[SQUEAKING] [RUSTLING] [CLICKING]

JOHN DOLHUN: Good afternoon, everyone. Welcome to The Ellen Swallow Richards lab part two. There is-- the room will be a little bit more sparsely populated today. I wonder why. Lab reports are due, right?

So we're going to get started. And I'm going to start off by talking about probably the most important parameter in measuring the health of the river, and that's the dissolved oxygen. So when we look at the dissolved oxygen concentrations out there, for a river to really be healthy, we're looking for something greater than probably about 8 PPM. If the concentrations fall down to 5 or less, then you'll start to see the fish move around erratically, because they're trying to get their oxygen.

So less than or equal to 5, and we have some stress placed on the aquatic life. If the dissolved oxygen concentration goes below 2, even for an hour or two, you're going to have fish kills. So let's take a moment to actually see how this oxygen gets into the water, and this is the big picture.

So if you look at this, you've got two processes going on. You've got the oxygen coming in from the atmosphere, and it's diffusing into the water system. You also have down below, you've got aerobic biodegradation taking place.

So we've got oxygen diffusing into the water, and then we've got aerobic biodegradation. These are the two main processes that are taking place, and both of these processes are going-- they have different kinetics, yet they're coupled together, coupled in the sense that if aerobic biodegradation uses up the oxygen, more oxygen will start to dissolve in. There's a transfer driving force that actually lets more oxygen dissolve into the water.

The rate at which oxygen is dissolving is actually proportional to the deficit of oxygen that's in the river. And that deficit is equal to the equilibrium level that we would expect minus the actual level that we actually find. So the amounts of oxygen that we're talking about are very tiny.

And there was a brilliant chemist by the name of William Henry, who actually, back in 1803, came out with Henry's Law. And what Henry's Law says is very simple. The solubility of a dissolved gas in a body of water is proportional to the partial pressure of that gas over the water.

There's even a Henry's Law constant here, and there are tables of these. There are thousands of them. For every solvent, for every gas, for every temperature combination that you can think of, there is a Henry Law constant that you can plug in.

Let's think about, what is the partial pressure of oxygen on any given day? How do we calculate that? Yes, yeah. Alec.

AUDIENCE: So if you have total pressure, and then you know the makeup of the atmosphere. You can just break it down. So say your pressure is 1 atmosphere and you have 20% oxygen in the at-- that's not right. But 20% oxygen in the atmosphere, your oxygen would be contributing 0.20 atmospheres to the pressure [INAUDIBLE].

JOHN DOLHUN: Very good. Very good, Alec. So Alec said that you've got 20.9% oxygen in the atmosphere. So we can convert that to a decimal. Partial pressure of oxygen on any given day would be the percent of oxygen that's in the atmosphere times the atmospheric pressure on that particular day minus the pressure of the water vapor that we're talking about. So that gives us a handle on this.

I want to show you probably the simplest example of Henry's Law. I brought a bottle of Coke here, and these are hard, right, because they pack them-- there's a head of carbon dioxide gas over the liquid. Now when I open this, I'm going to release the gas, so the partial pressure of CO_2 over this liquid is gone. We should start to see bubbles coming out of the solution, right?

[HISSING]

Oh good, it's not an explosive one, so that's good. So there are the bubbles of the dissolved carbon dioxide gas starting to come out. And that's a simple example of Henry's Law. Can anybody think of a more complex example? Anybody a diver here? Kelly, you're a diver?

AUDIENCE: Oh, no, I'm just [INAUDIBLE].

JOHN DOLHUN: OK. When you--

AUDIENCE: Yeah. It's like the bends, right? Because as you go down further, the pressure increases so that more nitrogen can dissolve in your blood, you said?

JOHN DOLHUN: Yeah, very good. Yeah, so you've got gases in your blood. And when you go down, when you're diving down under a great pressure, when you start your ascent, all of those gas bubbles are going to come out. And they have to go somewhere. They could migrate anywhere in your body, and it could cause a rash. It could cause some type of joint pain. You could end up with paralysis, even death.

