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CATHERINE So this is on oxidation--reduction go.

DRENNAN:

GUEST SPEAKER: OK, so the demo involves a strip of magnesium. And the magnesium is a pretty reactive metal. Right now it's coated with magnesium oxide. And what I'm going to do is I'm going to light it on fire. When it gets lit, it's going to glow very, very brightly, and it's going to react with this block of carbon dioxide.

And magnesium oxide is what you have up there. Both of those magnesium oxide molecules will react with one molecule of carbon dioxide to form just graphene, carbon, as well as magnesium dioxide. So you're going to get a white and black powder inside of here. And it's going to be really bright and really dangerous.

And can we lower the lights? Is that something we can do, because that makes it a whole lot more fun?

[SIDE CONVERSATION]

CATHERINE You can try it one more time, OK? All right. OK, good try.

DRENNAN:

So oxidation-reductions are really fun. And so we're going to continue and tell you about those on Friday.

So oxidation-reduction, or redox-- so redox is the shorthand for oxidation-reduction. Let's bring out your hands, but let's bring the noise level down. People have still been complaining about the noise in class. All right, so a redox reaction, a chemical reaction that involves an exchange of electrons. And so we saw what happens when we mix this with some dry ice.

All right, so let's do some definitions. We always start units with definitions. And most of you know some of these definitions already. Oxidation, electron loss. Reduction, electrons are

gained. This one you may or may not be familiar with. Oxidizing agent, it accepts electrons. So it oxidizes other things, but it itself is reduced. And reducing agent donates electrons. It reduces other things. It itself is oxidized.

And before you turn the page, I just want to try to get you to remember these two terms. So one way that I like to think about it is that a reducing agent might be somebody who is a little bit shady. They might walk up to someone and say, in a crowded space, don't look at me, just listen. Do you want to buy an electron? So that would be an agent of reduction going around.

Now, there could be ones that are a little bit shady, selling electrons to the highest bidder. But there could also be agents of good. I told you about free radical species. There's a lot of free radicals in the universe. They can be dangerous because they have a single, unpaired electron, which makes them really highly reactive. So there might be an agent, a secret agent of oxidation. And it might want to shield the universe from the power of free radicals. And so it will go along and it would oxidize that, take that electron away and oxidize and become reduced itself, take that free radical away. Because the free radical species might not know how dangerous it is. It's for everybody's good to shield you from that power.

You might also have an agent of reduction that's a secret agent, that might go up to the free radical species and give it an extra electron. And then it has two. And then it's more stable and less reactive. So you might have these agents.

So that's how you should think about it. Think about an agent of oxidation wants to go out and oxidize for the good of mankind, or an agent of reduction that wants to go out and reduce other things. So that might help you remember those definitions. We'll hope so.

All right, so let's take a look now-- we always start every unit with definitions, and then we have a set of rules and guidelines. So let's take a look at the guidelines. Luckily, these guidelines are really pretty straightforward. And you really shouldn't have a lot of difficulty keeping them in mind. Some of them we've really seen already, we just didn't realize that they were going to be the guidelines that we would have.

So guidelines for assigning oxidation numbers-- in a free element, every atom has an oxidation number of zero. OK, that seems pretty straightforward. H₂, oxidation number, zero.

Rule two-- ions composed of only one atom, also called monatomic ions, have an oxidation number that's equal to their charge. If this wasn't true, it would be deeply disturbing. Lithium

plus 1 ions has an oxidation number of plus 1. I think we're good.

Also in number two, in forming compounds, compounds Group 1 metals have an oxidation number of plus 1, lithium and all of its friends, sodium, et cetera. Group 2 metals have an oxidation number of plus 2-- calcium for example. And aluminum is plus 3 in its compounds.

This is also something that we've seen before. Those Group 1 and Group 2 metals are not all that exciting. They pretty much want to get their noble gas configuration and they want it to lose either one, if you're in Group 1, or two electrons if you're in Group 2. And so that's what they do here. So this is good to keep in mind because it's going to help you figure out the oxidation number of other elements for which there are not such hard and fast rules.

