[SQUEAKING][RUSTLING][CLICKING]

JEFFREY GROSSMAN: All right, how are you guys doing?

[CHEERING]

Thank you. You know why? This is such an amazing day. All right, first of all, we're going to talk about phases. And there are two lectures that I have introduced this semester that are new, that have never been seen before, until right now. One of them is this one. Right, Dane, thank you.

And then, check this out, this morning over the wires comes this news. And it's the Nobel Prize in chemistry this year was given to 3 of sort of the founders of lithium ion batteries. And the other new lecture that I'm giving this semester, later on, is on the chemistry of batteries. It's like coincidence?

[LAUGHTER]

I don't think so. Did someone in Stockholm know? I don't know. But it's really, really cool. And this is awesome. And there's so much great stuff online now about lithium ion batteries, because of this. I mean there was before, but there is a lot of press on this now. It's very interesting to read about. OK.

So now, what is happening? I say the word boiling all the time. Right. I've been saying it all the time, especially the last week. And I mean at home, too, I'm saying boiling a lot. And it's because it's a proxy for bond strength, intramolecular bond strength and boiling, I've been kind of going back and forth. Right.

But so, today, we're going to go, what is happening when you boil? All right. Well, you know that you're changing a phase, as you can see from this highly accurate drawing of water molecules in the gas phase, water molecules in the liquid phase. But what's really under the hood? And of course, the story starts with Otto.

Now who was Otto? Otto Von Guericke. Otto Von Guericke in the 1600s. He was like-- you know what-- Aristotle. Aristotle, he was all about like nature abhors a vacuum. Does it really? I don't know. I don't-- you know, does nature really hate? Like, let's try some stuff out. Let's try to make a vacuum, and see what happens.

So he figured out a way to take these two hemispheres, these two big bowls, and kind of pull all the air out. There he is. He looks very happy there. Right. And he's pulling-- because he's pumping the air out inside of here, and look he did. He went around the countryside in Germany, and he would attach-- gesundheit-- just attach eight horses to one side, and eight horses to another, 16 horses pulling.

Eight that way, eight that way, and they couldn't pull it apart. This is what he did. It doesn't look like he's got a lot of crowd, but maybe he did better. I don't see any houses either. I don't know where he is. But that's the experiment that he did all throughout the countryside.

And he called it the power of the vacuum. Nature abhors a vacuum, the power of the vacuum. Power of the vacuum. It's not the power of the vacuum. Oh, Otto, it's not the power of the vacuum, because vacuum has no power, and nature doesn't hate it.

But pressure-- so atmosphere has power. Technically, it has force. Atmosphere has pressure. What he showed in that experiment, and if there's one experiment I want you to remember from this class, it's not this one.

[LAUGHTER]

But this is a cool one. But what he showed wasn't the power vacuum, he showed the power of one atmosphere of pressure. He showed the power of pressure. And so we start. Pressure is force over area. Typical units would be like newtons per meter squared, right. And if you're going newtons per meter squared, then this is called the pascal.

And for reference, if you have one pascal, one pascal, it's about the pressure of a dollar bill, $1 resting on a table. I don't know. Why do I use that one? Because I remember it, $1 resting on a table. And so that's the force per area of the bill resting on a table.

So if you had 100,000 of them, right, if you had 100,000 of them, and you stacked them up, then that's one atmosphere. That's just a different unit. It's about 100,000 pascals. So the power that he was exhibiting in this experiment is the power of the atmosphere.

We've got-- we've got atmosphere all around us, pushing on us. It's pushing on us. And so, we got like a car worth of atmosphere, literally, on us, but it's pushing on us in all different directions. And we're pushing back. Luckily, we've got stuff inside of us, so we don't like implode. Right.

That force is strong, and that's a demonstration of how strong, how much one atmosphere is. Right. It's force per area, the power of one atmosphere. Well, that's also directly related to boiling. And how do we get from here to boiling? Well, so we need a couple of concepts, right.

So first, we need the concept of evaporation. And I'm going to use this-- I'm going to use this example of like water molecules, right. So this is a-- this is a glass of water. Right. And so, if I have a glass of water, I'm going to draw them like these little things-- yeah, and now molecules kind of near the surface-- well, it could leave. And that is-- that's a molecule, not a v, right, and that's evaporation, evaporation. Evaporation.

But, you know, the other way you could go is, you could say, well, what if there was this liquid here, and you've got these molecules, kind of like that. And you had one here. And it kind of comes back. Well, that's condensation. You guys kind of know this, right? Condensation. Evaporation and condensation.

