

The following content is provided under a Creative Commons license. Your support will help MIT OpenCourseWare continue to offer high-quality educational resources for free. To make a donation or to view additional materials from hundreds of MIT courses, visit MIT OpenCourseWare at ocw.mit.edu.

**CATHERINE
DRENNAN:**

So problems set 5 was due now. And problem set 6 is due next week. It will be posted later today. Welcome to the parents. We will have clickers too. Thank you for coming. So today's topic, we're continuing with chemical equilibrium. And depending on when the clickers come, the fun video may be at the top of the lecture.

But before that, today is going to be a clicker competition. And hopefully, for the parents in the room, your child has told you what these clicker competitions are. The parents will be in their own recitation. If you win, I think you have to come back next week for snacks at recitation. So we'll explain that. But some recitations have been doing very well in the clicker competition, but they have not won.

So I just want to share with you another extra-credit clicker assignment. And this extra-credit clicker assignment is something you can do this weekend with your parents. So again this is optional. This is to win for your recitation extra-credit points to get into the championship at the end of the semester to win the geeky T-shirt. The geeky T-shirt is itself a prize, or a extra-credit thing to get in.

OK. So in class, I have been showing science videos. And these are the stars of the science videos. They're MIT undergrads, graduate students, post-docs, and faculty. And they've been talking about how they use basic chemical principles in their research and how that research is important for addressing a real-world problem.

So the extra credit assignment is that each of these individuals also has a two to-- I think the longest one is about-- five-minute personal video. And so the extra credit assignment is to watch one or all, as many as you want, of these personal videos, and then send Allena an e-mail with your recitation number in the subject line saying one thing that surprised you in a personal video or one thing to which you could really relate, that someone said in their personal video. So it can just be one or two sentences. It doesn't have to be long.

When I gave this assignment before, I'll give you a little help on where you might want to start.

We had the most comments about Ben's personal video and Samuel's personal video. And some people also commented on my personal video. So these are personal videos that kind of tell people's path to science or people's path to chemistry. So

I thought this might be a fun thing to do. It's available online here. You can watch some of these- again, two-to-five-minutes-- with your parents. And that might also help your parents understand a little about the people who are here at MIT. So this is something, and it's due November 3. But I thought it might be a fun thing to do homework with their parents on visiting weekend. That's an awesome thing to be able to do.

All right. So clicker competition. Recitations 7-- Dan must unseat him. Here's one that parents might know the answer to. We'll see. And again, you can share. But remember, your parents are competing against you. All right. Let's just try 10 more seconds. Very good.

So that's the answer. So that was yesterday. And so this, every year, it's National Chemistry Week on the week of Mole Day. So this is national chemistry week. So if I can get most of the way through lecture, I'm going to share with you a fun video at the end that celebrates national chemistry week. So hopefully we'll get there. But first, let's talk a little bit about chemical equilibrium, the subject for today.

So chemical equilibrium-- it's all about staying calm. So we want the rate in to be equal to the rate out. We like to be at this equilibrium place. But sometimes, often, all the time-- frankly, all the time, stress is applied to our equilibrium situation. And systems must respond in such a way to alleviate that stress.

So Le Chatelier's principle suggests that, when a stress is applied to the system, the system will respond in such a way to minimize the stress. Again, this is about keeping calm, minimizing stress. And if you think about what direction the reaction will go to minimize the stress, you can predict the direction of the reaction.

So today, we're going to talk about some of the stresses to the system. Last time, we talked about adding and removing reactants and products. This time, we're going to talk about gases. And we're going to talk about changing the volume. That's our first stress to the system.

So if you decrease the volume of a gaseous system, it will cause an increase in the total pressure. And we know this from the ideal gas law and other relationships also between

pressure and volume. So the ideal gas law is something most MIT students are familiar with, because it is, in fact, the T in MIT.

So PV equals nRT . where P is pressure, V is volume, T is temperature, n is the number of moles of that ideal gas, and R is our ideal gas constant. So this relates pressure to volume at a constant temperature and relates pressure, volume, and temperature and also number of moles to reach other. So this is a very important expression, and we won't actually be talking very much about the ideal gas law. But you need it in this unit, because in this unit, we're thinking about stresses to equilibrium situations and how reactions are going to respond.

So if you decrease the volume of a gaseous system, causing an increase in total pressure, the system's going to respond in such way to minimize the stress or decrease the total pressure. And so Le Chatelier's principle would predict that, if possible, the system's going to respond in such a way to reduce that total pressure. So you increase the pressure, reduce the pressure. Minimize the stress.

