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PROFESSOR: So, you have now 10 seconds, go ahead and click in your answer. OK. We have a little bit of a mixed response here. So for those of you coming in late, I'm going to give you a second chance, and we say now what if the pK_a was 4. Everything else is the same, but the pK_a is now 4 -- what would the answer be? OK, go ahead and click in your response for a pK_a of 4, everything else being the same. OK, let's do 10 seconds. Interesting. More people got that right. So it's the same answer for a pK_a of 3 as a pK_a of 4, but it gave people the opportunity to think about that a little more and see what their neighbors had voted and make a decision based on that. And so more people came up with the right answer.

So let's consider now for a minute why this is true. So if you can switch to my lecture notes, I have a couple of slides on this that are not in your handout but that are related to things we've talked about. So just a brief sort of reminder of where we were with acid based titrations. So, we last time talked about what's happening at different points of the pH curve, and introduced this concept of 1/2 equivalence point where the pH equals the pK_a , where you have equal number of moles of molecules that are protonated as deprotonated. And another way to sort of think about this whole thing is that in the beginning at pH 's that are lower than the pK_a , you have more protonated molecules. And if pH is much above the pK_a , you have more deprotonated molecules.

So, we can think about things that way. So let's think about the last question that I asked where we have a pK_a of 4. So, if you gave someone a sample of molecules where the pH was equal to the pK_a of the molecule, you'd be giving them a sample where there would be equal amounts of protonated and deprotonated. But instead, if you gave a sample where the pH is much above the pK_a , then you're going to be giving them a sample that's mostly deprotonated.

And I emphasize about using, in terms of titrations, using Henderson Hasselbalch for buffer, but it also can be applied to thinking about this type of problem. And so you can be thinking about what sort of ratio are you going to get of your protonated to deprotonated when you have given a particular pK_a and a particular pH . And so, if the pH is really very far above the pK_a , then you're going to be largely deprotonated, and here's the math for these two numbers. If you're at a pH that's really below the pK_a , then your molecules are going to be protonated. And so, you're going to be thinking about this kind of thing when you get to organic chemistry, and also biochemistry. Biochemists spend a lot of time thinking about the pK_a 's of amino acid side chains, and there are often mechanisms that people propose of how an enzyme works, and they're proposing that some residue is going to

play a role as a catalytic acid or a base in an enzyme mechanism.

And often people are proposing things that really just don't make much sense in terms of the p k a of that molecule, that they're proposing, that it's going to be giving off a proton, but given the p h of the enzyme in the body and given the p k a, it wouldn't be protonated. So how is it going to give off something it's not going to have. So these are the kinds of questions people talk about in biochemistry, and if you're in a biochemistry seminar and someone's talking about something like this, there will probably be a hand go up in the audience and say "What do you think the p k a is of that amino acid that you're proposing that role in this enzyme mechanism?" And there have been people who are pretty high profile who've gotten themselves in trouble over p k a issues.

So now, you should all be ready to raise your hand in those seminars and say "What do you think the p k a is of that residue?" So, you'll be hearing about p k a's later, and the thing that makes me very excited is that some other people who are taking chemistry here, may not be as aware of the sort of biological role of p k a's. Now you know something, you've had some of these in your problem-sets, and so when you go into those advanced classes and there's a discussion of p k a, you'll raise your hand and impress my colleagues with your tremendous knowledge of p k a's. So I'm very excited about that. And please send me an email when you get extra points because of knowledge of p k a's. I want to collect that information and use it for evil purposes. No. But anyway, you can think about p k a's of molecules now, which will be very handy to you later on.

So it's not just -- some people tell me, if I promise never to titrate a weak acid with a strong base, do I have to take that part of the test? Well, you know it's not just about acid based titrations. Some of the things you learn in that are actually relevant to other things that most of you will see later in your career.

All right. So, I love enzymes, enzymes are great, I'm a biochemist. Acid base is very important in biochemistry. And the other thing that is very important in biochemistry is oxidation reduction. So, in these two units that we're doing now that'll be on the third exam, you're learning a lot of the basic principles that apply to how enzymes work. And so we started last time talking about some rules. Rules of assigning oxidation number. So this is sort of the very basic knowledge that you need to know to go on and do oxidation reduction problems. And here are those rules, and now let's see look at some examples of how we're going to apply these rules.