Oh, but this is not what's interesting. I mean it's interesting, but it's-- but it's more interesting now if you close it off. And so, let's do that with these super amazing animations here-- they're there. Look at that. OK. So now what I've done, I've got a liquid and a solid. And in both cases, there's a liquid. You could have solid. Molecules can come off.

We'll get to that. How? Why? They can come off. But now I've closed it. Now if I close the container, what would happen if I start with nothing? So now I start with a liquid, and I close it in a container. You see-- look at the molecules. They're kind of coming off. And maybe they'll hit the walls or bounce around, maybe at some point, they'll come back into the liquid. Right. OK.

So evaporation went that way. Condensation went that way. Now-- ah-- now, if I were to plot, if I were to close this and plot, for example, a rate. And I say, well, I want to know how quickly evaporation is happening. It's just going to keep happening. I mean I'm starting with, you see in the videos-- these videos are real-- you see it doesn't stop happening. right. It keeps happening.

So if I have no vapor in there to begin with, well, then, OK, evaporation is going to go like this. All right, maybe this is time, it is time. And then-- so that's evaporation. But then, at a certain time, at a certain point, I'm going to have enough gas molecules, gas water molecules in there.

So it starts out at 0, and they're going to build up in there, and at a certain point, you got that. Right. And so, this would be condensation. OK. We're starting with the basics, condensation. Right. And then at a certain point, they reach what's called dynamic equilibrium, dynamic eqm, eqm is equilibrium.

So the dynamic equilibrium is really important, because what happens there is nothing is stopping. It's not the thing stopped. You can see it in the videos. I found it online, so it must be right. So like, you know, it's continually going. It's a dynamic process. It's a dynamic equilibrium, but they are equal.

That point that happens in these containers, you can measure the pressure. That is very important. Right. Because it's exerting a pressure on the container. It's exerting a pressure on the liquid. Pressure. Atmosphere. Pressure, inside of a closed container, vapor pressure. P vapor. That happens when you hit that dynamic equilibrium.

Now why am I telling you all this? Well, the vapor pressure is going to be how we understand phase change. And so-- and so-- you kind of get a sense already, kind of get a sense already, that this is going to be related-- this is related, you know, this depends on the bonding. Right.

So I just told you that water was really special. So that means that how it-- it's the bonding between the water molecules and how often can it get out of the liquid, it's related to the bonding. Right. So if you take molecules that we talked about on Monday-- these are just from Monday-- propane and butane, now you understand, right?

Why is there this difference in the vapor pressure of protane-- propane, and the vapor pressure of butane? Well, it's how they bond together. And you know exactly why. Because these are only able to do [INAUDIBLE]. And so, this has a-- has a higher surface area. We talk about this in terms of the boiling point, remember, as a proxy for the intermolecular bonding strength. And now I'm giving you the vapor pressure.

So this is the vapor pressure of these things is different. Those are measured-- those are given a kilopascal. So if it's water-- so the water vapor pressure-- right-- let's see, make sure I get this value-- yes-- P-- I'll just put a Pv for water, is about 2.3 kilopascal. So it's a lot lower than those. Right. Those are-- now, in atmospheres, and this is going to be important, 0.02 atmospheres.

Just switching around units, that's all. It's same thing. Gesundheit. Right. So now, if a liquid is- you can think about this as, well, this is a proxy already. I can understand this conceptually. Right. It's a proxy for the intermolecular bonding. If you take something like glycerol-- well, glycerol is-- let's draw glycerol just for fun.

Carbon, carbon, Oh-- let's see, it's got another carbon, OH, it's got another OH, look at all that, OHs. It's got some hydrogens down here. It's got another hydrogen in there, and it's got two more hydrogens out here. No, I am not paying attention to VSEPR right now. I'm just getting it on the board.

But look-- but look at that. Right away, you're like, oh, look at those hydrogen bond opportunities. Yeah. This could bond more strongly than water to itself, right. It's got all that hydrogen bonding. It's got even more than water. And so the vapor pressure of this is 0.01 pascal. So much lower.

And as a general rule, we like to compare everything to water. So if something has a vapor pressure that is higher than water, so it's called volatile, Pv greater than Pv water, it is volatile. It's a volatile liquid, like butane and propane. If it's less than Pv water, then it's called-- here it comes-- all those chemists-- non-volatile. Look at that. Look at that beautiful naming originality. OK.