So divers have to really, really time their ascent. That there's a name for that. Kelly said the bends. That's one of the names. It's also called DCS, Decompression Sickness.

That's a great example of Henry's Law. What about if we think about temperature and dissolved oxygen? Let's actually take a look at that. Here is a graph of oxygen concentration and temperature. What stands out to you here? Sean?

AUDIENCE: It decreases exponentially with temperature.

JOHN DOLHUN: Yeah, as the water is warming up, look what's happening to the dissolved oxygen. Why do you think that's happening? And why, when the water is cold, can we hold so much oxygen in that cold water? Kelly.

AUDIENCE: Is it because, as the temperature increases, the average kinetic energy of the molecules increases, and the distribution broadens, so more oxygen has the ability to escape?

JOHN DOLHUN: Good, good. As the temperature increases, the kinetic energy of the water molecules is increasing, and they're starting-- think about when you boil water. Bam, they're bubbling out. That hydrogen bonding network is gone. All those intermolecular forces of attraction are gone.

Now think about this side when the water gets really cold and you form this hydrogen bonding network. I mean, it expands, right? Pipes break. That's why our pipes break. It's because of the hydrogen bonding in the ice crystals when they expand out. But look at the crevices here in these things.

So the oxygen molecules can swim in and out of those, and they can become weakly trapped and pinned in. And there you've got a dramatic increase in the amount of oxygen that can dissolve in the colder waters. The other thing to keep in mind about the water is the heat capacity of water is much greater than the heat capacity of air. It takes a lot more energy to change the temperature of water.

I mean, you can-- I've gone out when it's 85 degrees, and I want to go for a swim across from my house. And I jump in, and it's only 67 in there. So the water hasn't caught up to the temperature of the air.

Now there are a lot of different equations out there that actually would let us calculate the potential of a body of water to hold oxygen, theoretically, how much could be out there on a given day. And the more complex the equation, the better the results. I found a couple empirical, simple equations that give really good results, so I'm going to go with these. And they're all based on altitude and temperature.

If you look at the equation, it's all about the atmospheric pressure, the water vapor pressure, and the temperature of the water, which is in Celsius. So if we look at that equation, if the pressure, the atmospheric pressure goes up, right, what happens to the dissolved oxygen concentration? If you're trying to calculate the saturated level of dissolved oxygen, and the pressure goes up, the atmospheric pressure, what happens to the DO concentration? I promise this is not a trick question. This is-- Alec.

AUDIENCE: It increases.

JOHN DOLHUN: It absolutely increases. Yeah. Now what happens if we have a high water vapor pressure that day? You might be on a mountaintop stream. Water vapor pressure is very high. What happens to the DO? Someone?

AUDIENCE: Decreases.

JOHN DOLHUN: Sam, decreases. Good. And the other situation, which if the river temperature is high, then it's intuitive, right? Your DO is going to be smaller. So that's how you use those equations.

And then what you can do with the equation is you can calculate a thing called the percent saturation level, which is the actual DO that you get in the lab divided by the potential of water to hold oxygen based on the temperature and pressure above. And that saturated level is going to tell us a lot about the condition of the water. Usually-- I mean, you could have from 90% to 110%, something like that. That's normal.

If you get to 120% or greater, that's a problem, because there are different diseases that the fish can-- they have an oxygen bubble disease fish can get if it becomes too saturated. So this is something that you'll do these calculations, and the TAs will carry a pH meter. They'll record the air temperature, the atmospheric pressure, the river water temperature each day you go down, and they'll put that on the whiteboard.

Here is some simple data that will allow you to interpolate the water vapor pressures based on temperatures. And you can go in between these, and you'll get very good results. If you can't find it on here, you could use this equation. But the T here is the temperature of air, and it's in Kelvin. And back here, this is the temperature of the river in Celsius. So don't get confused with that.