Number three, oxygen-- oxygen in most compounds is minus 2. Again, not that surprising that it would want to gain those two electrons. However, there's some exceptions. Oxygen is just an element full of exceptions. If it is in peroxides-- hydrogen peroxide, O_2 , 2 minus-- then it's going to have an oxidation number of minus 1. So peroxides, and later we'll see superoxides are exceptions. But if it's not a peroxide or a superoxide, it's going to be minus 2.

All right, rule number four, hydrogen-- hydrogen is mostly plus 1. You probably could guess that. But there actually are some exceptions for this. And you're going to see a lot of those exceptions in this unit, and pretty much after this unit, not so much ever again. The exceptions are for the Group 1 and the Group 2. So the Group 1 and Group 2 rules supersede the fact that hydrogen is usually plus 1. So if you have hydrogen with lithium, lithium is going to be plus 1. So hydrogen will be minus one. Sodium is the same here. And calcium, that's Group 2. Again, hydrogen would be minus 1. So Group 1, Group 2 rules supersede hydrogen.

All right, what about our halides? We have fluorine. It wants to be minus 1. Other halogens-- chloride, bromide, iodide-- are negative when they are in salts, such as sodium chloride. But if oxygen is there, the oxygen rules are the ones that follow. So here, if you have Cl with oxygen, then you're going to have a positive oxidation number for the chlorine, because oxygen is going to be minus two. And you'll see a number of these so-called oxyacids in this unit.

All right, six-- this one, you probably already know. But in a neutral molecule, the sums of the oxidation numbers must equal zero. So the sums must be neutral if the molecule is neutral. And if you have something that is a polyatomic multiatom ion, the sum of the individual oxidation numbers needs to be equal to the overall charge on that ion. So again, we've seen things like this before. So we should have no trouble with these words.

OK, so let's just look at an example of this. Suppose we have our friend NH_4^+ . What is hydrogen's oxidation number going to be in this molecule? It will be plus one. So it's not with a Group 1 or Group 2 element, so it's going to be plus 1.

What about nitrogen? There were no rules about nitrogen, so we have to figure out based on the other rules what the oxidation number of nitrogen will be. What is it going to be here?

AUDIENCE: Minus 3.

CATHERINE DRENNAN: It's going to be minus 3. And the reason for that is because of the rule we just talked about, that the sum of the oxidation numbers needs to be equal to the charge on the molecule. The charge is plus 1. So we have four hydrogens, $4 \times +1$ minus 3 is going to equal plus 1. So nitrogen would have to be minus 3. Nitrogen can have a lot of different numbers, so you want to pay attention to what's around it and whether there are rules or guidelines about those other elements.

All right, seven, oxidation numbers don't have to be integers. They almost always are, but there is one notable exception, and it's from oxygen. Perhaps not a surprise, oxygen is always causing trouble.

So in superoxide, you have O_2^- . So what is the oxidation number on each oxygen in that O_2^- molecule? Yep, so it's minus $1/2$. So oxygen is usually minus 2 except in peroxides and superoxides.

All right, so let's take a look at some examples. And I'll write some of these on the board. And here are all of your rules. So you can yell out what you think I should be doing with these as we go along.

OK, so we have lithium 2 oxide. So here we know a lot about this. What's the oxygen going to be-- minus 2. The overall charge on the molecule is zero. And we have two lithiums. And what's the charge on each one?

AUDIENCE: Plus 1.

CATHERINE DRENNAN: Plus 1. And so $2 \times +1$ minus 2 is 0. Great.

Let's do another one. We'll do PCl_5 next. So what do we know about Cl? What charge do we

expect here?

AUDIENCE: Minus 1.

CATHERINE DRENNAN: Minus 1, and there are five of those. So chlorine likes to be minus 1. Exception is with oxygen, but it's not with oxygen here. The overall charge is going to be zero, because it's a neutral molecule. So what does that mean about the charge on our phosphorous?

AUDIENCE: Plus 5.

CATHERINE DRENNAN: Plus 5, right.

OK, so now let's try HNO_3 . So let's do oxygen first. What charge do we expect on oxygen?

AUDIENCE: Minus 2.

CATHERINE DRENNAN: Minus 2. So we have three of those minus 2. Overall this is going to be zero. Nitrogen can be varied, so let's do hydrogen. What do we expect hydrogen to be?

