CATHERINE DRENNAN: All right. Let's just do 10 more seconds. All right. Does someone want to tell me up here-- hello everyone up here-- how they got the right answer? Over-- everyone over there, I guess. OK, I'm coming over there. So for the answer we have an MIT chemistry bag. That's quite special.

AUDIENCE: Well, since your delta H value acts as the y-intercept, you know that it's negative because it's negatively. And then you also see that you're ascending positively, so that you know your slope is positive. And so delta S also has to be negative to make it positive.

CATHERINE DRENNAN: Interesting. That was a different answer than I would have given, but that works really well, which is always great to ask people. So you could also think about this in terms of your going from a situation with a negative delta G to a positive value. So you know that temperature is going to make a difference there. And so that would also tell you how the temperature changes is also going to tell you the answer to that. So that's great.

All right. So a couple of announcements. As you can see, see what I am wearing on my legs. So I was pleased with the performance on exam 2. The average was 84.7, not quite as celebratory as 87.4, which was the average on exam 1. Yeah, those are real numbers. That's a little weird. But anyway, I was still very happy with the performance on the exam.

I would like for exam 3 for the average to be back at 87, though. I like it to be very close to 90 because I want everyone to have an average around 90 in the class to demonstrate excellent knowledge of chemistry. So I try to write an exam where I feel like if you have excellent knowledge of chemistry, you can get a 90% or above. So this is my goal.

I all want you to have this fundamental knowledge of chemistry that you can go out and solve the energy problem, that you can go out and be ready to tackle the next great challenge in health. There's so many challenges facing us, and we don't even know what they are right now. So the Ebola situation, I think, is a real case in point, that we don't know what the next challenge facing us is going to be.

So we need to be ready with our chemistry knowledge. The only thing I can tell you about the

challenges facing us is that knowing chemistry is going to be really important in tackling those problems. So everyone needs to have an excellent understanding of chemistry to go on and do well and solve the problems of the world. So 84.7, really good. You're getting there. Awesome.

So celebratory tights, or leggings. Not really sure the difference between tights and leggings. Anyway, I was very pleased. But next unit, next exam, has thermodynamics, chemical equilibrium, and acid-base. We're already done with thermodynamics and moving on. So it's just for the rest of the semester, we're about halfway there, the rate seems to accelerate.

So problem set 5, thermodynamics. It's a little shorter because you have fewer days to do it. So we wanted to make sure that you would be able to get it done by Friday at noon. But that's when it's due. It's all thermodynamics. We've already covered all the material on the problem set. You can do it already right away.

All right. So also in my wardrobe-- I'm commenting a lot on my wardrobe today-- you'll notice maybe that I've worn this shirt once before, and my goal was really to have a different shirt every time. Now, this shirt is appropriate for today's lecture.

But I just wanted to say that I did try to get a different shirt. And I ordered new shirts on thermodynamics and chemical equilibrium this year, and one of the shirts had to do with entropy. And I just thought I would share with you briefly, because it's a good review of thermodynamics, what happens when you do things related to entropy.

So just to sort of show you what happened. In September, I was super organized. I ordered my shirts. And on October 3, they shipped. They went through Indianapolis to Massachusetts and to Newton Highlands about somewhere around October 8 or 9. That's where I live, so that's really good.

But the shirts were about entropy, and then look what happened. Wait-- rerouted for delivery to a new address. Jersey City, Cincinnati, Ohio, Warrendale, Pennsylvania. And then yesterday they were in Springfield, Massachusetts, to which I say, entropy! This does not happen when you order T-shirts about kinetics. That's all I'm saying. Those arrive record time. Thermodynamics is not always your friend, but you can embrace it.

And I'm going to try to calm myself down because we should be at equilibrium today. Today, the topic is chemical equilibrium. It's the calm part of thermodynamics. We'll get back to

entropy. There will be some mention of entropy.

But I'm going to try to calm down. I'm going to remember my reaction quotients and my equilibrium constants and think about chemical equilibrium and embrace the T-shirt I am wearing. Because I feel powerful in this T-shirt, like I can control thermodynamics with the power of the Green Lantern, with chemistry knowledge at my fingertips, and I will move forward.