And now I'd like to spend a few minutes talking about the actual method that we're going to use in this lab. This method was actually discovered by a graduate student by the name of Winkler back in 1888. And the method has withstood time. What he did is he developed this series of oxidation reduction equations to actually measure the dissolved oxygen concentration in saltwater. And he came up with this series of equations. And it's been 130 years, and this is still here.

Today they have the fancy dissolved oxygen meters, which we're not using, but you can just drop it in and get your DO. They used this method to calibrate those meters today. So you're using-- you're going to be doing the real chemistry here. The whole method is based on that if you've got oxygen in the water, iodide is going to get oxidized to iodine, and you're going to be able to titrate the iodine with sodium thiosulfate.

So let's just go through the method just a little bit here. So we start with manganese sulfate, and we form a white precipitate. But if oxygen is present in the water, oxygen is an oxidizing agent here. It's going to oxidize this manganese hydroxide to this tetravalent manganic species. So you've got manganese over here, which is a plus 2, and over here, you've got a plus 4.

There are a couple-- in the literature, there's some argument about this. Some people feel the species is a trivalent, and they're saying it's MnOH3. So on this side, you'd have a trivalent manganese species. Other people are saying no, it's MnO2. It's a hydrated form of mnO2. It's tetravalent.

What I've done is I've combined both these to give you this manganic species here, which works. This is tetravalent. So notice that when you've formed this, this is like a brown flock. You've captured the oxygen at this point. Now you've got to dissolve that, so we add sulfuric acid to dissolve that brown precipitate.

And then there's iodide in there. So this oxidized manganese will oxidize iodide to iodine. We can then detect it with sodium thiosulfate, and the iodine gets reduced back down to iodide. If you look at this, for every oxygen, every one oxygen, you make to manganic species. Each manganic species gives you one I2, and each I2 requires two thiosulfate. So that's a 4 to 1 ratio, thiosulfate to oxygen.

So let's take a look at your first day, which is the standardization of the thiosulfate. We have to know exactly what the concentration of that is when we titrate the river water to actually get-- to home in on the exact concentration. So we're going to be using a primary standard, potassium biiodate. Does anybody know what the qualities of a primary standard are?

When you pick a primary standard, it's pretty important. You want something that has certain properties. Anybody ever worked with a primary standard? No?

OK, well, first thing, notice how big this is. It has a very high formula weight. So that's actually a good thing. Not all primary standards have that. When you mass it out, you're going to have less error on the balance because of the mass of this thing.

Also, it has to be something that's pure. And this is like 99.9% pure. The other thing is you need something that does not have any water vapor attached, no attached water. The TAs put this in the oven. So no hydrated water.

And then the final thing is, you want-- it should be stable at room temperature and when heated. And oftentimes we'll look at the cost, too. That's another factor that comes into play.

But what you're going to do is you're going to weigh out 0.0818 grams of this. And if you divide by the molecular weight of that, you'll get moles. And you're going to make a 100-mil solution. So divide that by 0.1 liters, and you'll get your concentration. You're actually making up something like a 0.0021 molar solution of that standard.

And then you're going to use that to standardize your thiosulfate. We're also going to be using starch in this reaction. So why do you think we need to use starch when we're titrating something? Alec.

AUDIENCE: Starch reacts with iodide, and it creates the blue color so that when we react-- when we were titrating, it could turn really pale.

JOHN DOLHUN: OK.

AUDIENCE: So it was really hard to tell if it was clear or not. So I think the starch helps indicate whether or not you've gotten rid of all of the iodide. Iodide, yeah.

JOHN DOLHUN: Good. Good. That's very good. So starch actually reacts with iodide and iodine, gives you that blue-black color. So if you're titrating something yellow to clear, sometimes it might be harder to see. So what Alec said was your solution actually reacts with starch. So I actually brought some of the solution in here, and I brought a piece of bread in. I don't know if this is going to work, but bread has starch in it, doesn't it?