AUDIENCE: Plus 1.

CATHERINE DRENNAN: Plus 1, right, because it's not in Group 1 or Group 2. So what does that leave for nitrogen?

AUDIENCE: Plus 5.

CATHERINE DRENNAN: Plus 5. So we have plus 6 then minus 6 is 0.

All right, so why don't you try the next one as a clicker question. All right, 10 more seconds.

So oxygen is minus 2. It's not in a peroxide or a superoxide. The overall charge is zero. So that means that each nitrogen must be plus 1. So we have plus 2 minus 2 is 0.

AUDIENCE: Professor?

CATHERINE DRENNAN: Yeah?

AUDIENCE: What exactly is a peroxide?

CATHERINE

A peroxide-- so there really only goes the hydrogen peroxide example. So it'll say that it's a

DRENNAN:

peroxide, or a superoxide. You don't have to worry about that. It's going to be clear. OK, good.

So now let's talk about some reactions that involve oxidations and reductions. And to talk about these reactions you need to know how to do this. And then we're going to talk about balancing reactions. And you need all of these skills.

All right, so first I just want to tell you about something called the disproportionation reaction where you have one element that is going to be both oxidized and reduced. And this is a good example to allow us to think about what's happening in a different reaction.

So in this reaction, the element that is going to be both oxidized and reduced is Cl. Sodium is going to be a spectator ion. And the oxygen is going to stay in its minus 2 form here. And so we have chlorine going to chlorine in this compound and in this compound. So we can break this down and think about what's happening with this species going to this, and then also going to Cl minus.

So let's think about what the oxidation number of chlorine is in ClO minus. So first, oxygen, I already told you, is going to be minus 2. The overall charge on the molecule is minus 1. So chlorine has to make this equation work. So what is the oxidation number of Cl?

AUDIENCE:

Plus 1.

AUDIENCE:

Plus 1, thank you. So yeah, plus 1 here.

Now let's take a look at this. So oxygen, we're going to have as minus 2. There are three of them. That's going to equal, overall, minus 1. So what's the oxidation number of chlorine here?

AUDIENCE:

Plus 5.

CATHERINE

Plus 5, right. So we have a case now where we're going from plus 1 to plus 5. Is that an oxidation or a reduction? That's an oxidation. So we're increasing the number. It is losing electrons. It's becoming oxidized.

DRENNAN:

So down here, we already have done all of our work and know that this is also going to be plus 1. So in this case, we're going from plus 1 to minus 1. Is that an oxidation or a reduction?

AUDIENCE: Reduction.

CATHERINE DRENNAN: That's a reduction, and that's good because it wouldn't be a very good disproportionation reaction otherwise. Something has to be oxidized. Same thing has to be oxidized and reduced. So this is how you think about this type of reaction. So to consider whether something's being oxidized or reduced, you need to be able to dissect out what the oxidation numbers are.

All right, now we're going to balance. So balancing is going to become very important for doing problems later on, balancing redox reactions. You balance differently if you're in acidic conditions versus basic conditions. So first we're going to do acidic conditions. And we're going to run through the steps for doing this.

First step is we're going to write two half reactions for this. So we have iron in here and we have chromium in here, and we're going to separate those out. So we're going to think about what iron is doing separately from what the chromium is doing.

So first, let's write out our chromium and think about what is happening to it here. And that is a clicker question.

All right, 10 more seconds.

OK, so let's take a look at why two is correct. For one thing, you just have to remember the definitions. For another, you have to recognize what's happening to the chromium.

So let's take a look at the chromium now. So we have oxygen at minus 2. There's seven of them. And the overall charge on the molecule is minus 2. And so what does that mean for the oxidation number of this chromium?

AUDIENCE: Plus 6.

CATHERINE DRENNAN: It's going to be plus 6, because we're going to have 12 minus 14 equals minus 2. So we're going from plus 6 over here to plus 3 here. So is that an oxidation or reduction? So that's a reduction.

So I kind of like this unit coming after acids and bases because a lot of those problems, I think, make people's brains hurt a little bit. And you get here and it's, like, basic math. So it's kind of exciting.