But this is the thing, right, I've given you this concept of vapor pressure. I've talked about it a little bit. You can relate it to something you just learned on Monday. But I still have not come back to boiling. And that really is it. It's that boiling is actually understood by vapor pressure. That is how we understand boiling.

And so you can think about it-- let's just give it to you conceptually. Right. So if you had-- let's go back to those buckets there. Think of it-- we'll go back to like evaporation and condensation. But now, instead of being in a closed container, right, now I'm outside. So if I had, let's say I had this container here. And I've got all these waters in it. Right. But now I'm outside.

So I've got the power of one atmosphere, right. One atmosphere. Well, you know, now you see, like, OK, if a molecule is able to evaporate from the liquid, if a molecule is able to leave the liquid bonding environment, breaking the bond, that's why we can relate it to intermolecular forces, right. It's breaking those bonds in the liquid, and becoming a gas phase molecule. It's changing its phase.

Well, you know, then the question is, is it like going to-- what's it going to do? You know, is it going to kind of hang out here? And are the forces of the atmosphere, is the atmospheric pressure going to kind of hold it around there? Or, is it going to be able to leave?

Well, this is a very simple conceptual picture. Right. If you thought that the forces of the atmosphere were going to always hold it there, then you'd say the atmosphere is a container. And it's not. But, actually, evaporation only happens because a breeze can come along and blow this away-- right-- out in the air.

So those forces are super important, because as soon as the vapor pressure is greater than the atmospheric pressure, well, now you can get a sense of what-- the vapor pressure is greater than the pressure outside over that surface of the air here, then the molecule is just going to fly right out and power through it. That is the definition of boiling.

Boiling is a pressure effect. It's that the vapor pressure of this material, which is related to how strongly it's bonded in the liquid compared to the gas phase, that that goes higher than the surrounding pressure. That's the definition of boiling. So you can see why water doesn't boil. Like having a bucket of water, and I watch it, it's not going to boil. 0.02 atmospheres is not greater than one. Right.

But so I haven't gotten there yet. I haven't made it boil. But I want to make a boil. And that is where we go next, because that is really what temperature is. See, the vapor pressure is really a link between temperature and boiling. It's the length that you need to understand between temperature and boiling.

The reason why-- when you increase the temperature of the water, the reason why it boils is because you change its vapor pressure. That's why it boils. That's what boiling is. OK. It's when you get it above the atmospheric pressure. No. OK.

So the-- so we got-- OK-- so are we-- where-- there we go-- we'll come to that in a minute. So if I were to plot now, the vapor pressure-- let's do this carefully here. OK. I'm going to make a plot of the vapor pressure, Pv, right, versus temperature. So I just told you there's a dependence on the temperature. Right. So versus temperature.

Let's go to-- oh, I don't know-- let's go to room temperature. So around here, and I just told you that at 0.022 for 25 C. OK, good. So 0.02 for 25. 0.02 for 25 C. So that's the vapor pressure. That's the temperature. But I also told you that boiling-- if I'm at-- so this is 0-- but I also told you that boiling-- so this would be like one atmosphere.

Let's suppose the units are atmospheres. Right. OK. So I told you that boiling happens-- now we know that water-- we know what the answer is, right, up here. That's got to be boiling for 100 C, right. That's where the water goes above or meets the atmospheric pressure, and then goes above it. So once the vapor pressure-- so if this is the vapor pressure of water, this is what we have so far. But we need more than that.

So you can say, well, why is it-- so we know this-- why is it that if I go to Denver, the-- so like this is in Boston-- this is in Boston. This is Boston. But if I live in Denver-- how many people live in Denver-- well, I mean you're not living there now, but-- but, welcome to Boston. Oh, ho, ho.

And if I go like this-- now see the thing is, now you boil at-- this is-- this is in Denver. So you guys don't have as much air in Denver. The pressure in Denver is 0.83, 0.83 atmospheres. But I just said that, you know, you boil when you have the vapor pressure is higher than the atmospheric pressure.

So that means, whatever this curve is, it means that I'm going to boil at a different temperature. And it turns out that's 100, and if you're in Denver, it takes longer to boil potatoes, because the boiling temperature is 95. That's why. But it's all because of this. It's all because of how evaporation happens as it relates to the atmospheric pressure.

Yeah, but-- the thing is, though, what we need-- and by the way, you know, you can go the other way, too. Right. You go up. If you go below sea level, then the atmospheric pressure is higher. So if you go to the Dead Sea, your water boils at 101.4. If you go to the top of Mount Everest, it's 70. You never boil potatoes there. Well, you can, but they'll never cook.