So let's look at a reaction of a gas to another gas. So here we have the molecules of gas P_2 going to one molecule of P_4 . And if we had a container with a particular volume and we compress that volume, the reaction that is at equilibrium is going to respond. And here, a decrease in volume is going to shift the reaction toward product.

So why is this? Why is it going to shift it toward product? And it's because, for every two molecules of P_2 that are consumed, only one molecule of P_4 is formed. And thus, the shift toward product, shift toward P_4 , is going to reduce the total pressure, because you have fewer molecules of gas. So each of them is a little bit happier in this now decreased Volume

So you can also just think about what direction, where are the moles, and how this is shifting. But you can also think about this in terms of Q and K , and I highly recommend this. I'm a big fan of thinking about the relationship between Q and K . So let's think about this now.

Let's just suppose that the volume is decreased by a factor of 2. So we were at equilibrium. Q was equal to K . K is the equilibrium constant, or the reaction quotient, at equilibrium. But now, we have changed the volume by 2. So we can think about what has happened to Q .

So Q is products over reactants at any time point. Product is the partial pressure of gas P_4 -- that's our product-- over the partial pressure of the reactants, the partial pressure of P_2 gas. But there are two of them, so we have to remember the stoichiometry of the reaction.

So both of these are affected. But because of the stoichiometry of the reaction, Q is now a half. So Q has changed. We have decreased Q by a factor of 2. So Q is now less than K . So tell me what happens to ΔG in this situation. Let's try just 10 more seconds.

Yep. So it is a negative value, and let's take a look at why. So ΔG is negative. And we recall this relationship that $\Delta G = RT \ln(Q/K)$. So when Q is less than K , you get a negative value for ΔG . And when you have a negative value for ΔG , the reaction goes in the forward direction. So a negative value for ΔG means it's spontaneous in the forward direction. The forward direction is toward products. And this will happen until Q equals K again.

So you can think about this based on the equation. Always ask yourself, this Q changing? Is Q less than K or greater than K ? And that'll predict the direction to which the reaction will switch. So it's going to go in the forward direction if ΔG is negative. All right.

Now let's increase the volume and think about what happens in the container. So we have a container here. We increase the volume. So now, we're going to shift to the reaction toward our reactants. So the increase in volume is going to lower that total pressure.

And we can think about this switch. We have one molecule of P_4 on the products, two molecules of P_2 here. So for every one molecule consumed in the backward direction, two molecules of P_2 are formed. So the shift will compensate for this change, and we're going to increase the total pressure, which was decreased before, by this increase in volume.

So again, you're shifting it in such a way to minimize the stress. Even if the stress is a decrease in pressure, you want to be back at your equilibrium place. So you'll shift to then increase the total pressure just to compensate in any way, minimize the stress. Whatever the stress is, you want to minimize it.

All right. Let's do something slightly trickier. Let's add an inert gas to this reaction. So we're going to add an inert gas to a container, increasing the total pressure at constant temperature. Why don't you predict what's going to happen? All right, 10 more seconds.

Yes, this is a tricky one. Since it's a clicker competition, we won't repoll, and we'll put up the answer. Ooh, by a slight edge. So this seems very counter-intuitive, so let's discuss it. So nothing happens. But why? So Q depends on the partial pressure of these gases.

And under the scenario described here, the partial pressures were actually not changing. So Q was not changing. And therefore, there was no shift. So to understand this, we have to do a little partial pressure review. So partial pressure of gases. The partial pressure is the pressure that each gas would exert if it were by itself. That's the definition of partial pressure.

So if we have oxygen at one atmosphere and nitrogen at one atmosphere, and you add them together, and now you have a total pressure of two atmospheres-- 1 plus 1 equals 2-- but the partial pressure of O_2 here is identical to its partial pressure here, its partial pressure as if it was there by itself. It has no interest in this inert gas. It ignores it and feels no different as a result of it. Very just sort of isolationist with these ideal gases. So the partial pressure of O_2 here is the same.

And so we can look mathematically at this expression. So the partial pressure of gas A is equal to the number of moles of that gas, RT over volume. The total pressure is equal to the partial pressure of gas ABC, how many ever gases there are. But the partial pressure of each gas, like O_2 here, just depends on the number of moles of that gas and the volume of that container, assuming constant temperature.