So chemical equilibrium is calming, but it's also a dynamic process. We can't forget that the reactions have not stopped. It's just the rate of the forward reaction is equal to the rate of the reverse reaction. So there's no net change, but the dynamic process is still happening.

So my goal in life is to have equilibrium in my life. The amount of work-- the rate at which work comes into my office equals the rate at which I complete the work and it leaves my office. Most of you should also have this goal. The rate at which problem sets come in equals the rate at which you complete the problem sets and the problem sets get turned in. You don't want to get yourself in a situation where that equilibrium is messed up, where the rate of things coming in is just nowhere at all equal to the rate at which things are completed and going out.

So goal-- chemical equilibrium. We want to be at this calm place where we're still working hard. It's still dynamic. But the net change is good. It's in our favor. The amount of work we're doing equals the work coming in. So I like chemical equilibrium. I feel like it brings hope to thermodynamics.

So let's look at an example of a reaction, and we're going to talk a lot about this reaction in chemical equilibrium. So we have N2, nitrogen, H2, hydrogen coming together to form ammonia, and it has a delta G0 of minus 32.9 kilojoules per mole. So let's think about what's happening in the beginning of this reaction, when we're just starting this reaction.

So we can plot concentration on one axis, and we can have time on our other axis, so time increasing over here. And we can think about what happens when we're starting our reaction. We can have hydrogen starting out. We have some hydrogen in a certain concentration, and it's going to come down and level out. So that's our H2.

We'll also have nitrogen as a reactant. It comes down, levels out. And we're starting with our product. We have no product. We're just starting the reaction with the reactants. So product increases and then levels out. So we're decreasing our reactants, increasing our product, but

the lines go flat.

The reaction is still happening. The reaction in the forward direction is happening, and then in the reverse direction is happening. But there's no net change at equilibrium because it's reached a state where the rate of the forward reaction equals the rate of the reverse reaction.

And let me also just point out some arrows and some nomenclature here. So we have this double arrow, and you can see this a couple of different ways. But this indicates the reaction's going forward, and the reaction is going back. And that's necessary if you're at equilibrium. The definition of equilibrium-- rate forward equals rate back at equilibrium. So you'll be seeing those arrows a lot.

All right. So let's think about delta G now in terms of what's happening at different times in this reaction. So if we're just starting out this reaction, and you don't have a lot of products, so you're in this case of pure reactants. At this point, you're going to be moving in the forward direction. You need to make products.

And so at this point, the forward direction is negative. It's spontaneous. You're spontaneously making your products. You don't have any at this point. And so your delta G is going to be negative. So up here with pure reactants, your delta G is negative.

Now, if you started with pure products-- you could do that as well-- the NH3 will dissociate and form the reactants. And so that would be over here. And so when you have excess products, now the forward direction is positive. It's not spontaneous because it's the reverse direction that is spontaneous.

So over in this part of the graph over here, our delta G is positive. So if you have pure reactants, you spontaneously go to products. If you have pure products, you're going toward reactants. And when you have the right amount of reactants to products, the right ratio for equilibrium, then you get to equilibrium and delta G equals 0 at that point. So delta G is going to change depending on your ratio of products to reactants.

So delta G changes as the proportion of the reactants and the products change, and there's an equation to describe that. So we're back now to having a few more equations in this unit. We noted a lot of people forgot to bring calculators to exam 2. There weren't so many calculations. We're back to more calculations again.

So delta G equals delta G0 plus RT natural log of Q. So let's think about these terms. So delta

G is Gibbs free energy at some particular point, at some ratio of product to reactant. Delta G0 is Gibb's free energy under standard conditions, in their standard state. Q is our reaction quotient, products over reactants. And R is the universal gas constant, and T is temperature. So this equation describes how delta G will compare to delta G0 at a particular condition, a particular ratio, of products and reactants.

So let's talk about this reaction quotient Q. So let's look at a made-up reaction-- A plus B goes to C plus D. And think about that reaction in the gas phase. When we think about that reaction in the gas phase, we're going to be talking about partial pressures. And so Q over here is products over reactants, but here we're writing it out in a more complicated long way.