So how do you-- how does this get connected? That's really the question. What is-- so what is this? Right. Where does that-- where does that dependence come from? And that's where we need something called kinetic theory. Now I'm not going to teach you kinetic theory. And I'm not going to say that you responsible for knowing it. But I'm showing you what it means.

What kinetic theory does, which you would learn in, say like a thermal class, right. Well, kinetic theory does is it gives us this relationship between how many molecules have a given energy, in this case, it's a kinetic energy, right, and the temperature. That's what can in theory does for us.

So let's see, if I had-- if I had to write this in a kind of just simple dependence way, what it's telling us, in, general is-- so this would be kinetic theory-- kinetic theory of gases, if you want. That's what we're talking about there. OK.

This would give us the following. This is what you get is that the log of the fractions-- fraction of molecules with some energy, E, OK, with energy-- I'll write it out just so there's not confusion-- energy, E, the log of that, the natural log of that, is going to go as that energy divided by temperature.

It gives you a connection between what I'm plotting here really. That's where that plot comes from. OK. The fraction of molecules with kinetic energy, E, goes as that energy divided by temperature. The log of that. That's what kinetic theory gives me. That fills it in, except that I-- it still isn't vapor pressure. Right. This still isn't vapor pressure.

And so for that, you need one more piece. And that is given to us by Clausius-Clapeyron. So the Clausius-Clapeyron equation-- ah-- Clapeyron-- does not-- I'm not going to derive it. You don't need to know how to derive it. But it goes from the kinetic theory to vapor pressures.

So this one tells us that the log of the vapor pressure equals minus-- and here we go-- this is the next really important concept-- the heat of vaporization-- and we'll get to this in a second-- divided by RT plus a constant. So that's the ideal gas constant. That's the temperature. And this is the enthalpy or the heat of vaporization. So it's the energy to vaporize a liquid. Good.

Now, often, this is given in per mole. So it would be like the energy of vaporizing a mole of a liquid, but it's an energy. It's an amount of energy needed to turn it, whatever you got to put into it, to turn it into a gas from a liquid. That's that enthalpy of vaporization. And R is the ideal gas constant, in this case. You know, so if it's R-- if R is in the equation, it's got to be per mole, right, because it's R. Remember, we've already talked about that. Since it's R, it's the ideal gas constant.

So that's, now, OK, so that is how you get between-- that is how you get this relationship. You get it, and, again, this is an important equation. You get it from the kinetic theory gases, but this is what the outcome is for us. This is what matters, is the Clausius-Clapeyron equation.

And so, and so now, you can actually now, if somehow you knew how much energy it took to vaporize a mole of this stuff, well, now you've got a curve. That's all I need to know, right, is how much energy is-- so now, I can say, well, let's see, that's OK-- I got-- I got, maybe butane is going to go like this. And then glycerol might go like that. Maybe this is propanol. Right. Propane, glycerol, and so on. Yeah.

But I'm still missing this concept. So I can now get dependence, but how do I get delta H? That's the next piece of the puzzle. Right. How do I get delta H? OK. So, you know, so far I've got this really cool result, right. I've got that P-- so far I've got that-- I'm going back to bowling-- I'm going back to boiling. Right.

For boiling, I know that-- so the ion-- so P vapor is related to the IMFs, right. I know that P vapor is related to the temperature. And I know that the atmospheric pressure is important. Right. Atmospheric pressure is related to boiling. Yeah. OK. That's redundant, because you've got to get higher than the atmospheric pressure to boiling. That's how we defined it.

So, but, again, this specific dependence is what I wanted on temperature, and I have it, except that I now need to know what this delta H is. OK. So here we go-- why is this related to pizza? Everything's related to pizza at the end of the day. But see, now we need to talk about delta H, and what it means.

And you know, the way-- the easiest way to show what it means is to plot how this relates to phase change. How heat, right, that's how we're changing phase here with temperatures. We're changing the temperature. We're adding heat into the system. What does it do? What does it do?

Well, you know, so one type of plot that you can make for that would be temperature versus thermal, thermal energy input. Now a lot of times what you'll see in textbooks is they call that q. It's an energy term. So it's like in joules, right. Let's say. But it's how much thermal energy you put it in, then you can draw it out like this. Right.

So if you draw it out, then you would have your phases. So this would be like a gas, a liquid, a solid, and these would be phase changes. This, and that. So I'm adding thermal energy into the system. It's raising the temperature of the system. That's why these are going up. But the solid is getting hotter, and then, all of a sudden, it melts. OK. Good. So it melted. It melts.