So I'm going to put someone on this bread. Wow. Look at that blue black. So something's going on here. It's reacting with the starch, right? Can I have a volunteer? Some brave person?

I know you're all tired, but one of you. One of you, come up. Come on up, Maida. Maida, right? It's Maida, right?

AUDIENCE: Yes.

JOHN DOLHUN: OK. Maida, stand up front here. So what we're going to do is we're going to open this up. And put a pair of these on, Maida. And I'm going to let you hold this beaker and face your fellow students.

And just hold it nice and-- and what we're going to do is we're going to add something to it. It's clear, right? OK, keep your eye on the beaker. Maida, don't drop but whatever you do.

AUDIENCE: I won't drop it.

JOHN DOLHUN: OK. [CHUCKLES]

OK. So keep your eye on that beaker. Don't take your eyes off the beaker. Now what-- don't worry. Keep watching it, Maida. Don't get nervous.

Oh! Now that's what I'm talking about. You see that? This is what we're talking about.

And I mean, you're titrating. You can't see the end point. You put this in, now that last drop of titrant, it's going to turn it clear. You're going to be able to see your titration. And thank you very much, Maida.

So what's going on with the starch? Let's take a look at this. Starch is made up of about 25% amylose, which is the linear helical form. And it also is made up about 75% of the branched amylopectin. What happens is we have iodine and iodide present in our solution, and when those two come together, they actually form this pentaiodide anion.

So you've got some I2, some I minus. Remember, I2 is amber. I minus is clear. When you get these two together and they insert into this helix, the amylose helix of starch, and what amylose does is it forces the pentaiodide anion to go in linearly into that helix.

Then the energy spacings change. So you've got-- the way the wavelength of light hits that, you're going to see blue black. It's all happening inside of the starch with the amylose. That's the key thing.

There are different people-- some people still believe it's I3 minus. There's somebody else out there saying no, no, it's a polyiodide. There's always going to be some controversy in the-- but they're working on it. I like I5. And someone actually made an inorganic complex to make it look like starch, and they proved it was I5. But even after this paper came out, there's still a lot of controversy, so.

So for your standardization, pretty simple. You're going to go through, and you're going to just follow these steps. The TAs will go through this with you, very simple. And you're going to start off with probably something like this, kind of like a reddish solution. And then you're going to start adding your thiosulfate. And gradually, your titrating the iodine in this solution to iodide.

So here you've got a more yellow solution. What you want to do is you want to find a spot when it turns yellow to add your starch. And the starch has to be bubbly hot on the hot plate. And if you add it too early, there's so much iodine in there, it's going to destroy the starch complex. You won't get a reaction. But if you wait until it's so pale yellow, what happens, Thomas?

AUDIENCE: It turns blue. It makes the complex successfully, it turns it dark blue.

JOHN DOLHUN: It turns it dark blue. Yeah. But if you wait too long, right--

AUDIENCE: [? I ?] [? mean, ?] if you add too much, [INAUDIBLE]

JOHN DOLHUN: Did you-- you had a reaction yesterday where your starch was an globules, right? Is that--

AUDIENCE: Oh yeah. It formed a film because we didn't keep it hot.

JOHN DOLHUN: Oh, the starch was not hot enough. OK, that explains that.

AUDIENCE: [INAUDIBLE]

JOHN DOLHUN: Yeah.

AUDIENCE: We put it in and stirred, and it made a bunch of little blue specks that [INAUDIBLE].

JOHN DOLHUN: Yes, yes. Yeah, I've seen that before. So you want to keep your starch hot, and just be patient. And then you just take a plastic pasture pipette, take a swig of it, shoot it in, and you should get your blue black. And then look. You've got your blue black. You put that last drop of thiosulfate in, and you've got your clear solution.