So let's go through this. So we have this P sub C. Well, what is that? That is partial pressure. So partial pressure of gas X, and this is partial pressure of gas C, and C is a product and D is a product. So we have products over reactants. Partial pressure of gas C over P reference, partial pressure reference, which is 1 bar. The reference is 1 bar, so we're dividing that by 1. And this is raised to the little c, which has to do with the coefficients of the reaction.

Then we have partial pressure of gas D over our reference raised to the small d-- those are our products, products are C and D-- over reactants-- partial pressure of gas A over reference raised to little a, partial pressure of gas B over reference raised to little b. But 1, the number 1, the reference is 1. So most of the time you'll see the following expression-- Q, our reaction quotient, partial pressure of gas C to the stoichiometry little c partial pressure of gas D raised to the stoichiometry d, over reactants, partial pressure of A little a partial pressure of B little b.

And you will note that because we had 1 bar and these are in bars, our units are going to cancel here. So Q does not have units. I knew this should be very exciting for you. You won't be losing one point for lack of units with Q. There are no units. So this is very exciting. So there's the expression that you'll mostly see for Q. And now we can kind of forget that the reference is there. We can ignore the reference. You just don't freak out later if you have no units. It's OK.

All right. So we also could be talking about solution, and here we're going to talk about concentrations. In a lot of these problems, they're talking about things in the gas phase, and they're giving you concentrations. Don't worry about it. It's all OK. So here our concentration reference, C for concentration, is 1 molar. And you'll see this term-- C in brackets. That means concentration of.

And so if you see that like this, you would express this in words as the concentration of C-- again, that's a product-- over the reference-- we'll get rid of the reference in a minute, but we'll keep it for now-- raised to the little c, concentration of product D over this concentration reference of 1 molar to the little d, concentration of reactant A over this concentration reference raised to the little a, concentration of other reactant B over our concentration reference of 1 molar raised to the little b.

And then, again, this is concentration. And we can get this expression, which is usually the one you see, of just products over reactants. But I will make one point here, that you need to know how to balance equations to be able to do these. So in the beginning of the book there was something about balancing, limiting reagents, stuff like that.

If you feel like you did not master that, you want to go over it. You need to be able to balance. And in a lot of these problems you may be thinking about limiting reagents again. All right. So that's Q. You can see it as partial pressure-- and I'm going to do a little partial pressure gas review on Friday-- or concentrations with these brackets.

So at equilibrium, delta G equals 0. And it's a dynamic process-- rate of the forward reaction equals rate of the reverse reaction. Q now is the equilibrium constant, because the ratio of products to reactants at equilibrium is the definition of the equilibrium constant. So when delta G equals 0, Q equals K.

So we now can go back to this expression that we had and rewrite this for the situation at equilibrium, when delta G equals 0. So when delta G equals 0, this is 0 and Q is K. K is products over reactants at equilibrium. And we can rewrite this or rearrange it now-- delta G0 equals minus RT natural log of K. And that's a very important equation that you'll be using a lot in these units. So it relates delta G0 with our equilibrium constant K, and it depends on temperature.

All right. So K-- same expression as Q, except for something very important, which is that the concentrations are the concentrations at equilibrium. So it has the same form as Q, but it's only the amount of products and reactants at equilibrium.

So in the gas phase, we would write K in terms of our partial pressures. And again, this is this little symbol to remind you these are the concentrations at equilibrium. And in solution it would be expressed in molar or something else. So K is, again, products over reactants, but only

those concentrations at equilibrium. Whereas Q is at any point, any concentrations of products over reactants for Q. For K, it's those concentrations at equilibrium.

All right. So now let's think about Q and K together. So we can rewrite this expression again. We just derived a new expression for delta G0, and that was minus RT natural log of K. So we can say delta G equals, and now have this other expression, minus RT natural log of K plus RT again-- gas constant and temperature-- natural log of Q.