And then it's a liquid, and, all of a sudden, it vaporizes. Right. Vaporize. But you can go the other way. If you go the other way, it's condensation out here. All right. Or freezing in here. Melting, freezing, vaporization, condensation. But I like to draw it this way, because this way we're going to see very nicely what delta H is.

So here's-- now I'm going to draw it with temperature on the x-axis. So temperature-- and this is going to be H. And I'm drawing this as the thermal energy-- let's see, I'm going to write it in the same terms-- right. So this is thermal energy. Now in this case, thermal energy, we use this thing called enthalpy, enthalpy.

But you don't need to worry too much about it, for this class, you can think about it as the thermal energy content of the material. So if I have a change in that, wells, it's because I've lost or gained thermal energy. Right. If you take a thermal class, you'll learn a lot more about enthalpy. And that this is true only if you have constant pressure conditions, but that's not relevant here.

I just want you to know, this is like this. OK. But this is the total thermal energy in your material. And by the way, you know, q, heat, the word heat is q. That is the definite-- heat means thermal energy transfer. Thermal energy transfer, a flow of heat. That makes sense.

But you cannot say that this has 10 joules of heat. No, don't say that. You will sound like you're from another school.

[LAUGHTER]

You don't have-- you have that much enthalpy. Yeah. But not that much. Heat is a transfer of thermal energy, please. OK. Now here we go. All right. So let's suppose I were to draw this. Well, now I'm going to have the same-- I'm going to have phases. Right. And so here we go, here we go, here we go. And this is my gas, and this is my liquid, and this is my solid.

There's a couple of cool things, right. So now-- so that's the thermal energy, H. But so now, if I were to take part of the solid like that, and just look at it, while it's a solid. I'm adding thermal energy to it. I'm increasing its temperature. We got a way to relate those things. It's called the heat capacity. All right. It's called the heat capacity.

So if I look at that, OK, I can do that. Let's go here. All right. If I look at that, then I can relate those by-- well, I'm using q, because it's a change, right, that's the change in the enthalpy. All right. That's the energy transfer. It goes as the heat capacity times the change in the temperature. That s just means it's the heat capacity of the solid.

Literally, that's the definition of heat capacity. Right. This would be like units of joules per kelvin. OK. So and, example of heat capacity, I love pizza, and so an example of heat capacity is the following. You burn, hopefully, you don't. But if you take it out of the oven, and you take a bite, right away, you will burn the roof of your mouth on the cheese. On the cheese, because cheese has a lot of water in it compared to the crust.

And water has a really high heat capacity. What is heat capacity? It's the capacity of it to have thermal energy. And the water has a really high one, and so, when you take this out, everything's at the same temperature for just an instant. They're all at the same temperature. If you eat it right here, please don't, you burn your mouth on the crust the same way as on the cheese.

But since you're waiting, now the crust gives out-- it doesn't have much heat capacity. So it all goes away. All that thermal energy leaves very quickly from the crust, but not the cheese, because it's got so much more. And if we want to get a little more serious about it, we can talk about the oceans. And so, I put this in here. Because we talk a lot about climate change, what you're talking about are the temperatures of the atmosphere. We talk about that a lot.

But what gets sort of shoved aside is the temperature of the ocean, and that is rising, and it's not rising that quickly, luckily, but that's because the ocean has such a high heat capacity. That's where most of this thermal energy is going. It's going into the ocean. Right. And we need to start talking about that a whole lot more, because it's a lot higher than it used to be.

This is the historical reference, pretty flat. A volcano erupts, 500 years ago, you get a blip this big. Now we're here. Right. So the amount of thermal energy this ocean is holding for us is very high. OK. Now so that's high capacity. That's kind of an important part, but, look, the other thing we have here-- ah, I keep hitting the wrong-- is this.

And that's what I want to talk about. This is that delta H. Now in this case, it's the change in energy of fusion, because you're fusing to make a solid. In this case, it's called delta H of vaporization. You could call it condensation vaporization. That's the thing we needed. And now you see it so beautifully. All right.

You see it so beautifully right there, because, now, you see-- so what is it that causes me to need to put a lot of energy into a material to change its bonds so that it can vaporize? It's all about the bonding strength. It's all about the-- everything comes to the bonding strength. Right. And so I have these kind of two things, right.