So, first let's look at a compound that has lithium and oxygen in it. All right. What do I know about lithium's oxidation number? What's it going to be? Plus 1. So, things in group 1 are going to be plus 1, and I have 2 of them. So we're going to be using one of those rules up there to assign it. What about oxygen? What's oxygen going to be in this molecule? Minus 2. So it's not in a peroxide here, so it'll be minus 2. And then we have plus 2 minus 2 is equal to 0, and that's good, because there's no charge on this particular molecule, so all of the oxidation numbers should add up to 0. That's also one of our rules.

All right, so let's look at PCl_5 . So, what we know about oxidation number of chloride? What are you guessing? Minus 1. What's the exception to that? When it's with what? When it's with oxygen it can be different. All right, so there's 5 of those. I haven't told you anything about phosphorous, but what can you guess it's going to be?

STUDENT: 5.

PROFESSOR: I love that enthusiasm, good. So, it will be 5, and so overall, we're going to add up to 0. So you won't always have a rule about everything in your molecule, but you'll be given something that you can get a handle on and then predict what the other thing is going to be. So let's do a couple more.

So, what about HNO_3 ? So let's start with oxygen -- you told me about oxygen, what's oxygen going to be here?

STUDENT: Negative 2.

PROFESSOR: Negative 2 and there's 3 of those. And we need to have this all add up to 0 again. So, what about hydrogen?

STUDENT: Plus 1.

PROFESSOR: Plus 1. What's the exception to that rule? When it's with a metal. So what does that leave for nitrogen? Plus 5. And that will all add up. All right, so you're very good at this, so let's bring in that clicker competition, and you can tell me about this molecule. All right, 10 seconds.

Excellent -- not necessarily for the competition since most people got it right, but I love seeing people get it right. So, 2 times plus 1 minus 2 is 0, so we know our states here. And I'll just tell you another thing you can do in doing these problems, if you recognize that something is often seen has a unit here, you can also think about this as H^+ plus and NO_3^- minus, and you would get the same answers to the problem. So, if it's a little complicated and you want to break it apart, you can do that too, you'll get the same answers there.

All right, so as I said, in this unit it's a lot of adding and subtracting, so this part is not too complicated, but you just have to be paying attention and doing your math, simple math, correctly. All right, so we looked at some examples. So now we're going to give some definitions. What is oxidation? What happens when something is oxidized? Yup. So, when you oxidize something, electrons are lost. What happens when you reduce something?

STUDENT: Gain electrons.

PROFESSOR: You gain electrons. Most people are good with those definitions. Here are some that give people a little bit more trouble. Oxidizing agent. So it is an agent of oxidation, so it accepts -- an oxidation agent is something that wants to oxidize something else, so it gets reduced itself. So it's an agent of oxidation. It runs

around trying to oxidize other things, but it itself is going to get reduced, and then you can probably figure out what a reducing agent is, it's an agent of reduction. And so it itself will be oxidized. It'll try to reduce something else. It'll run around trying to reduce something, trying to give off its electrons, donate its electrons so that it can be oxidized.

So, keep these definitions in mind, because you're going to be using them a lot in this unit.

So now we're going to use them, we're going to look at a reaction, and we're going to think about what's being oxidized and what is being reduced. And this particular type of reaction, a disproportionation reaction, the same element can be both oxidized and reduced. So let's break this down into two equations. In every reaction there's going to be an oxidation and a reduction, and so you can write those separately. And poor sodium, sodium is getting kind of a jip. A lot in these last few units in acid base, and here again, it's a spectator ion here, so it doesn't even make it into the equation. And it was ineffective as a conjugate acid, it hasn't really been doing very much recently. But that's OK.

All right, so let's look at what's going on here and try to figure out what's happening in terms of what's being reduced and what's being oxidized. So let's start over here, let's think about what the oxidation numbers are in this particular molecule. So what would be true about oxygen here and chloride. What do we know about this combination of things in terms of oxidation numbers? So let's start with chloride, what's that going to be here? So, we have a minus 2 here. And the whole thing is going to be minus 1, and so what is that about chloride? Plus 1, right.

So this is one of the exceptions. Chloride is usually minus 1, except when it's with oxygen, and here we have an overall charge of the molecule of minus 1, so it all has to add up to minus 1 and it does.

All right, so now let's do that side. So what's going on here? Are they going to be similar or different? Let's start with the oxygen, what's that going to be? So we have three negative 2's. It's not a peroxide, so it's negative 2 for oxygen. The overall has to be equal to minus 1, so what is chloride here? Plus 5, right. So that's unusual, but that's what it is. So we use our rules. Chloride is usually minus 1, except when it's with oxygen, and then it can be something different.