So you turn around, and you're writing this down in your notebooks, right, and then you look back, but it's starting to turn blue black again. What should you do? What would you do?

AUDIENCE: Add another drop of titrant.

JOHN DOLHUN: Add another drop of titrant. That's what I would do, Thomas. But it's not correct.

AUDIENCE: [LAUGHS]

JOHN DOLHUN: But I would do the same thing. I would add another drop of titrant.

AUDIENCE: Oh, what if it just turned blue by itself, [INAUDIBLE] and you're just [INAUDIBLE]

JOHN DOLHUN: There's a side reaction going on in the air. When you get this clear, you've got all iodide present in there. But what happens is-- so you've got iodide there, but you also have oxygen in the air. And the oxygen is oxidizing the iodide to iodine, and we don't want this iodine coming from the air. We want it only from the river water, right, from the oxygen in the river.

So this is a side reaction that can go on if you let it set. So just ignore it. Just take your end point, your final neutralization, and you're good.

Take a look at the stoichiometry here. One biiodate make six I2. Each I2 reacts with two thiosulfate. So it's 12 to 1 thiosulfate to biiodate. What you're going to do is you're going to do three trials, and they all should agree to within 2% or 3%. If they don't, do an extra trial or two.

And what you're going to do is you're going to find your mean, your standard deviation, and your confidence intervals and then give all that information to your TA. Yesterday the TAs did a beautiful job. They put everybody's on an Excel spreadsheet, had it all averaged out. They really did a good job yesterday. The students really homed in on the exact concentration of the thiosulfate.

So it's pretty simple. You do your calculations right there in the lab. And then you have a choice of using the class average or using your own results for the dissolved oxygen that will take place on the next day.

So collection of water samples. So you're going to be going to the river for day two. And we have these poles. If you have long arms, you'll take a short poll. If you're short and have short arms, then we have real long poles. They're like hockey sticks.

So you get your stick, and what you do is you insert your-- your BOD bottle snaps into the clip that I've got inside. You take the stopper out, and then you go to the edge of the dock. Don't fall in, and don't push anybody, OK. Remember the cyanobacteria and all that out there.

And take gloves with you, because when you put this under water, it's got to be completely submerged. You have to reach over the dock. You've got to stopper it underwater. So you've got to have gloves with you, all right.

So I stoppered mine underwater. I bring it up, take it out, and then I look at the-- I take the bottle, and I do this test. I don't see any bubbles. That's a good sign. That means it's perfect.

If you see bubbles, you've introduced more oxygen in there. Your DO concentrations are going to be too high. Pour it out, and restart again. OK, so that's the collection.

Then you bring the stuff back to the lab, and you're going to treat the bottles in the lab. This here is wrong. You don't want to use these digital pipetters to treat the water, because you'll be introducing oxygen into those water bottles. What you're going to use is your 10-mL glass pipette.

And what we're going to do here is each pair of students will have four bottles. So you're going to-- first, you're going to treat it with the manganese sulfate, 2 mLs of manganese sulfate. That's the first reaction on that oxidation reduction Winkler series. And the way you do it is-- what I would do is I would take up 10 mils of manganese sulfate to prevent any airflow, any air.

And then what you're going to do is you're going to take the stopper out. You're going to go just below the surface of the liquid and put in 2 mils. And then your partner can stopper it. You go to the next one. You do all four of them.

The last 2 mils in here, shoot it into waste. And then do the same thing with the alkaline-iodide-azide reagent, 2 mils in each bottle. And you should look at the bottle at this point. After you add those reagents, there should not be any bubbles. If there's a bubble, it means you've introduced air, and your DO values will be higher by doing that.

So once you've treated with these two steps here, you've essentially trapped the oxygen. You're going to have this brown flock like this. Now you just shake that a bit, and then you're going to dissolve it now by adding sulfuric acid, 28 drops of sulfuric.