So if the partial pressure isn't going to change-- so if you add inert gases, increasing the total pressure, the partial pressure is unchanged. And when the partial pressure is unchanged, Q is unchanged. Remember, Q equals the partial pressure of the gas products over the partial pressure or the gas reactants. So if partial pressures don't change, Q does not change. And if Q doesn't change, then Q is the same as K , and there is no shift. So we have to always ask ourselves, what's happened to the container? What's the partial pressure? Did the partial pressure change?

Let's try another example. Let's add another inert gas. But this time, we're going to add it to a container. And the total pressure and the temperature are kept constant in this case. So this is another clicker question. Why didn't you see what you think is going to happen here? All right, let's do 10 more seconds.

All right. This is a good clicker day. Sometimes whenever the answers are all 90%, then it's decided that these questions will determine it. All right, so let's look at this. So we're going to shift toward reactant. Again, let's consider why. All right. So in this problem, we're adding inert gas. Total pressure and temperature are constant. Well, temperature was constant before. But total pressure wasn't in the last example.

So let's go back to our picture here. So if the total pressure was to stay constant-- and normally, if you add, one atmosphere of O₂ and one atmosphere of N₂, the total pressure is two atmospheres, because total pressure equals partial pressure of one gas plus the partial pressure of the other gas. So if the total pressure is constant, what has to happen? Tell me. What happened?

So the volume must have changed. So if the temperature stayed the same, the volume of the container must have increased. Because otherwise, the pressure, the total pressure, should have changed. And it didn't. So the system was manipulated. The volume must have increased.

So now, let's think about what happens when the volume increases. When the volume increases, the partial pressures decrease, because they're now in a bigger volume, because partial pressure depends on number of moles. It depends on volume at constant temperature.

And so if the partial pressures decrease, then Q is also going to change. And when Q changes, the reaction is going to respond. And in this way, it's going to respond in such a way that you want to increase the pressure, because we have this decrease. So now we want to increase. And so we're going to switch from one mole to two. And so the reaction is going to shift toward the reactants.

So in doing these problems with partial pressures and gases, you always have to ask yourself, did Q change? And whether Q changed is whether the partial pressures change. And partial pressure will definitely change if you change moles or if you change the volume.

So often in these problems, you're saying, did the volume change? Volume change, partial pressure change. Q is different. The reaction is going to shift. So there's lots of examples on this in problem set 6. So you can go work on those. All right. So that is about volume. We've had temperature constant. Can't do that forever. Now it's time to change the temperature.

All right. So changing the temperature of an equilibrium mixture by adding heat is going to shift the reaction in such a way that some of that heat is absorbed. So again, this is consistent with Le Chatelier. Minimize the stress. Add heat, absorb heat. Remove heat, create more heat. Again, you're minimizing the stress to the system.

So now let's think of different types of reactions. We have endothermic and exothermic reactions. And let's think about the effect of changing the temperature on those reactions. So

let's think about raising the temperature of an exothermic reaction. Is that going to shift the formation of reactants or products? What do you think? Reactants. All right, let's consider why.

So for an exothermic reaction, that means that the reaction in the forward direction is exothermic, which also means that the reaction in the reverse direction is endothermic. So for an exothermic reaction-- exothermic in the forward direction-- heat is produced when you go from reactants to products. And in the endothermic direction, heat is absorbed. So if you raise the temperature, you're going to want to minimize the stress. So you want to shift it in the endothermic direction. You want to shift it in the direction that heat is absorbed, which, in this case, is toward reactants.

Now let's think about an endothermic reaction. Raising the temperature of an endothermic reaction is going to shift toward products. And if we put our pictures again, an endothermic reaction is endothermic in the forward direction. So heat is absorbed from reactants to products. And in the reverse direction of an endothermic reaction is the exothermic direction in which heat is produced. So if you raise the temperature, then you're going to want to shift in such a direction to minimize that stress in a direction to absorb heat, which, in this case, sends you toward products.

So here, the predictive tool is our friend delta H. So delta H tells us whether reactions are exothermic or endothermic, whether delta h is negative or positive. So for some things, the prediction is what's happening to delta G. Here, our predictive tool is the sign of delta H. So let's do a clicker question, and tell me what direction you think this particular reaction is going to go if you add heat. All right, 10 more seconds.