Within one phase, within one phase, I can relate delta H to delta T. And that relationship is here. I'm in the solid phase. I can relate them. That's actually called sensible heat. Sensible heat. It's the energy that you would give off when you cool, or it's the energy that you'd absorb when you heat.

But now I've got this other thing, which is at the phase change, at the phase change. Well, now, the temperature doesn't change, but I'm putting all of this energy in, or getting it all out. That's what delta H is. All right. So delta T is 0, because the temperature is constant. And delta H is sort of, generally, called the latent heat, latent heat.

If you're talking about the phase change between a solid and liquid, it's the heat of fusion, heat of melting. If you're talking about between the liquid and gas vaporization, or condensation. So that's that energy change. And it's related to, well, what we've been talking about, but here it isn't a nice diagram. Right. Now, where do we go from here?

Well, there's something so cool about this graph that I can show you really easily. I can also tell you how to do this, which is that things can-- these are the phase change transition temperatures, but things don't need to transition at those. In fact, oftentimes, you might get something like that. Not as much in the gas, but liquid to solid, you certainly can get it. This is called super cooled.

Why? Because, for some reason, [INAUDIBLE], for some reason, it's remained a liquid below its phase change temperature. Here's water, super cooled, poured into a glass. It's one of the most fun things you can do in life. There it is. It's a liquid. As soon as it hits the glass, it's like, wait a second, I want to be a solid. I'm over here. I'm over here. I need to be a solid.

And you can do this. You can make-- I can tell you how to make it. You can make super cooled water at home. It's really fun. But that was what you're doing there, now you know, is you're bringing this down below the phase change. It's remaining liquid.

Now all of a sudden, it hits the glass, and, but what am I doing? I need to solidify. All right. OK. This is the reason why all this matters is heat is a big deal. 90%, you now, this is-- you can study this. There are wonderful charts to study. But depending on what you want to do, heat is almost always involved. Heat being a transfer of thermal energy.

And so 90% of the current energy budget, basically, goes through the middle in some way. That's a lot. But then, if you look at things like this, which you shouldn't, because you can't read it, but if you blow up-- these are all the inputs on the left, all the ways we make energy. Here's how we use it. And here's the result.

Here's where I want to get the punch line is, rejected energy. Whatever these units, doesn't matter, it's 60%. What is rejected energy mean? Wasted. All this precious fossil fuel stuff we're burning and all the solar energy we're collecting, anything we're doing here, 60% of it goes into heat, wasted heat.

And so the reason why this matters is because there are these materials that we can use to try to capture some of that. Right. There are materials we can use to try to capture some of that, and it all comes down to chemistry. These are called phase change materials. And if you plot the melting temperature, that's this-- that's this melting temperature, right, T_m, versus how much of its delta H, how much delta H you have, you get all sorts of classes and materials.

But, you know, so you might want a certain melting temperature from wherever you're operating, but then you might not get a high enough delta H, or you might want really high delta Hs, but then the thing doesn't melt until 700c, which is way too high. So we got to fill this out. We need materials here. We need materials here. That's a call to chemistry. That's a call to chemistry.

So then, it all has to do with the phase change. It's a weird name. These are called phase change materials, but that's true of all materials. But, here, when you're talking about storing thermal energy, we use the term phase change materials. OK. Now the last concept, and I'll just touch on, and then spend a little time on Friday on, is the concept of a map.

These are maps. These are maps. But you can go even bigger. And back in the day, back in the day, PVT was a big deal. You went to a party, you were going to talk about PVT. You were. And you know, and so, everybody had their own way of looking at it. Gay-Lussac, P versus T.

But look, if I do P versus T, and hold volume constant, right, I might get like a line there, like P-- if this is pressure, all right, and this is temperature, you might get something like P versus T is a straight line, if you hold other stuff constant. Right. That's good. But that's not what I'm interested in today. I'm interested in where the phase boundaries are. Right.

This is a materials map, and it gives me the boundary, that's what this is all about. That's what this was all about. Because, look, that's why I went through all this. You boil. What is boil? Phase change. Phase change. You boil when Pv is greater than P at [INAUDIBLE].

So that is the phase boundary. It's when you have both the gas and the liquid coexisting. Where is it? Ah, there it is, because I can get my pressure of vaporization, if I only know this, or vice versa. Right. And then I get that. It's up here. That's the phase boundary. Oh, yeah, that's all cool, but this is where it's at. I want to know where stuff changes phase. All right.