So, here chloride's going from plus 1 to plus 5, so what's happening to it? Is it being oxidized or reduced? Yup. So it's going from plus 1 to plus 5, so we have an oxidation going on. So to tell whether something is an oxidation or not, you need to figure out what the oxidation numbers are and then see what's changing in the course of that reaction.

All right, so down here, we've already done this one, so we can put that down here. So we have minus 2 for

oxygen, overall minus 1, and so chloride is plus 1 again. And what's the oxidation number on the other side for chlorine? Minus 1. I just have to look, and so that's minus 1. So here we're going from plus 1 to minus 1, so what is that? That's a reduction. And if we had figured out that they're both oxidations and something we would have done incorrectly, because in these reactions you're going to have oxidations and reductions. So, here, and this is disproportionation. So, you have chloride in one state, and in another, it's undergoing a reduction. OK, so that's how you sort of think about what's happening in these types of reactions.

So now we need to balance reactions. And this is very important in getting the correct answer to the later problems, and so we'll go through an example of how you're going to balance. You need to think about whether it's an acidic solution or basic solution, and we'll talk about that at the end.

So, first here, we can look at the two $1/2$ reactions going on, we have iron, we have chromium, and so we're going to look at those separately. So here is the first one, and let's think about what's happening over here. So, what would our oxidation number of the oxygen be here? What is happening to it? So, figure out what the oxidation numbers are of the chromium on one side, you know what it is on the other side, and then tell me what's happening to it. I hear some murmuring. People in one recitation aren't helping out people in the other recitation, are they? All right. Do you need more times? You good? You clicked in? Let's do 10 seconds.

OK. People did pretty well on that one. Let's look at the answer to that. So, let's go back to my Powerpoint. And so, oxidation number for oxygen here, minus 2. The overall charge on that is minus 2. So you have to figure out what the math equals there. And so, if you run through the math, then you see that you're going from a plus 6 to a plus 3 state, so we have a reduction. OK, 75% did that, got that, good.

So, most of you should be able to get this one now, iron plus 2 to iron plus 3. Just yell it out, what is that? Yeah. So here we have our oxidation.

Now let's balance this. And I left some nice blanks in your notes, but not so many that you won't be able to keep up and you should feel free to yell out the answers and we'll go through balancing this pretty quickly.

So there's a couple of different rules. I'll just say some books have things differently. If you can get the right answer, you can use whatever procedure you want. I have found in the past that this particular procedure a lot of people find to be the easiest. So I'll teach you this one, but you're free to use whichever ones work well for you.

All right, so the first thing we want to do is balance all elements that are not oxygen or hydrogen. To make it equal on both sides, we're going to do oxygen and hydrogen later. But what do we need to do up here to balance our non-oxygens? We need to add a what? A 2. So we need to add a 2 over here, two chromiums here, two on the other side. What about for iron? Nothing. All right, so that was pretty simple. Now, in this procedure you add water

to balance the oxygens. So, what are we going to do up here. How many waters do we need to add? So we need to add 7 waters. And the bottom one, bottom one's pretty easy so far, we don't have to do anything. So, go ahead and write in your 7 waters.

Next step is we're going to balance the hydrogens that we just added, and we can balance here with h plus, and I say here's one place that books are different. Some of them balance in the more sort of technically correct way with hydroium ions, but then your oxygens get unbalanced again, so it's OK to use the simpler approach, and just balance with h plus. So how many h plusses do we need to add to balance the top part? Yup, so we need to add 14 over here. Again, the bottom one, nothing to do. Pretty simple.

Now we need to balance the charge, so we just added h plusses, so we just added some charge to this, and now we have to balance the overall charge, so you want the charge on one side of the equation to be equal to the charge on the other side of the equation. So how many electrons are we going to have to add to this top equation to balance the charge? I heard the answer. 6. So we have to add 6. I told you, this unit involves adding and subtracting in your head, and you always want to check your work when you're doing this on an exam, because it's really sad when you lose points for something that is adding and subtracting. So you want to make sure that you don't lose points on these on an exam.

So what about the bottom -- finally we get to do something with the bottom one, what do we do? So, we add how many electrons? 1, right. OK, so now our charge is balanced.