Acid for the sulfuric, you don't go below the surface. You just open the lid, and you drip, drip it in. The acid is heavy. It's going to fall right to the bottom of the bottle. And now you should have a bubble in there. You will have a bubble after you add the sulfuric, and that's OK.

So when you're done with this, you're ready to titrate. And we've got to-- we're supposed to titrate 200 mils. That's what you titrated when you did your standardization as well. But we've got to make up for the 4 mils that we added there in the beginning. We've displaced something there.

So you want to titrate 200 mLs times-- you've got a 300 mL BOD bottle, and you've taken and added 4 mLs of stuff to it. So if you do the math to make up for that, you really have to titrate 203 mLs. So there's a couple ways you could do this.

You can use a 100-mL graduated cylinder, and the error on that is about plus or minus 0.5. You fill that up twice, put that in your container. The last 3 mils, you could use this for accuracy to get your last 3 mLs in, and then you're good. But you can also just titrate 200 mLs and then multiply your answer by a correction factor, which would be 203 over 200. And that will give you the same answer as titrating the 203.

So the math is pretty simple here. You're converting the moles of your titrant to moles of oxygen, grams of oxygen. You divide by the liters. If you're doing 200 mLs, that's 0.2 liters. And then you get your milligrams per liter in PPM.

I kept mentioning we're using the-- we've got to add this alkaline-azide solution. We're actually using the azide modification of the Winkler method. And what is that? Well, there are NOx gases in the atmosphere.

What's the nastiest NOx gas that you can think of? I'll give you a hint. It's a big greenhouse gas, and it's not CO2.

AUDIENCE: Ozone?

JOHN DOLHUN: Ozone is a gas up there, yeah. But that protects us. Our ozone layer is a layer of protection. But what's-- yes.

AUDIENCE: Is it NO2 or--

JOHN DOLHUN: The other way.

AUDIENCE: Or N O? I know that [INAUDIBLE]

JOHN DOLHUN: All right, please don't laugh at this. It's laughing gas, N2O. You're all-- I mean, the atmosphere is loaded with laughing gas. I know. Everybody's laughing now. It's great.

Where does it come from, N2O? All the fertilizer. 1% of all the fertilizer in the world goes up in the atmosphere as laughing gas. It's not from a dentist's office. It's from fertilizer.

I know. I don't want to think about dentists. I was just at one yesterday. I don't like dentists.

But every ton of laughing gas that's in the atmosphere. is like 300 tons of CO2. And it stays up there for over 100 years. And what does it do? Well, who said ozone back there? There you are. OK.

Laughing gas reacts with ozone, and it forms dinitrogen dioxide. And this dinitrogen dioxide only lasts for about an hour. And immediately, the N2O2 reacts with air and water to form nitrites.

And here we are talking about greenhouse gas in the world, and this is affecting our little tiny reaction in the Charles River experiment, because these nitrites, look what they do. They get in and oxidize iodide to iodine. And we don't want our iodine coming from some reaction from laughing gas, right, from nitrite. We want it coming from the oxygen in the river.

So we add sodium azide. And that zaps the nitrites, convert them to harmless nitrogen and water. So that's the theory behind why we're using the azide modification. All that for these nitrites.

So let's talk for a moment about pH, pH in natural waters. Let's see here. Pulling down some board, here. Who knows the definition of pH? Remember your chemical principles from a long time a-- Alec.

AUDIENCE: I think it was log of concentration of hydrogen ions, [INAUDIBLE] ions.

JOHN DOLHUN: Very good. The negative log of the concentration of the hydrogen ions. So pH. And I mean, you all know the pH scale is 0 to 14, and acid is less than 7. And I hope you know what color litmus paper changes, right, with pH.

That was a question on a TV show with Regis Philbin,*You Want to Be a Millionaire.* A guy actually got up to the million dollar question. And I happened to turn it on just at that point, and they ask him, what color does litmus paper change in base? And the guy said, oh, I don't know, but can I call my lifeline? And they said yes.