And, of course, you can go further. And I've got to show you something here. There's a couple more things about-- these are called phase diagrams, right. These are maps of materials, and they are beautiful things, and incredibly important. And there's a couple last points, one of which I want to share now, and then the other I'll share on Friday.

But the one I want to share now is the triple point. So look at this. Here's the gas liquid boundary line. That's what we've been talking about. You also got the solid liquid boundary line. That's here. You got gas liquid here, delta H- here's delta H. You also have sublimation. So you go straight from a solid to a gas, and you have deposition.

You go straight from a gas-- but look at this. There's one point where they all co-exist. One, and only one where all three phases exist at the same time. That's pressure. That's temperature. And what's so cool about this is there's only one for this material, and so it, actually, is a really nice way to have a repeatable, reproducible point, which I'll get to in a second.

But first you got to see this. This is a triple point. And what they're doing is they're lowering the pressure and the temperature. They're changing both. You could just do this with one or the other. Yeah. There it is. OK. How did they get it to boil? Lower the pressure. You now know why. You lower the atmosphere inside the container. You lowered that, so you went like this. You went down.

And all of a sudden, you converted some of the liquid to gas, because you went down in pressure. And then you said, well, OK, but wait a second, I'm only seeing these two phases. So let me do this, and go over that way, getting closer, closer, OK, what's happening now? Then do this, right? And you're trying to find this special point here. And they're going close to it. They're going close to it.

And then a couple minutes go by, that, look, it's freezing. It's freezing. It's-- but wait a second, I thought I said triple point. And I did. It's boiling and freezing at the same time. It is boiling and freezing at the same time. Yes, that is actually what all materials can do. That is incredible.

Somebody just said what. Thank you for making my day. That makes my day. That is a triple point. We got a one or two more things to talk about phase diagrams. We'll do that on Friday. Have a great rest of your day.

This is where we left off. And so, I want to start right here. And because a triple point is just so incredibly cool, I've found another random video, and I thought I'd show that. So this is where we left off. And here it is. And what they're doing, they might be controlling the temperature, they're definitely controlling the pressure. All right.

They have a liquid in here. Now watch what happens. We know now from the phase diagram that you can find those boundaries on the diagram. So they're lowering the pressure, and at some point, yeah, nothing's going happen. There. OK. Now it's playing.

And so at some point, this is going to start to freeze, because I'm crossing over the boundary. All right. I'm crossing over the boundary. And so when you cross over, when you get to a phase boundary, you've got two phases coexisting. Oh, and there it starts freezing, so you've got the liquid and the solid phase coexisting, but there's-- but now can it coexist with, also, the gas.

And at one point of pressure and temperature, it can. There it is. It's kind of boiled a little bit. I like this one, because it kind of goes big boil. So here it is. It's trying to freeze, and now it's boiling. And now, watch, it'll refreeze. All right.

And then it's going to boil again. It's all three phases at exactly the same time. So you can Google it. There it is. It boiled, again. And then it's going to freeze, again. It can do all three at once. That is cool. All right.

And how do we get there? How do we get there? Well, what I thought-- on Wednesday was the first time I told you there's two new lectures this fall. That was the first time I gave that one. That was one of them. The other one will be on the chemistry of batteries.

And because it's new, I, you know, I want to make sure that you feel your oneness with the concepts that we learned. And so I have made some summary slides for you, that I'm not going to take time to go through. But that tells you how we got to here. That's the end of the lecture on Wednesday. And I want to make sure that how we got there, and what the key points were.

There are five key points. And I've created a slide here, again, this is Wednesday, so I'm not going to go through this in detail, but I'll leave this in your notes. One point was boiling and vapor pressure, and what that means. The second point is the vapor pressure versus temperature, which we get from kinetic theory and Clausius-Clapeyron gives us the relationship between that vapor pressure and the temperature through this thing called the vaporization.

The thermal-- this, remember, this change in the thermal energy. Right, this delta H, this change in enthalpy. OK. And by the way, you can-- these are these like vapor pressure versus temperature curves we drew. You can take two of them, for example, that gets rid of the constant in that equation, two pressures, two vapor pressures, two temperatures. You can calculate that vaporization energy. That's that interrelationship we talked about.

The third point is the thermal energy that delta H is related to the phase change. It is the energy of the phase change. Right. So we talked about that. And then, I couldn't find a good-- so I just drew this myself here with fancy PowerPoint-- and what you can see here is a heating curve. And you can draw it in two different ways. You can draw it as the temperature versus enthalpy, or you could draw it as the energy input versus temperature. These are equivalent. It's just not mm, and mm. Right.