So now, we want to multiply up one of the 1/2 reactions so that the electrons are going to cancel, and so what do we have to multiply by the bottom equation so that the electrons cancel with the first? 6. So we have 6, 6, and 6. And now, we're going to add those together and make the appropriate cancellations. So here is our overall equation -- we have the 6 electrons, the 14 protons, we have our chromium oxide compound, we have the 6 iron 2 plusses, on the other side, the 2 chromium 3 plusses, the 7 waters, the 6 iron 3 plusses, and the 6 electrons. So, we should be able to cancel the electrons, so we cancel those out. And now we want to double check that, in fact, it's balanced. Again, this is very important to do on the test, it's really easy to make some kind of math mistake, but you should be able to figure it out at this point. It won't be balanced if you've made a math mistake. So there should be 14 hydrogens over here, 14 over there, 2 chromiums, 2 chromiums, you have 7 oxygens here, 7 oxygens there, 6 irons, 6 irons. And the charge also should be balanced on both sides. So you can double check that and if it's good, then you're done. And this was in acidic solution, so we should have again this number and a charge of plus 24 on each side, double check, make sure it's correct.

So, that was acidic solution and we ended up with an equation that had h plusses in it. You can also be asked to do it in basic solution, and again, books have different approaches here. What I like to do is the simplest thing,

which is to use the same steps all the way up to here to get your same answer that you had for the acidic solution, and then neutralize it at this point. So, adjust the p h in quotes by adding hydroxide ions to both sides, and so, if we have 14 h plusses here, we can add 14 o h's on one side, and 14 o h minus on the other side. And then we can add those guys together. So we have now 14 waters over here. And we have still our 7 waters on this side, and then the 14 hydroxides on this side. And now, we should be able to cancel out some of those waters. So, we had 14 over here, 7 over here, so we can just have 7 on this side. And now we have an equation that looks like it's in basic solution. So instead of having h plus, which you would have in acid, you have hydroxide for basic solution.

So you can follow the same rules. Again, some books do it differently. I think this is the simplest way to do it, the least likely to make those kind of adding and subtracting errors. So those are sort of the fundamentals you need for this unit. You need to figure out oxidation numbers, looking at the composition of a molecule, and you need to be able to balance equations. When you can do that, you can go on to do other things. So, at this point -- oh, let me just say, those are the two answers there, so you can see acid and then in base.

So, at this point then, I want to try to do a little demo to show you that when you do do oxidation reduction reactions, cool things can happen. So, it's more exciting, perhaps, in real life when oxidation happens than on paper. So I'm going to turn it over now to Dr. Taylor and Dr. Patti Christie who's here to help us with this demo. And they will -- go for it.

PROFESSOR: OK, so this is mostly visual, there's not too much we'll need to say. Basically what we're going to do for you is oxidize magnesium. So we have a source of oxidation, which is going to be carbon dioxide, which is dry ice, and what we put inside the dry ice is solid magnesium. So we're going to set the magnesium on fire, and then put the oxidizing agent on top of it, and you can see the rest of what's going to happen there. And you can actually also determine if this is an exothermic or an endothermic reaction while you're at it.

[EXPERIMENTING]

[APPLAUSE]

PROFESSOR: Thanks to my helpers.

All right. So oxidation reduction reactions can be quite interesting, and don't try that at home.

So, we're going to continue on now, you know the basics and we're going to talk about electric chemical cells. We're going to introduce Faraday's law, and a thing I love to do, is come back to my friend Gibb's free energy. So, thermodynamics is all over the place in chemistry and you can't get very far away from it. So hopefully, by the end of the class today, we'll come back to free energy.

All right, so what is an electric chemical cell. It's any device in which electric current, which is a flow of electrons through a circuit, is either produced by a spontaneous reaction, or used to bring about a non-spontaneous reaction. And so a battery is technically a collection of cells in a series, so that the voltage that each cell produces is the sum, the battery has the sum of the voltages of each cell.

So, let's take a look at what some of these might look like. Here's a little cartoon of a simple version. We have a beaker -- 2 beakers with different solutions in them, 2 electrodes, a salt bridge across, and as electrons transfer through this, you can read a current on an amp meter here, you can read some kind of voltage coming off. And so, I like to talk about how good you guys have it at MIT with the Web and handouts and everything and how students in the old days at MIT, if they're going to do their problem-sets at night, first they had to build a battery to get electricity to be able to see. So you guys have it easy, you have these electric lights and all this fancy stuff now here.

So, let's sort of break apart sort of a components. Here is my beautiful picture that I drew of this system, that's why I show you a better cartoon first. So this is a beaker that you -- here's one beaker with one solution on this side, here is the other beaker. So we have two electrodes put in, we have a salt bridge, and then we have a wire across where we can measure voltage.