All right, so the iron is even simpler. We're going from plus 2 to plus 3. So that's an oxidation.

So one thing is being reduced. The other thing is being oxidized. That's what makes it a redox reaction.

All right, so we know what's happening in terms of what's being oxidized, what's being reduced. And we've broken these into half reactions. Now we need to balance.

All right, so the first thing we're going to do is we're going to balance our elements that are not oxygen and hydrogen. So we're going to balance chromium and we're going to balance our iron.

What do we need to add to this top to balance chromium? We need to add a 2. And we don't need to do anything to our iron. We have one iron on one side, one iron on the other side. All right, so all you need to do is insert the 2.

Next step, we're going to add water to balance oxygen. So how much water do we need to add to the top reaction to balance the oxygen? So here we're going to add seven waters to balance the oxygen. And again, we have nothing to do down there.

All, right, so let's keep going. Next, we want to balance the hydrogen. We're just added those waters. And so now we have to balance hydrogen. So we're going to balance the hydrogen by adding H plus. Some books have you balance the hydrogen by adding a more correct species, which is hydronium ion. But when you do that, you screw up your oxygens that you just balanced. So that is infinitely harder to balance equations that way. So I say let's not do that. So let's just use H plus.

So how many H pluses do we need to add over here to balance the hydrogen? 14-- you've added 14. And again, you don't have to do anything to the bottom expression.

Next, we are going to balance the charge by inserting electrons. So over here we have 14 minus 2 plus 12. And over here we have plus 6. So how many electrons are we going to need to add? We're going to need to add six. So we've added six over here, because they have a negative charge. So now we should be balanced on both sides.

And now we finally get to do something with the bottom expression. And we can add an electron over here to balance the charge. So plus 2 here, plus 2 net on the other side.

All right, now we want to multiply so that the electrons are going to cancel. Because this is an overall oxidation-reduction reaction, our electrons should cancel. So what do we need to

multiply the bottom equation by 6, and there we go-- 6 here, 6 here, and 6 here.

Now this is one of the places that people often, in exams, make a mistake where they write 6 and then put a parentheses around it indicating that you're multiplying the whole thing by 6, which is fine. But then they forget to actually carry that out on the next step. So when you multiply, make sure that you continue to-- sometimes it's better to just write it all out, because otherwise it's really easy to forget and not multiply, and then things won't make sense at the end.

All right, so now we can add the half reactions and make the appropriate cancellations. So we've now added all together six electrons, 14 H plus. We have our chromium oxide compound. We have now our six iron plus 2s, our two chromium plus 3s, our seven waters, our six iron plus 3s, and our six electrons. And if we did everything right, then the electrons should cancel. And we can double check. It's always a good idea to double check that it's balanced.

So we have 14 H's here, 14 H's there. Chromium two, two. Oxygen seven, seven. Iron, six, six. This can catch a mistake that you might have made. And then check the charge-- plus 14, minus 2, plus 12 on one side. Plus 6, plus 18, so 24 on both sides. That's pretty good, right? That's pretty simple. So balancing.

That's an acidic solution. We know it's in acidic because we have 14 H plus. That feels quite acidic.

What do you do for a basic solution now? So for a basic solution what I recommend is just do the exact same thing that you did for your acidic solution. And here again was the answer that we have. And then, quote, neutralize the acid and make it basic. Adjust your pH by adding OH minus or hydroxide ion concentration to each side.

So if we do that now, we're going to add 14 OH minus to this side and 14 OH minus to that side. Now 14 OH minus plus 14 H plus are going to give us 14 water molecules over here. And on this side we still have 14 OH.

This is not simplified, because we have 14 waters on one side and seven on the other side. So we can take that, cancel out these seven. And we have seven left over here. So we have seven waters now on one side and 14 hydroxides on the other.

So if we look at both of these expressions in acidic solution and in basic solution here, in acidic

solution we have 14 H plus on one side and seven waters on the other. In basic we have seven waters on this side and 14 OH minus on this side. Everything else should be the same, and they should all add up.

All right, so there we go. That is balancing oxidation-reductions. It seems really straightforward, but it's very important. We'll see next week that we need to use this information to be able to do problems.