And we can simplify it, bring out the RTs, and now we have delta G equals RT natural log of Q/K. And this is, again, a very important expression that you'll use a lot. Because it tells you about delta G, whether the reaction is going to be spontaneous in the forward direction or the reverse direction, depending on the ratio of Q and K.

This is a very important equation for chemical equilibrium. So let's think about what this means, what comes out of this. So the relationship between Q and K, if Q is less than K, what is the sign of delta G? You can just yell it out.

AUDIENCE: Negative.

CATHERINE DRENNAN: Yep. So it's going to be negative. Just mathematically, you can look at that expression. So it'll be negative. And again, we give you all the equations on the equation sheet. That will be there. So you just need to know how to think about it.

And so the forward reaction will occur. And so if we think about this, it's going to mean that at equilibrium there are more products than there are at this time for Q. There's less products in the Q expression. K is greater than Q. So we need to make more products. So delta G will be negative, and the forward reaction will occur.

So when Q is greater than K, delta G is positive. And so when Q is greater than K, that means there's more products now in Q than there are at equilibrium, too many products. So we need to shift it in the reverse direction. So delta G will be positive. So again, thinking about the ratio of Q and K tells you about what direction is going to be spontaneous. Is it spontaneous in the forward direction or in the reverse?

OK. So let's continue doing a little example. We can do a little calculation here on the board for the same reaction. We're given a value of K. And now we're told some partial pressures of these gases and asked which direction the reaction will go.

So let's write the expression for Q. So Q, again, is going to be equal to products over reactants, and our product here is our NH3. So we're going to be talking about the partial pressure because it's a gas of NH3. And am I done with the top part of this expression?

AUDIENCE: No.

CATHERINE DRENNAN: No. What do I need?

AUDIENCE: [INAUDIBLE].

CATHERINE DRENNAN: Yep. Again, you need to remember the stoichiometry of the reaction. So now we have the partial pressure of N2 on the bottom, and I'm good, and the partial pressure of H2. And again, I have to remember that there are three H2s in this balanced reaction. So we have the partial pressure of H2 to the third. I can put in my numbers. 1.1 squared over 5.5 over 2.2-- these numbers may be made up, that's OK-- equals 2.1 times 10 to the minus 2.

And we're back to thinking about significant figures a lot again in this unit. So I have two there, so I'm going to have these two here, and I'm good. But now this is the value of Q. And I want you to tell me with a clicker question, what direction is this reaction going to go? All right. Let's just take 10 more seconds.

All right? So recognized, for the most part, that yeah, most of these numbers, that Q is a bigger number than K. And so then you have to think about what that means. And so when Q is greater than K, then you're going to go toward reactants. And so you're going to dissociate the product until you achieve equilibrium again. And so this means when you have a big Q number, you have too many products compared to the equilibrium state, and you want to dissociate your products. So you're going to go spontaneously in the reverse direction.

Now, I just want to mention one point. For the first part of the course, I only taught that material once before. But the second part I've taught a lot. So I have lots of years of experience of what people do wrong on exams on this part, and I will share one thing. A lot of students-- and faculty too, I'm this way-- right-left challenged. So they write left when they mean right. They write right when they mean left.

And so for those of you who are like me and have this issue, draw an arrow. When people draw an arrow, they always draw it in the direction that they mean, or they say toward reactants or toward products. And so I've seen so many times on the test they'd explain the

answer beautifully, and then write the wrong direction down.

So if you are right-left challenged, try toward reactants or toward products or draw an arrow. So that is my suggestion to you, as someone who's also very bad with saying the right direction, left or right, when I mean that direction. OK.

So what does K tell us? So K tells us about the ratio of products to reactants at equilibrium. And if you have a very big number for K, it's going to tell you something about the ratio of those products to the reactants, and there'll be a lot of products for reactants. So let's just think about this for a minute.

So when K is greater than 1-- K is products over reactants at equilibrium-- that's going to mean high products. So you'll have a higher number there. But if you have a small value for K, then you're going to have low products at equilibrium. So this is a good thing just to think about and check yourself when you're doing these problems. Does my answer actually make sense?