So, let's think about what is happening on each side. So in one beaker you're going to have an oxidation reduction, and in the other beaker you're going to have a reduction. So if we think about what's happening over here, you could have a zinc system, is oxidized from zinc 0 to zinc plus 2. And so, this is sort of a view of what would be happening at the electrode, you have zinc solid or zinc 0, and you've also have zinc plus 2 in solution. And so, as you take a zinc solid atom, going into zinc plus 2, you can have an oxidation reaction. And here would be the equation going down. So zinc solid to zinc plus 2 with two electrons.

Those electrons can go through to our other beaker on this other side, and so here we have a copper solid electrode or a cathode in this case, and we have copper plus 2 in solution, which is being reduced to copper solid, and so we're plating on to our electrode on this side. And so, here we have the reduction reaction, copper 2 with two electrons going to copper solid.

So our oxidation is happening at an electrode called the anode. The reduction is happening in an electrode called the cathode. And as these reactions occur, you're changing the charge on either side. So, you have this salt bridge, and so to neutralize the change in charge, you would have negative ions come down here, because we're producing more plus 2. And on the other side potassium is going in as we're going from copper plus 2 to copper solid, again, to balance this change in charge that's occurring. So those are the basic components of a simple electrochemical cell.

So, we could talk about the cell -- here's some nomenclature that represents the cell. So, the one that I just showed you, you have one electrode, a zinc solid, and we also have zinc plus 2 in solution. You have a single line between the solid and the zinc plus 2 ions in solution, that indicates there's a phase boundary, so you're changing phase from solid to aqueous. Then if you see two little lines that tells you, okay, one part of the reaction's in one beaker, the other part is in the other beaker, that represents the salt bridge which separates out the two $1/2$ reactions. And then on this side, we're going from copper 2 in solution, single line which indicates the phase boundary, to a copper solid over here.

So the amount of charge that goes through the system, it depends -- and how much of the zinc that is consumed or the copper that is deposited, is proportional to that charge, the number of electrons that go through the system, and this is called Faraday's law.

So let's look at what is happening a little bit here, so we have a little movie that shows the oxidation reaction. So in green is the electrode, and then here are the water molecules, and let's look at what happens when there's this oxidation. So here things are floating around, and there go the electrons. And this pops off into solutions, so if this is zinc, so then as this is happening, we have a zinc plus 2 now, it's aqueous. There go electrons, we have our oxidation. Here comes off zinc plus 2 floating around, there's another zinc plus 2 that just came off. And so our electrode here is being consumed as this oxidation occurs. I think it's really cool, the electrons, you can see them as this green going along in this movie. So, this is a little thing to think about what's happening.

So what's happening at the cathode, we have a reduction. So if we have our copper ions are going to plate onto our electrode. So here in blue is the electrode, and again the water molecules, here's a little copper 2 in solution. And when it gets electrons and gets reduced, it's going to start plating, so let's take a look at that.

So, here's -- oh, there comes the electrons. And so now it becomes part of the electrode there, the electrons come in, and so you're building up, you're adding, your plating on to your electrode. Now it got its electrons, it got reduced, got reduced again, and it becomes solid, and plates onto the electrode. And we see the electrons coming in.

So again, the amount of current that flows, the number of electrons that flow are going to be proportional to the chemistry that's happening at the two electrodes. So you can figure out how much zinc would be consumed, how much of the zinc solid electrode could be consumed, and how much copper would be deposited on the copper electrode for a particular amount of current. So say we have 1 amp of current for 1 hour, what are we going to do to our system? So we can calculate this out, so we use this equation to find out how much charge is going to pass through the system, we have this term q , the magnitude of charge and it's in Coulombs, those are our units, and that's equal to the current in amps, and just for unit conversion, amps are Coulombs per second, so that's very

convenient because we're going to take the current and times second, so times time. And so if we work that out, we had 1 amp and we had 1 hour, so we're going to have 3,600 Coulombs that are our magnitude of our charge.

So once we know that, we can convert, using Faraday's constant from charge into moles of electrons that had to pass through the system to generate that kind of current. And here is our friend Faraday's constant, and so that's in Coulombs per mole. And so, we can we can convert the number of Coulombs into the number of moles of electrons that were passed through the system with that amount of current, 1 amp for 1 hour.