All right, so end of 14. Now to electrochemistry-- so today's handout.

So we're talking about oxidation-reduction. We're continuing talking about oxidation-reduction. I just want to remind you how really important oxidation-reduction reactions are.

So we've talked about photosynthesis in this class. We've talked about fuel cells. All of these reactions that make energy involve oxidation-reduction, loss of electrons, gain of electrons. Basically the human body is a big battery. We have lots of oxidation-reduction reactions going on inside of us. This is how we live. This is how we make energy. This is how we survive.

All right, so electrochemistry is the study of redox reactions, usually thinking about how they occur at an electrode. So there are two types of reactions that you will be considering. You might be considering a reaction that generates an electric current, electron flow, or electricity, from a spontaneous reaction. And you might consider one where you need to drive a reaction that is nonspontaneous using an electric current.

So a spontaneous reaction, delta G is what, positive or negative?

AUDIENCE: Negative.

CATHERINE DRENNAN: Right. So a spontaneous reaction, delta G is negative. It's less than zero. A nonspontaneous reaction is positive. Just a little review there for exam three.

All right, so electrochemistry, study of these redox reactions. Electrochemical cells are devices where you generate electrons and they flow through the cell. And so you have the flow of electrons through a circuit. And again, that can be generated by a spontaneous reaction. And if it is generated by a spontaneous reaction, that's called a galvanic cell. Or you can use an electric current to force a nonspontaneous reaction to go. And that's called an electrolytic cell.

And it's important to remember those definitions because a lot of problems tell you, oh, in this

electrolytic cell. And that's telling you key information about what's going on in that problem. So if you don't have these committed to memories, you will find yourself in trouble. But it's not too bad, just two terms.

So a battery then-- so I have a battery here. A battery is a collection of galvanic cells. And it's going to give us-- it's going to use spontaneous reactions to give us electric current.

And again, electrodes, the definition for those, they're just the conductors through which the electrons travel. And there are two types. Let's quiet down again. I hear the noise. Two types of electrodes-- we have anodes and cathodes.

All right, so let's look at an electrochemical cell and consider, again, the anodes and the cathodes. So this is my beautiful depiction of an electrochemical cell. This might be a better depiction of what it looks like, but I think mine is not bad. Just to help you out, that's supposed to be a beaker that has fluid in it. This is also a beaker that has fluid in it. And this is my electrode. That is an anode in this case. There is the anode there, the electrode there. Here's another one. That's the cathode, another type of electrode. It's over here. This is my salt bridge. The salt bridge here is this plastic tube. And here I have a wire connecting my electrodes here. And you can't really see the wire. But in both cases it goes through some kind of volt meter that measures the electric current.

All right, so now we get rid of this better picture. But you can think about this pictures or that picture if that helps. And I always like to joke that you think you have a bad here at MIT, in the old days, before you could do your problem set at 2 o'clock in the morning you had to build your own batteries to be able to do it.

All right, so anode-- electrons are produced from oxidation at the anode. So the electrons produced here flow out of the cell. And in this particular type of electrochemical cell, zinc solid is oxidized to zinc plus 2.

So if we look at this over here, this is my little beautiful depiction of the blow up of the electrode. We have our zinc solid here. And we have zinc plus 2 in aqueous solution. So when the electrons leave the zinc solid, so that the zinc then leaves the electrode and goes into solution, forming zinc plus 2, those electrons go out.

So the electrons will then go into the cathode. And notice I had a typo in the handout. I find this pretty funny. I don't know if it's me or spellcheck right now. I wrote elections are consumed. Oh

man, it's been a long week, between the national elections and our T-shirt elections. Anyway, I didn't mean elections. I meant electrons. Yes, of course, electrons are consumed by a reduction reaction. Of course I noticed that after I had printed hundreds of copies.

And this particular reaction here, copper plus 2 is reduced to copper solid. And if we look over here, we have copper plus 2 in aqueous solution. When two electrons come in, those two electrons go to the copper plus 2 and reduce it to copper solid. And the copper solid actually plates on the electrode.

So we can measure the flow. We can have a meter that measures the volts or the flow of electrons, the electric current.