And if you see a big value for K, you're like, great. If I want products, I want a big value for the equilibrium constant, means that this reaction is going to give me a lot of what I want, a lot of my products. Because again, K is products over reactants at equilibrium.

All right. So let's do an example here. So here is an example where K is greater than 1. It's actually 6.84 at room temperature. We have a delta G0 of minus 4.76 kilojoules per mole. Two NO2 molecules going to N2O4. So we're going to start with 1 bar of our reactants and no products.

So we can look at what that's going to look like in our plot of concentration versus time. And here the concentration is indicated as partial pressure, which is one type of concentration. So we're starting only with our reactants. It will go down. We'll have no products in the beginning. At time 0, no products, and that will grow up. The curves will level off as you reach the equilibrium. Again, the reaction's still going in the forward and reverse directions, but the rates are equal so there's no net change in the concentrations then.

So now we can actually do some math and figure out how this changes. And so this is one way that you can set up these kinds of problems. And we will be doing these types of problems in chemical equilibrium, acid-base equilibrium. There'll be many examples, so it's good to become familiar with them if you haven't seen it before.

So we can calculate the partial pressures at equilibrium using this information. So when we started, we only had our reactant. We had 1 bar, 1.000 bar, to be specific, and we had no product. As the reaction goes, we will form product, so it's plus X. Some amount of product is going to be formed. What is the change in partial pressure as you go to equilibrium? What happens to this? What changes? What do I put in this area?

AUDIENCE: [INAUDIBLE].

CATHERINE DRENNAN: What?

AUDIENCE: [INAUDIBLE]

CATHERINE DRENNAN: I heard minus--

AUDIENCE: [INAUDIBLE].

CATHERINE DRENNAN: So I heard minus X. I heard somebody say [INAUDIBLE]. It's minus 2X because you have to remember the stoichiometry of the reaction. And so we have two of these going to one of those, so it's minus 2X. And then at equilibrium, we have 1.000 minus 2X is our concentration of our reactant, and our concentration of product is plus X.

So now we're going to solve for this, and we need to write an expression for the equilibrium constant. So why don't you write that for me. All right. 10 more seconds. I'd like 90%. Oh, so close. OK. So again, the trick here, you have products over reactants, and you have to remember the stoichiometry.

So we can go back over here. So we have partial pressure of our product over partial pressure of our reactant. And we have the stoichiometry there, and now we can continue to plug in. So the concentration, the partial pressure, of our product is X at equilibrium. The partial pressure of our reactant is 1 minus 2X, and that whole term is raised to the 2. So it's all that whole term is squared. And if you solve for X, you will get 0.38 bar. So you definitely want to remember calculators and things on this exam.

And that value, then, is our product. So that's the answer to the partial pressure of the product because that's what X is. Another thing that I've seen on exams is that people solve for X. They're happy to solve for X. But then they don't remember what X was. So keep track of what

things belong to what.

So then we want to also find the partial pressure of our reactant. That was 1 minus 2X. So here we have 1 minus 2 times X. And you can see the significant figure fun that one can have in this because we have multiplication, division, subtraction, and pretty soon we're going to have log significant figure rules. So there's going to be a lot of fun. And of course, then this is our reactant over here.

So let's just go back to the diagram, which is up above in your notes, on the same page in your notes, and plug these in and think about what this means. So we've done the math. And we see that our reactant at equilibrium is 0.238 bar, and our product at equilibrium is 0.381 bar. So K is greater than 1, more products, and you see that that works out.

So if you were asked just to explain what you expected, you could say K is greater than 1. I expect more products. But if you do the math, and you'll often do the math, you can calculate what the partial pressures are at equilibrium of the reactants and of the products.

All right. So now let's think about the relationship, again, between delta G0 and K. So here is our expression we saw before. We can also rewrite this to solve for K. So sometimes you will be given information about delta G0 and asked to calculate a K at a particular temperature. But we can also think about what we would expect. So if K is large, what is going to be true about delta G0? And so why don't you tell me what you think. All right. Let's just do 10 more seconds.