Now, how is this going to relate to the zinc consumed or the copper deposited? So we know the total number of moles that passed through the system. And so, then we have to think about for every 1 mole of zinc that was consumed, how many moles of electrons went through the system? So why don't you yell out and tell me what you think that is. What do you think? 2, right. So we are going from zinc solid to zinc plus 2. So, we need to do -- to consume one, we need two electrons. So then we can look up the atomic weight of zinc and calculate the number of grams.

We can do the same with copper. So for one mole of copper deposited, how many moles of electrons needed to pass through the system? two, right. And then we look up the atomic weight, which is very similar, and actually with significant figures, it's the same amount, which won't be true for most things.

All right, so these problems are not too complicated. So let me just tell you a little bit more about types of electrodes that can be used. It's not always true that you have to use electrodes that are going to be consumed or have species deposited on them. You can use an inert electrode, such as a platinum electrode, and here is an example of that. So here, this cell has a platinum electrode, and the reaction that's happening, you have actually two things in solution instead of going from a solid to a solution. So this is just another example of a type of cell.

So let's think about, then, what equations we would write for this chemistry and what's going on. So over here at the cathode, what happens at a cathode, oxidation or reduction? I'm hearing somewhat of a consensus. That's correct. Reduction. That's the thing you're going to need to memorize, it'll be important in doing problems later on. So, what is the reduction reaction that you imagine would be happening if you have copper plus 2 and copper solid? So, if you're reducing it, you're going to be reducing copper plus 2 with two electrons to copper solid. So, same reaction we saw before.

So what about over here. This is the anode, which has the oxidation reaction going on, and what oxidation reaction can you imagine happening with chromium plus 3 and chromium plus 2. What's going to what? So, you're going from plus 2 to plus 3 with one electron here. So that would be the oxidation reaction.

So let's think about how we would actually write this down then. So the notation for this type of cell, we're going to

have a platinum that indicates the electrode, a single line that indicates a phase boundary. And in solution, you have the chromium plus 2, chromium plus 3, those are both aqueous, so they have a comma between them. So this is on one side in one of the beakers, so that's the reaction at the anode. And then we have a bar here that indicates a salt bridge. And then in the other beaker, we would have the copper plus 2 and the copper solid, separated by a single line, because there's a phase boundary between the aqueous and the solid. And then here are our two equations. At the anode, we have copper plus 2. Aqueous going to copper plus 3 aqueous in an electron at the cathode. We have copper plus 2 aqueous and two electrons going to copper solid. So those are our two equations and how we would write that down.

Another example of an electrode is a hydrogen electrode. And this is very common. In fact, most potential standard reduction potentials, or as they're also known, oxidation reduction potentials are measured against a standard hydrogen electrode, and it'll be abbreviated s h e. And so, if you see it was measured against s h e, you will now know what that means. So that's a standard hydrogen electrode.

So there's a couple of different variations. It can be used at the cathode or at the anode. And so, if you're using it at the cathode, h plus is reduced, and at the anode, h 2 is oxidized. And often you have a platinum system there as well, for the standard hydrogen, so platinum is commonly used in a standard hydrogen electrode.

So, let's just sort of look at a little picture of this. So here, we may have like a glass cylinder here, we're pumping in h 2 gas. And here you can see it sort of bubbling down. We may have a solution of hydrochloric acid, which would have a lot of h plus in it, and so this would be a hydrogen electrode on one side, the other side may have something more common or something that we saw before with the zinc solid and the zinc in solution.

So since on this side we have the cathode, of we can write about the h plus aqueous, a bar for the phase transition to h 2 gas, and then we indicate that there's also a platinum electrode going on here. And the other side is what we saw before, the zinc solid and the zinc plus 2. So you're just introducing different types of electrodes that you may be seeing in this particular unit. And again, you want to remember what reaction's going on at the cathode, and at the anode.

All right, so just very briefly, I will mention cell potential, which has many different names -- cell voltage, EMF it's often called. And so the flow of electrons generates a potential difference between the electrodes in the current. And this can be related back to our friend, delta g. So in this equation you have this potential difference that's generated by the flow of the electrons, and that times Faraday's constant times the number moles of electrons, and if you have a negative sign here, that's equal to the free energy of the cell. So we can relate that back.

And a little bit of information. We have the standard terms as well. We can have the delta e nought, the cell potential, cell voltage, you'll see many names for this. And that's when the products and reactants are in their

standard states. The units we're talking about here are in volts, sometimes you'll see things reported in millivolts as well. And let's end with one final clicker question, and then we can announce our clicker winner for today. OK, 10 seconds. Most people should get this right.