All right. So is this an exo- or endothermic reaction? Exothermic, because we have the negative sign. And so then we can think about what's going to happen. So you add heat. You're going to want to shift it in the direction to absorb the heat. So you're going to shift it toward reactants in the endothermic direction, absorb the heat. So again, delta H is the predictive tool. You have to remember negative delta H exothermic, positive delta H endothermic. And then just think about shifting to minimize the stress, always minimizing the stress. All right.

So our equilibrium constant K. It's a constant, except that it changes with temperature. So the equilibrium constant has a temperature dependence. So big K, the equilibrium constant, can change with temperature. Reaction rates can also change with temperature.

So when we talk about kinetics, we're going to have little K_s , which are reaction rates. And they're going to change with temperature as well. So how does the equilibrium constant change with temperature? And you might feel a derivation coming on, and you would be right.

So let's think about what we know, equations that involve the equilibrium constant. We saw before that ΔG^0 equals minus RT natural log of K . What else do we know about ΔG^0 equations? What else does ΔG^0 equal? It equals our friend ΔH^0 minus $T \Delta S^0$.

And we can rearrange this expression, solving for the natural log of K . And if we assume-- and it's a good assumption-- that ΔH^0 and ΔS^0 are independent of temperature, that means that K is going to change with temperature. And we can think about two different temperatures and come up with a unified equation.

So if we consider a reaction carried out-- temperature T_1 and temperature T_2 -- we would have natural log of K_2 equals minus ΔH^0 over R and over temperature 2 plus ΔS^0 over R . And we have the same thing now but for equilibrium constant 1 and temperature 1. So we can combine these, and things will cancel.

So subtracting those two equations gives you this. And it's a named equation in chemistry, which means that it's probably pretty important, because a lot of things don't have names. So we have the natural log of the equilibrium constant K_2 over equilibrium constant K_1 equals minus ΔH^0 over R , which is a constant, in brackets, 1 over temperature 2 minus 1 over temperature 1. Van 't Hoff equation.

And later in the semester, I'm going to ask you to recall this equation. And I'll have a special prize for the first person who comes up with the Van 't Hoff equation, so you can keep that in mind. I think it's a lecture 34. So we have a ways to go. All right.

So let's think about what is true, then, if ΔH^0 is less than 0. First of all, is that exothermic or endothermic? Exothermic. All right. So now we're going to think about increasing the temperature of an exothermic reaction. So temperature 2 is going to be greater than temperature 1. And we already thought about what was going to happen. But let's make sure this equation is valid and works for what we just saw.

So if we increase the temperature here, what's true about this equation? Well, what's true is you have a minus sign here. ΔH^0 is negative for exothermic. And if your temperature 2 is greater than temperature 1, that's also a negative. So overall, you're going to have a negative

value. And if you have a negative value here, that's going to mean that K_1 is greater than K_2 , which means that, at your new equilibrium, there'll be fewer products.

All right. Well, this works with what we were just talking about. If you increase the temperature of an exothermic reaction, it shifts it toward reactants to absorb the heat. So there would be less products. We've shifted toward reactants. So you can just think about what's happening and moving in the direction to minimize the stress. And you can also look at the equation and mathematically figure out what's going to be true about the equilibrium constants. It's great to have a backup plan.

All right. So if we decrease the temperature, T_2 is less than T_1 . So now we have our minus. We have a minus. But now this is a plus term. And so that's going to mean our K_1 is less than K_2 . We have more products here. So if we decrease the temperature of an exothermic reaction, the reaction responds in such a way to minimize the stress, add more heat. So you switch to the exothermic direction. And so you should have more products at this new equilibrium.

All right. So now, why don't you tell me what should happen for the scenario where ΔH_0 is greater than 0? Which of the following should be true? All right, 10 more seconds.

Yeah, so most people are right in that-- yeah. I think this is a good example of how, when things are expressed different ways, sometimes it's hard to see that it all is consistent with each other. So yeah, all of those turn out to be true. All right, so let's look at it over here.

So this is an endothermic reaction. When we increase the temperature of an endothermic reaction, you still have this minus sign. But now this is plus. But when we've increased the temperature, this is minus. So you get an overall plus in signs, which means that K_1 is less than K_2 . Increasing the temperature of an endothermic reaction shifts it toward products. If we decrease the temperature, the net sign is negative here. So K_1 is going to be greater than K_2 . We decrease the temperature. It shifts in the exothermic backwards direction, and we have less products.