So it would be a large negative number. And we can think about this, that if you really are lying on the side of products in your reaction, then that would be consistent with a bigger value and a negative value. So you can think about-- again, we talked about in terms of formation, is the thing that's being formed more or less stable compared to its elements, the same sort of idea. You can think about the relative stability and whether you'd expect more products or more reactants at equilibrium based on these values.

So let's do an example and prove that this is true. So let's consider baking soda again. Baking soda works really well for this unit on thermodynamics and chemical equilibrium. So again, we have our baking soda going to CO2, which is very important, the CO2 gas, which helps our bread rise.

And we calculated last week that the delta G0 at room temperature for this process was plus

36 kilojoules per mole. And that would mean that it's not spontaneous in the forward direction, which is really bad for our bread, because we need the CO2 gas to cause it to rise. But the good news was that if we remember to turn on the oven in baking bread and put it at a normal temperature in the oven, that the delta G0 is minus 15 kilojoules per mole. So then it becomes a spontaneous reaction.

Now we can think about this in terms of our equilibrium constant K. So if we do the math here, at room temperature K would be 4.9 times 10 to the minus 7. That is a small number. That means that there is very little product at equilibrium at this temperature. Very little CO2 gas to be able to be used to make our bread rise.

But now if we calculate K at this value, at a negative number for delta G, K is now 55. So we have quite a lot of product then to be used to make the bread rise. So you can think about things in terms of delta G's and whether something's going to be spontaneous and give you the product you want.

But the equilibrium constant gives you that information as well. If it's a very small number, you have very little product. And if you want product, if you're trying to industrially make something, that's a very bad thing. But if you have a big value for K and this negative delta G0 value, then that's good if you want a lot of the product at equilibrium.

So equilibrium doesn't just matter for things like baking soda or forming ammonia from nitrogen and hydrogen. Chemical equilibrium applies to large molecules as well, such as enzymes in your body that are catalyzing reactions. So I'm going to share with you an In Their Own Words. And this is Nozomi Ando, who I will just mention is an MIT undergraduate-- was a graduate of MIT, was an undergraduate here in physics, majored in physics. And she is now a professor of-- do you want to guess?

AUDIENCE: Chemistry.

CATHERINE DRENNAN: Chemistry, that's right. She's a professor of chemistry at Princeton University. So here is a good example of what can happen. And she was very happy-- actually, I think 5.111 didn't exist then. I think it was 5.11 when she was here. But nonetheless, so here in her words about chemical equilibrium and the proteins she studies.

[VIDEO PLAYBACK]

-My name is Nozomi Ando, and I study a protein called Ribonucleotide Reductase, or RNR for

short. It catalyzes the reaction of converting ribonucleotides, or the building blocks of RNA, into deoxyribonucleotides, or the building blocks of DNA. It's the only means of getting those letters for DNA, so it's important for DNA synthesis and repair and replication.

It knows, for example, when there's an imbalance in the pools of the letters for DNA or if there's a lot of letters of the RNA. And this controls the sort of state that RNR is in. And RNR can be in an equilibrium of active and inactive states that are sort of regulated by the alphabet soup in the cell.

When it's active, it's very compact, but then it has to make a really dramatic structural change to go into an inactive state. I have this imagery of Transformers because it's just so dramatic. So when it's active, it's compact, like when a Transformer's a car. And when it's inactive and it makes a sound, che, che, che, che, and then it expands into a robot.

The letter A, or adenosine, pushes this equilibrium from the active to inactive state. And it tells RNR, OK, we have enough of the DNA letters, so stop. For humans, it's really important to study RNR because it's the protein that is essential for making letters of DNA. So it's essential for DNA replication, which is essential for cells to divide.

And we want RNR to function normally for our health. But in cells that are dividing too quickly, such as tumor cells, we want to slow it down. So actually, RNR is a really important target for anti-cancer drugs. But also, because RNR exists in every organism, we can start looking at differences between different species. So for example, it could be anti-bacterial and not just anti-cancer.

[END PLAYBACK]

CATHERINE DRENNAN: So that's an example of how you have a shift between two states, an inactive and active state, that it's just a chemical equilibrium. And binding one thing shifts the equilibrium one way. Binding something else shifts it back. So chemical equilibrium-- a lot of nature works by just suddenly shifting the equilibrium between different states. So understanding chemical equilibrium is pretty important.

