are right on axis. So if this is y-axis and x is coming out toward you, those orbitals are pointing right along that coordinate frame.

The other three orbitals look a little bit like dx squared minus y squared, but they're not on-axis. They're off-axis. They're in between the axes. So we have dyz. It has its maximum amplitude 45 degrees off of the y and the z-axis. So if this is z-axis here, there's no maximum amplitude along here. It's 45 degrees off. So it's right in the middle between the z and the y.

So dxz has its maximum amplitude 45 degrees between x and z. So that would be pointing the other way. And so I tried to draw this keeping the reference frame the same. It's a little hard to see the orbitals, but it would be kind of this. So we rotate that around, and so that's what that would look like.

And then dxy we have maximum amplitude 45 degrees in between the x and the y. So x coming out, y in the plane. And so this is, again, a little bit hard to draw. If I drew it absolutely perfectly and not tilted at all, you kind of wouldn't see anything. But that's what that would look like.

So again, the names of this, it tells you about the relationship between that orbital, that maximum amplitude, and the axis that we have defined. So this is very important to know that these guys are in between the axes, right in the middle, 45 degrees. And you'll see why in a few minutes why that's important.

OK. So just to practice, here are some slightly better pictures of the orbitals. And this is the coordinate frame over here, and now we have the orbitals inside that. So again, z is going up, y is in the plane of the screen, and x is going back and also coming out toward us.

So which is this d orbital? You can just yell it out.

AUDIENCE: dz squared.

CATHERINE DRENNAN: dz. Yeah, that's easy to remember. That's the unique-looking one. What about this one? First think about the plane. So it's the xy-plane. And then, is it on or off-axis? So which one is this?

AUDIENCE: [INAUDIBLE]

CATHERINE DRENNAN: Yeah. So this one is on-axis. You can see the maximum amplitude of the orbital pointing right along those axes. So it's right in the corners of that square there. And then what about this one down here?

AUDIENCE: [INAUDIBLE]

CATHERINE DRENNAN: Yep. So that would be dxy. So it's in the xy-plane, but it's 45 degrees off the axes. So it's in between the axes here. And what about that one?

AUDIENCE: .

CATHERINE DRENNAN: Right. So it's along both z and y here. And then this last one, which is drawn to kind of come out toward you, so that is along x as well. So that's dxz, and it's going up along z.

So you can look at the coordinate frame, which we'll try to keep consistent, and ask yourself, is it on-axis or off-axis, and which plane is it in? And that will allow you to name them and also to draw them.

So just to kind of give you more of a three-dimensional sense, there's these little movies that I'll show you now. And so you can get a better sense of that awesome doughnut. It's going to make you hungry. They even colored it like a really nice original doughnut that you would get at Dunkin' Donuts.

So the doughnut is in the xy-plane, and these other lobes are along z. So now we have dx squared minus y squared, and you can see that the maximum amplitudes, again, are along the axes. Key-- they're along the axes here. I don't know why it comes out towards you and-- I didn't, yeah. But it gives you a good three-dimensional sense of this.

All right. So dxy now, again, in the xy-plane. But instead of being on-axis, it's 45 degrees off-axis. So you can see, I think, in this really nicely, it's right between the axes, but it's not touching them. The axes sort of separate these orbitals.

And then we have xz. So now we're going up along the z-axis and in the x-plane. And here it comes at you again, 45 degrees in between z and x. And then our last one, we have yz.

So the shapes of those later three, actually even four of them, are the same. It's just a matter if they're on or off-axis and which plane they're in. So this is not too hard to draw.

All right. So why is this important? Why should we care exactly how the orbitals are oriented? And the reason that you should care about that is because it can explain a lot of the special properties of transition metals.