And also what's that salt bridge doing? Well, the salt bridge can help maintain the neutrality in the system. Because over here, we're generating plus charge. And over here we're losing plus charge. And so ions can flow through the salt bridge and keep the charges more similar in our two different beakers.

All right, so I know my pictures are gorgeous, but I have a little animation to show you that will also show you what I was talking about, oxidation at the electrode. Zinc solid is going to lose electrons and form zinc plus 2. And let me just show you that happening. So there are the electrons going away. And here is zinc popping off as zinc plus 2. And then the zinc plus 2 is going to be in solution. So there the electrons go again. And the zinc solid pops off as zinc plus 2.

All right, so now let's consider reduction at the cathode. So here we have our copper solid. Here we have a copper plus 2. And here the electrons are going to come in and join up with our copper plus 2. And it's going to plate onto the electrode.

There come the electrons. There they come again. And now we have added solid copper to our electrode. And just one more time-- and I think it's so cool that we can just watch those electrons move like that.

All right, so there's our electrochemical cell where we have an oxidation and reduction reaction. Instead of drawing this admittedly beautiful picture every time, we can have a shorthand representation for this. So we can say zinc solid. And then we have a little line here. That little line indicates we have a phase boundary. So that means we're going from solid to a different phase. And in this case we're going to aqueous, so solid zinc to zinc plus 2.

The two lines here indicate the salt bridge. So this tells us that the reaction on this side is in one beaker, and the reaction on this side is in our other beaker. The reaction in the other side is copper plus 2 aqueous, a single line for phase boundary going to copper solid. Now if we had, say, two aqueous species, we wouldn't have a phase boundary, we'd just have a comma. But if you have a phase boundary, you get the line.

All right, so it's a clicker competition. Let's consider another cell. It still has zinc, but now it has tin. Why don't you tell me which reaction is at the anode and which reaction is at the cathode? And also, which of those is oxidation and which is reductions?

All right, we'll just do 10 more seconds.

Great. OK, so some of these had parts of it right, but there was a mistake. We have AnOx, so that is wrong. So you had to pay attention to whether these match, and then also that the reaction is written consistent with the word. So only number two had all the various different parts right.

So we can come back here and you can enter-- so the reaction at the anode was a zinc reaction. And that is an oxidation, AnOx. The reaction at the cathode is the tin reaction. Tin plus 4, two electrons, tin plus 2 aqueous. And that's a reduction.

And while you have your clickers out, why don't you tell me how you would represent that electrochemical cell? All right, let's just do 10 more seconds.

So right. So I mentioned this very briefly. This was my tricky clicker question, because it's a competition. So there's no phase transition here, so you have a comma. And the top one a few people put down, but it's not solid here. So that one is not correct.

So here you can write down in your notes, you have this zinc phase transition, zinc plus 2 aqueous. And then we have our salt bridge, zinc plus 4 aqueous comma per zinc, tin plus 4 comma, tine plus 2.

And now we're ready for winners. We'll see if Sam's also ready to present the winners.

All right, we need a strategy to defeat Dan's recitation, but we know what T-shirt we're playing for now. All right, have a great weekend everybody.

All right, let's just take 10 more seconds. All right, does someone want to tell me how they got

the right answer? Yep?

AUDIENCE: You know that your hydrogen will have a plus 1 charge, and your oxygen will have a minus 2 charge because it's not a peroxide. And you know your sum is going to be neutral. So you add one.

CATHERINE DRENNAN: Yep, that's right. So this one, chlorine is usually minus 1. The exception is when it's with oxygen. And so oxygen is always minus 2 unless it's in a peroxide or superoxide. And hydrogen's always plus 1 except when it's with a metal like sodium, something Group 1 or Group 2.

OK, so today take out lecture 25 notes. We're just going to wrap those notes up-- that shouldn't take too long-- and then keep going with lecture 26. So we were talking about this cell. And we're going to talk about this electrochemical cell a lot.

So we had zinc on one side, zinc at the anode. And the zinc was being oxidized, generating electrons, which come around to the copper. And then the copper plus 2 ions were being reduced to copper solid. So the zinc anode was being consumed. And the copper cathode had copper solid being plated onto it.