So he called this prestigious biologist at one of the California universities. And they said red.

[LAUGHTER]

And the guy answers red, and he loses the million dollars. Oh, I was beside myself. So if you remember anything, remember what color litmus paper is in acid.

But in the river, the pH can go from 6.5 to 8.5. And the pH effects everything, from the solubility of metals-- if it gets too acidic out there in the river, the metals become more soluble, mercury, cadmium, arsenic, all those metals. So the fish uptake the metals, and we eat the fish. And with global warming, the acidity of the rivers is gradually drifting very slightly lower.

It also affects the forms of phosphorous. So remember, you've got PO4 to the 3 minus. But if it's slightly acidic, you might have the hydrogen phosphate or the dihydrogen phosphate. Or you could end up with phosphoric acid, depending on how acidic it gets.

What about photosynthesis and pH? Where are my biologists at, my resident bi-- there must be somebody in bio in here. Come on. Don't be afraid.

Yes. There you are. What's the equation for photosynthesis?

AUDIENCE: Water and carbon dioxide.

JOHN DOLHUN: Good, good. Water and carbon dioxide. Let's just write that down for a moment. So CO2 plus water, little bit of sun.

AUDIENCE: Oxygen and glucose.

JOHN DOLHUN: Oxygen and sugars, good. So CH2O, n. So does the pH go up or down during photosynthesis? I mean, it's like a 50/50 chance. Sean, you want to answer, don't you?

AUDIENCE: I would say the pH goes up.

JOHN DOLHUN: It goes up? OK. Anyone else? Alec?

AUDIENCE: It goes down?

JOHN DOLHUN: It goes down? OK.

[LAUGHTER]

There you go. I told you it's like a 50/50, right? All right, let's do a little experiment with this. I'm going to pour myself a drink here. This is good stuff. It's not Gatorade, I can tell you that.

This is a classic test in the medical profession, if you're going to medical school, for breath. And when I'm doing this, you cannot make me laugh. If you make me laugh, I'm dead, OK. Please.

I'm going to play some music to relax myself here. Let me just-- let me see if we can get this on here. OK. There we go.

So I'm breathing into this. We don't hear any music coming out. Going to--

[MUSIC PLAYING]

There we go. You're making me laugh, Sean. I may have to have you come up and help me with this.

[MUSIC PLAYING]

I may not be OK. The color's supposed to change.

[MUSIC PLAYING]

There we are. Good? I'm OK. So what just happened?

AUDIENCE: You blew carbon dioxide into the water.

JOHN DOLHUN: I blew carbon dioxide in the water. So we've got CO_2 plus water going to what? Yeah.

AUDIENCE: Carbonic acid.

JOHN DOLHUN: Carbonic acid. Good. H_2CO_3, which breaks down into H plus plus bicarbonate. Now photosynthesis needs carbon dioxide, right? So we take that out of the equation, which means we're also taking out the H plus. So does the pH go up or down?

AUDIENCE: Up.

JOHN DOLHUN: Alec? Up, up, up. OK? Good. And pollution and pH, same thing. If you've got pollution, you've got all this vegetation. So you've got photosynthesis, so the pH is going to be up.

This is Lassen Lake, Volcanic National Park, pH 2, 2.0. You do not want to stick your foot into this lake. And then you've got some safety of the chemicals, which the most serious is the azide reagent, which it's a neurological toxin. And if you ingest it, it can cause death. So please be careful with that.

Sulfuric acid, you know how bad that is. You don't want to get it in your respiratory system. And the manganese sulfate seems innocuous, but it attacks the central nervous system, targets blood and kidneys. So be careful with that.

Sodium thiosulfate is a respiratory irritant, can cause breathing problems. So if you feel like you're having breathing issues, it could be that. Biiodate can burn your eyes. Very, very dangerous stuff around your eyes. So all of these things, you have to be careful with. OK. So we'll see you Thursday for the last lecture in this series.