All right. So K changes with temperature. You can see this equation on your equation sheet or just think about minimizing the stress. So what is this useful for? Well, figuring out how to maximize the yield of a reaction can be very important in industry.

One process that is an important industrial process is the Haber-Bosch process, where you

take nitrogen gas and hydrogen gas and make ammonia. This process, which was developed a long time ago, is still used today. And each year, you make 1.6 times 10 to the 10th kilograms of ammonia, which you use for fertilizer and other things.

You know what other things ammonia is often used for besides fertilizer? Explosives, yeah. So if you look at the dates of these process development and the names of these people, they're German. And so we have World War I and World War II. I'm thinking that the push to come up with this process was probably not for fertilizer and the farmers at this time. This was an important part of the war effort, to get this process to work to make explosives.

So Haber over here also is considered the father of chemical weapons. He came up with gases, such as phosgene. And he would study the effects of how they killed people. And one of the ironies is that he was Jewish. And so a lot of his discoveries were used and resulted in his relatives dying. So he was an interesting figure in history. Carl Bosch, on the other hand, disapproved strongly of Hitler and was fired from his post. He didn't lose his life, but his career was ended. But he still had his Nobel Prize, so maybe it was OK.

So this is an interesting little history about chemistry. But this process is still used now. No one has come up with anything better. And we could think about, what are the challenges of making this reaction go? So it's an exothermic reaction. So what temperature would you want to use to get more products?

Right, you want to use low temperature, because that would favor your products. But low temperature is also sometimes not great, because it slows the rate, which is bad. So the compromise temperature that's used in this process is 500 degrees Celsius. So it's pretty hot. So this works, again, maximizing the yield of products. So what are other ways to drive this toward products? So can someone suggest another way [INAUDIBLE]? I don't think I have two [INAUDIBLE]. If you have your parents here, they might want this lovely residential water bottle.

[LAUGHTER]

What are other ways to drive this toward products? Maybe someone could run the mic up there.

Is this-- yeah, it's gone. You could decrease the volume?

Yep.

Yes!

Yep.

[LAUGHTER AND APPLAUSE]

[CHEERING]

So there's a couple things that you can do. One is you compress the volume. And the reason why this works is that, if you note you have four molecules over here-- 1 plus 3 going to 2. So if you compress that volume down, you're going to shift it toward the fewer number of molecules, which is 2. So that shifts it toward products.

The other thing that they do is they actually remove products from the reaction. So they liquefy out the product at some point. And when you remove the product, that also shifts the reaction toward more products to minimize the stress. So these are two of the things that are done.

So you have to have high pressure to compress that volume down and high temperatures. So this process costs a lot of energy. So we spend a lot of energy in this country doing this reaction. And the reason why is because it's really hard to split nitrogen. So what kind of bond does nitrogen have, N₂ have? Triple bond, right. It's very hard. It has a high bond order. It's very hard to break that bond. And so you really have to work at breaking it.

So if you're doing industry, there's lots of nitrogen in the air. So a lot of it is available. But it's just very hard to utilize it to fix that nitrogen into another form. So lots of nitrogen in the air. Nitrogen is needed by all living things, but it's really hard to break that triple bond.

But if you're a bacteria, you have a much easier time doing this than if you are a person. And bacteria have an enzyme called nitrogenase. Adding the "-ase" means it's an enzyme. So this is an enzyme that works on nitrogen. Most people call it "nit-raw-genase," but it's really "nitrogen-ase." And it splits nitrogen.

And it does this by these really complex metallocofactors. Lots of metals in here. You have iron and molybdenum. You also have inorganic sulfur. And it's those d-orbitals that let you do this. So you can use high temperatures, high pressure. Or you can get a transition metal with beautiful d-orbitals to do your work for you. And a lot of people are trying to come up with their own catalyst inspired by nature to get this to work. So this is an ongoing area of research.

And I just want to mention one of the leaders in the field. This is a husband-and-wife team, actually. And Markus Ribbe is a German with a very heavy German accent. And so Germans have been dominating this nitrogen fixing field for a very long time. But Markus Ribbe is like the nicest person you would ever meet in your entire life. And he's trying to figure out environmentally friendly ways to do it. So again, chemistry is a powerful tool to change the world. But let's change it for the better, please, and try to come up with ways to protect our environment.