So Faraday's law says that the amount in which an anode, like the zinc anode, is consumed and the amount of copper that gets plated on at the cathode is proportional to the charge that passes through the system. So according to Faraday's law, there is a proportionality between the consumption and also depositing of metal and the amount of charge going through the system.

So we can think about how much zinc is consumed and how much copper is deposited on its cathode or its electrode if you have a current of 1 amp flowing for 1 hour. So let's just take a look at how we do a problem like this. First step is we want to think about, calculate the charge that's going through the system. And so Q here, we have another Q -- not the reaction quotient, but the magnitude of the charge. And that's in coulombs. And that's going to be equal to the current times the time.

So we have the current, I . And the current is in amps. And t is in seconds. And conveniently, amps equal coulombs per second. So this equation works out. Our units will work out.

So Q is going to be equal to-- we have 1.0 amps. And it was 1 hour, so that's 3,600 seconds, or 3,600 coulombs, because we have an amp times second. And then here, we have this

number of coulombs. And now we want to convert to the number of moles of electrons that that is equivalent to, that that charge in coulombs is equivalent to.

And here, our Faraday's constant is our conversion factor. So Faraday's constant, we have 96,485 coulombs per mole is one faraday. And its this kind of funny looking F over here.

So we have our coulombs. So we can use now the Faraday's constant to convert. And we get 0.00373 moles of electrons. So that's the number of moles of electrons that would be going through our system that correspond to that amount of charge.

All right, so now we need to relate this back to the zinc and the copper. So step three, we want to know first the number of moles of the zinc and copper involved, and then convert that to grams because the problem asked us for grams.

So we have the number of moles of electrons that are going through the system. And so for every one mole of zinc consumed, how many electrons go into the system? What's the ratio, one, two-- how many electrons?

AUDIENCE: [INAUDIBLE]

CATHERINE DRENNAN: So one zinc solid going to one zinc plus.

AUDIENCE: Two

CATHERINE DRENNAN: Two. So that's going to be two moles of electrons. And then we need to look up, how are we going to convert our moles to grams? What are we looking up?

AUDIENCE: Molar mass.

CATHERINE DRENNAN: Yep, molar mass, molecular weight, atomic mass. So in our periodic table, grams per mole. And so we get 1.2 grams here for this equation.

All right, so for copper, so we have the same number of moles of electrons that are going through the system. So for every one mole of copper solid deposited on its electrode, how many moles of electrons have to come into the system? Louder?

AUDIENCE: Two.

CATHERINE DRENNAN: Two, yes. So that's two as well. And then we look up again, in the periodic table, the atomic weight, molecular weight. And it's almost the same, copper and zinc. So to the number of significant figures it's the same amount of grams. Now this will not always be the case. They're just very close to each other in their molecular or atomic weight.

OK, so we have two significant figures. Also our current in time only had two significant figures. So we only have two significant figures in this problem.

All right, so in this particular electrochemical cell, again, our anode is being consumed and the electrons that were generated ended up plating the copper on. Sometimes you're not going to want to do that. Sometimes you're just going to want to actually plate something with copper, for example. And so let's look at some of the uses for this kind of chemistry. So here's some electroplating.

[VIDEO PLAYBACK]

- This steel spoon will have a thin coating of copper deposited on it. The cell consists of an anode of pure copper, a solution of copper sulfate, and as a cathode, a spoon. There is no net reaction. The copper metal is transferred from the anode to the cathode.

CATHERINE DRENNAN: It's pretty nice.

[END PLAYBACK]

CATHERINE DRENNAN: OK, so is kind of electroplating is pretty common where you use these kind of reactions to get your copper, which is in that solution, to plate as copper solid onto something. So it could be a spoon. It could be an electrode. Yeah?

AUDIENCE: Why is it dangerous?

CATHERINE DRENNAN: Why is it dangerous? You know, I've seen these demos before, so I think that's just like security [INAUDIBLE] to say that. I don't think it's incredibly dangerous frankly, but I probably shouldn't say that. I'll get myself in all sorts of trouble. Don't try this at home. I don't know why.

AUDIENCE: Usually when you're doing jewelery plating, the chemicals that you use are relatively toxic.