And then, in this, you can find all so much information, like there's that delta H between the phases, right. Here is the liquid phase, the gas, the solid. And then, if you're in a single phase, we talked about that change in enthalpy related to the heat capacity. If you're at the phase change, the temperature doesn't change, but you're putting a lot of energy into it.

And that, finally, got us to phase diagrams. This is where we left off. So I'm leaving this all here for you to look at if you want. It's kind of a summary of Wednesday's lecture, the key points. This is that two-phase curve. Right. And this is-- on these curves, remember, the phase diagram is a map. It's this beautiful map of the material. Right.

It's a map of the chemistry of the thermodynamics of the phases of whether it's a gas, liquid, or solid. And these coexistence curves are where two phases coexist. Right. So you see at these temperatures and pressures, and that's why we started the lecture talking about vapor pressure, because the vapor pressure is such a nice way to get into the idea of how a material changes phase from liquid to gas, which is what we did.

And then, we also talked about how that triple point, that's the video that we left off with. And then there's one last thing I didn't mention, which is that there's this critical point here. What's happening out there? What's happening out there is really interesting, too. There is a temperature, right, above which-- so there is-- if you keep going up in temperature, there is a critical point where this boundary just disappears. It stops.

The liquid, there's the liquid, and the gas, and now there's no line. What happened? What happened is, you've reached a temperature where it, basically, can't exist as a liquid anymore. If you keep going, the thermal energy is just too high. But the pressure is too high, too. So this is a different phase.

And this phase it out here is called supercritical. This is called supercritical, where the material has properties of both. That's really cool. Right. So this material has maybe properties like it can flow like a gas, all right, but it's got a higher density than a gas, but it's still lighter than a liquid. And it's that same material. It's that same material, whatever it is, water. Right. That's called supercritical.

So these lines can end like that. And that's called the critical point. Now, there, you know, there are many reasons why these maps are important. I thought I'd give you one which is strawberries. Because, if you go and you buy-- some of us like to buy like dried fruit-- but if I take strawberries and I leave them out on the counter, they're not going to look like this.

You know, how do you know that? Well, you know that from the phase diagram, because if I draw my temperature and pressure, and I've got some phase diagram that looks like this, right. And so now, you know solid, liquid, gas. Well, you know, I could-- if I just-- if I just tried to dry these things out, that's this. That's that.

But so, I'm-- maybe, maybe, I'm going to, or maybe I want to do this. Oh, that's even worse. What if I go like this. I'm going to heat the strawberries up. I'm making jam. No. That's not jam. Right. That's nice kind of maybe still solid, it's not like mushy and gooey. How do I get it? Well, you use the phase diagram, of course.

You go-- because this kind of phase change, going across that phase boundary, it creates a lot of disruption in the material. It creates a lot of disruption. You saw it, right, this boiling. So if you do that for the water inside of the strawberries, you're going to destroy all the framework of the good stuff, right.

So what you do is you freeze it. You go this way, and then you go that way. If you do that, now the solid as it sublimates is much less of a kind of violent, disturbing phenomenon to the surrounding matrix of strawberry. And so you get these freeze dried fruit that way. It's much less disturbing, destructive way to get rid of the water. That's called freeze dried.

You can also go the other-- if I go out here, I've got all sorts of uses to use this supercritical region. I've got all sorts of ways where that becomes useful. That's how coffee is decaffeinated, because it's a different phase of this material. So the phase maps give us this information. They tell us how we can engineer properties and materials by engineering their phases. Right.

And if you pull out another thing, you know-- a lot of the textbooks and what we did is we look at the phase diagram down here. But if you pull out, you also can see how complex materials are. All right. This is water. We've been focusing on this part of phase diagram, now you plot it. Water has 17 phases, 17, not 3.

And just two years ago, the 17th was discovered, and it's a very, very light form of ice. All right. So you can get all sorts of really important information about materials by looking at their phase diagrams. You can compare different materials, like this is the phase diagram of CO_2 versus water.

And you can learn about sort of where their critical points are. You can see all the slope here in water is-- most material slope like that. Water slopes like that. And that's because of hydrogen bonds. Because the hydrogen bonds make it so that the solid is less dense than the liquid, which means, if I squeeze on it, I can-- then I go to the liquid. That's really weird.

If I squeeze on this, it's more soft. But here, it goes-- it can cross a phase boundary. That's very unique to water, and the fish are very happy about that, because ice floats. OK. Now, OK, now, so that's the kind of last part of the phase diagram stuff.