All right. So now we're going to apply stress to our chemical equilibrium, and we're going to talk about the principle of Le Chatelier. So here, a system in equilibrium that's subject to stress will react in a way that tends to minimize that stress.

And whenever I talk to MIT students about this, I feel like I need to really emphasize this point. MIT students experience a lot of stress, but often do not tend to react in a way to minimize that stress. They say, all right, I am just struggling with this double major. I don't know what to do. It's just so much work. So maybe I should triple major instead. No.

[LAUGHTER]

Ask yourself, what would Le Chatelier do? Minimize. Double major, single major. So this is a principle that can apply to your life. And I recently saw a chemistry major, and we had lunch. He was a former MIT student, and now he's a CEO of a company, small company, in Cambridge that's designing computer software platforms.

He was a chemistry major, and he's doing software design and building like little computer tablet things for restaurants. So I said, would you use your chemistry at all? And he goes, oh, I use some things all the time, especially Le Chatelier's principle. I'm all about minimizing the stress. So that was one thing that he really grabbed onto in chemistry.

So again, if you think about this, about minimizing the stress, you can predict the direction of the reaction. And in nature, this really works pretty well. So Le Chatelier's principle gives us a way to predict the direction the reaction will go if you ask, which direction will minimize the stress?

So let's look at an example. We're back to N2 and H2 making ammonia. So here's a slightly different plot than I drew over here. Now, this is a reaction sensitive to temperature, so the equilibrium constant's going to change with temperature. So every plot of this may look a little different depending on what temperature it's at.

But some things are the same. If you start with hydrogen and nitrogen, you'll have some of those to begin with. They will be above 0. But ammonia will start-- if you had no product, you'll have ammonia at 0 and it'll rise in. The other thing that should look similar is that as the reaction runs for long enough it'll reach equilibrium, and you'll have the lines level out. You'll reach an equilibrium state.

Now what happens if you stress that equilibrium state? So say you add hydrogen. And so this line adds the hydrogen, then it goes down. Then you're also going to use up some nitrogen as you're using up the hydrogen. You're going to shift it to make more product. Now, say, you make product. You add product-- sorry, you're adding product. It's going to shift to minimize

the stress, use up the product, have the product dissociate, and make more hydrogen and more nitrogen.

So here's the plot. Now, let's just think about what's happening at each step. So if we're adding more reactant, you have more reactant. You have too much reactant. Shift to minimize the stress, and you will shift the reaction toward product. Get rid of the reactant, use up the reactant. Let's get back to the equilibrium condition, minimize the stress.

Now, we can think about this in terms of Q and K again. So when you have reactants added, then Q is going to fall below K momentarily. And so that means-- if you think about our equation over here, recall this equation, this important equation. So with Q less than K, you get a negative delta G, and that's going to be spontaneous toward the right, toward products.

So again, you respond by making more products. You shift to the right. And you saw that over here. It's shifting toward products. You're using up the hydrogen. You're using up the nitrogen. You're trying to return to an equilibrium condition.

Now what happens if we add more product? And so when you're adding more product, Q is greater than K momentarily. You have too many products now compared to equilibrium. And when Q is greater than K, you get a positive delta G. And that means that it's spontaneous in the reverse direction, or non-spontaneous in the forward direction. So you shift toward reactants. You shift to the left.

So again, you can use this equation to think about what direction is now going to be favorable. And so here you added product. The product gets used up. It's shifting toward reactants, and you're making more reactants until you reach equilibrium again. And we have one minute left for a last clicker question. What happens when you remove products? All right. So let's just take 10 more seconds.

All right. So we'll just put it up over here. So that means that delta G is going to be negative in that case because you remove products. So Q is less than K, and so the reaction is going to be spontaneous in the forward direction. Delta G will be negative, and you'll move to make more products. Again, minimize the stress. You took away products. You need to make more. Minimize the stress. Do problem set 5 and minimize your stress.