CATHERINE DRENNAN: Yeah, but not that much worse than the lab stuff. Allena makes jewelry, so she knows all about this. But she does it very safely-- very safely.

All right, so let's look at a couple other kinds of electrodes. It is not always the case that an electrode is consumed or that an electrode gets stuff deposited on it. Some electrodes are inert. An example for this is a platinum electrode. It is inert. So it doesn't get consumed and it doesn't get reacted onto during the reaction.

And so here's an example of a cell that uses this. So over here now we have still our copper system. So this is our cathode, again, over here. And a cathode does what, oxidation or reduction?

AUDIENCE: Reduction.

CATHERINE DRENNAN: There's a reduction. And so this is our same reaction, copper plus 2, two electrodes, coated copper solid. So this is doing the plating again. But this time it's different. Before we have the zinc on an electrode that was being consumed. But now we have this platinum. And the reaction that happens at the anode-- it's still an oxidation, always an oxidation at the anode. But the oxidation reaction is chromium plus 2 going to chromium plus 3 and an electron. So here, the electrode is inert, and the reaction is all happening in the same phase, all happening in solution.

So we can also write the notation for this kind of cell. And so here we would have platinum solid, the single line, which means the phase difference. Chromium plus 2 aqueous comma, because now there's not a phase transition between these, a chromium plus 3 aqueous. The two lines, that means what? What do those mean again?

AUDIENCE: Salt bridge.

CATHERINE DRENNAN: Salt bridge. So this indicates that's one beaker. This is the other beaker. And here we have the copper plus 2 aqueous, single line for phase change and a copper solid. So again, our anode reaction and our cathode reaction.

All right, so this is another kind. You can also use platinum electrodes in combination with what's known as a hydrogen electrode. So these are often together, the platinum with the hydrogen. And many reduction potentials-- and when you get problem set eight, you're going to be looking up in your book a lot of standard reduction potentials. And many of them will say, measured against SHE. And a lot of people say, who is this SHE that measures all of these potentials in the back of books? And that SHE is a standard hydrogen electrode.

All right, so let's think about how this electrode is being used. So when it's used as a cathode, we have H plus aqueous phase transition to H₂ gas, and then another phase transition to our solid platinum inert electrode. And here when it's used as a cathode, H plus is being reduced. But it can also be used as an anode.

And this is a clicker question. Why didn't you tell me what would be happening when the hydrogen platinum electrode is being used as an anode, or reaction's being used as an anode? 10 Seconds. Let's see if you can get 90%-- 89, 89%. Oh, far away. That is the right answer though.

So some of these, platinum is inert, so it's not doing anything. And it's over here. So it's not doing anything. So we can rule out all the things that platinum are doing stuff.

All right, H plus being oxidized-- you ever heard of H₂ plus? No, so we don't want that. We know at the anode it should be oxidized, so we can eliminate all the ones that are reduced. OK, so we only have really kind of one good option if you say that the platinum is inert. So H₂ then gets oxidized in this case, forming H plus.

All right, so let's look at a cell like this that uses the hydrogen electrode. So here's an example. Now on this side we have our regular zinc system for our anode. And so over here we're using, as the cathode, the hydrogen electrode. And so here in this beautiful drawing that I made, we have a glass tube here where you're pumping H₂ gas in at 1 atmosphere. In the middle of the tube we have our inert platinum electrode. And then these are H₂ gas bubbles down here. And we have strong acid HCl, which is our source of H plus, in solution.

So over here again, this is our cathode. So we have a reduction at the cathode. And we have H plus aqueous, phase transition H₂, and gas, and then another phase transition to our solid platinum electrode. And over here is our standard anode, our oxidation, that involves zinc and zinc plus 2. So we could also have had the hydrogen electrode over here as well.

All right, so that gives you a number of different examples of types of electrodes that you could have and types of cells that you can have. And that is the end of this, today's-- or lecture 25 notes.

And as soon as you finish writing that down, we'll go and look at lecture 26. Again, we're continuing on the same theme. And now we're going to talk about the potential of these cells. So we've talked about the reactions happening at anodes and cathodes, how to write them.

And now we're going to think about the change in potential that occurs when you run these reactions.