All right, so there's one application of Le Chatelier's principle, maximizing the yield of a reaction. Another example comes in medicine and Le Chatelier's principle with respect to another metalloprotein, which is hemoglobin. So hemoglobin has a "heme," which is iron, more d-orbitals. If I'm this excited now, wait till I get to d-orbitals. It's very exciting. I can't wait. I love transition metals.

So anyway, so we have hemoglobin. So hemoglobin Hb plus oxygen yields oxyhemoglobin. It's hemoglobin with oxygen bound. And this is an incredibly important reaction, because otherwise your air, your O₂, would be in your lungs and wouldn't leave. But hemoglobin carries it to the body, where it can be used to make energy. So we really need the reaction. We want to maximize the product of that reaction.

So this disequilibrium can be shifted if you have less O₂ around, a lower partial pressure of O₂. And this can happen as you go up in altitude. The pressure of oxygen decreases, and that's a real problem. So any of you real serious mountain climbers? At least one. How about somewhat serious, climbed a mountain, maybe worried about your hemoglobin levels, altitude thickness? A few, OK.

So what can happen here is, when you have, again, lower pressure, then the reaction shifts in the opposite direction. It's like you have a lower amount of reactant, so you shift to make more reactant. And this is what happens. You have less oxygenated hemoglobin, and that can make you pretty sick.

So for some of you if you want to say, what can you do about this? How can you shift the reaction back the other way? Do you know what your body does? Your body is very clever. What does it do? Yeah, so it makes it more hemoglobin. So that's why you need to give your body a little time to adjust. Can't do it immediately. You have to give it a little time to make more protein. But when you increase the amount of reactant, it'll shift toward product. So when

you increase the reactants to minimize the stress, it shifts to make more products.

So this is how the body compensates. So I'm going to tell you very briefly about significant figures, and then I'm going to have time to show you a Mole Day video. And it's actually from UC Irvine, where Markus Ribbe is from. So today, we're celebrating UC Irvine chemistry department.

Anyway, I just want to mention significant figures. I have a confession to make. I didn't know the rules of significant figures for logs when I started teaching this course. And this will be information that I think now I'll have to purge from my brain on my deathbed.

But pay attention to these, and we'll have some clicker questions on them. So that'll be fun. OK. So let's look at these two-- log of two different numbers. Note the 3 here is the 3 over here. The 23, 10 to the 23-- 10 to the 23, Mole Day-- is the 23 here.

But those are not significant. That 3 and that 23 are not significant figures. These are the significant figures-- 7.310. So there are four significant figures here. Those four significant figures end up after the decimal point. So those are the significant figures after the decimal point. And if you go the other direction, if you have a 0.389 here, that has three significant figures. So three significant figures here. But if you have 12.389, you still have three significant figures, because that 12 ends up over here.

All right. With that, give it a try. This could be for the clicker championship of today. All right. Let's just do 10 more seconds. Should I show my video first, or do you want to do the winner first?

Do winner first.

Winner first, OK. All right. OK, so 3 after the decimal place. And let's look at the winner, and then we're going to end with a short video. Oh, man! All right. Dan, whoa. OK. All right, so we're going to end with a short video. And here it is.

[VIDEO PLAYBACK]

[MUSIC - "LET IT GO"]

- (SINGING) Snow looks white, 'cause it's scattering light elastically. Atoms are everywhere I look, though none of them can be seen. Molecules have always been a challenge to resolve.

Light microscopes don't work. Heavens knows we tried. Do they exist? Can they been seen? Maybe using some spectroscopy! For centuries, we didn't know. Well, now we know! Chemists know, chemists know what makes water freeze and flow. Chemists know, chemists know why fireflies can glow. All chemists know Avogadro order 10 to 23. The moles never bothered me anyway.

Angstroms measure distance, nomenclature hard to say. Though we are macroscopic, we always find a way! We love to see what we can do, synthesizing novel molecules from Exclusion Principle of Pauli-- chemistry! Chemists know, chemists know about ozone in the sky. Chemists know, chemists know sigma bonding versus pi. We understand Le Chatelier. Equilibrium!

There's chemistry to be found from the air down to the ground. We make these compounds form in flasks with bottoms that are round. We hope they crystallize, 'cause columns are the worst. Still we strive to do all this for the chance to be the first! Chemists know, chemists know why it's red at the break of dawn. Chemists know, chemists know where reactants go when they're gone. We need to know every single way to form or break bonds! The moles never bothered me anyway.

[END PLAYBACK]

[APPLAUSE]

All right, have a great weekend. Please return any borrowed clickers.