

MITOCW | MIT3_091F18_lec29_300k

We're going to pick up where we left off on Monday.

And where we left off was we were dissolving stuff.

We talked about how things dissolve.

Why does something dissolve maybe in this solvent versus that?

And how does it dissolve?

And what makes them-- and then we say, well, what if we had something and we wanted it to dissolve all the way until it couldn't dissolve anymore?

Well, that's called saturation, right?

And we talked about that.

We talked about how if you had a general reaction, maybe we had something like this, $aA + bB$ goes to and comes back from $cC + dD$, then this reaction can happen in both ways, right?

And so the general concept that we called a reaction quotient, reaction Q , is equal to-- remember, it's equal to the concentrations raised to those stoichiometric coefficients.

And it's the products over the reactants.

OK, so that's what we-- now then, we say, well OK, that's a thing that has to do with maybe-- where's that reaction going?

What's dominating it?

And remember, these are the stoichiometric coefficients, and they're related to literally the probability that that reaction can happen.

OK, and so if you want an intuitive reason for the exponents, you can think about this as, well, if these all have to get close enough in a given volume to react, then that's where those exponents come from.

But unlike the rate law-- and I said this Monday, and I'm saying it again today-- they are not things that are experimentally measured.

Instead, they just come from the reaction itself, OK?

So we're just getting back into the right mood here.

But then we said, well, but look.

OK, this thing is reacting.

It's going back and forth, whatever.

But what about once it reaches equilibrium, all right, equilibrium?

And when it reaches equilibrium, it's the same thing, but we call it K_{eq} , in equilibrium, eq'm, equilibrium, right?

So that reaction quotient has a very specific value, a constant, a constant, once this way is happening in the same amount as that way.

Right, that's the saturation point.

Well, if you're dissolving something, that's the saturation point.

OK, so it's where the precipitation-- remember, we wrote this as dissolution and precipitation when we were putting stuff in solution.

And that's where we got to.

And we went a little bit further, and we had a specific example.

And that's where I want to start today, and that's silver chloride.

OK, so here we go.

So we have silver chloride.

And if I write this all out for silver chloride, I've got $AgCl$.

And I've got a little pinch of it, like a little speck of silver chloride.

And I'm putting it into a beaker.

So it starts out as a solid, and I add some water to it.

And the water is l, because it's everywhere.

It's the liquid.

It's the solvent.

OK, it's the liquid that I'm dissolving it in.

And we say, well OK, that's going to go like this.

So that reaction, it's going to give us silver ions-- remember salts, we did salts-- and chlorine ions, and back to some more water.

OK, that's H₂O.

Now, the aq means aqueous.

Because it's a salt, right, these are now dissolved ions, because it's a salt.

So we write them like that.

And we know we got water everywhere.

So oftentimes, we leave the water out.

Oftentimes we don't really write the water, because it's on both sides.

And anyway, the water is just surrounding stuff.

It's not necessarily being consumed.

The concentration of the water isn't changing.

The water is the water in the liquid phase.

Well, OK, then we went even further and said, well, if you write a K_{eq} for this, then that would look something like the concentration of silver plus in solution times the concentration of chlorine minus in solution.

And that would be divided by the concentration of silver chloride as a solid.

And then we said-- but hold on.

But hold on.

In the solid phase, the concentration isn't changing, right?

As a solid, the concentration of silver chloride is silver chloride.

It's a constant.

And so then we said, well OK, if this is a constant, this is a constant.

Then we can also absorb it into the equilibrium constant.

And I somewhat confusingly wrote it as K_{sp} in both this form and the form that we know and love for K_{sp} , which is simply literally the solubilities of these times each other.

Right, and that is K_{sp} .

So I wanted just to clarify this, because a student asked a very good question, which is, which is it?

Is this in K_{sp} or is this in K_{eq} ?

Well, the big full picture for K_{eq} , you'd write it all in.

But the solubility product is focusing in on the dissolution of these ions.

And so you wrap this into the equilibrium constant, because it's another constant.

So that's the solubility product.

And that's where we got to on Monday.

The solubility product is a special equilibrium constant, where what we're talking about is-- what are the concentrations when you've reached saturation of this dissolution precipitation reaction, OK?

That's the solubility product.

That's K_{sp} .

So you wrap this in, and you get K_{sp} .

You can think about it as, well, maybe you could have called this some K_{sp} in a pre- K_{sp} .

And then you wrap this in, and then it's still just [INAUDIBLE], because it's still just a constant

representing this reaction.

OK, I just wanted to make this super clear, because this part leads to confusion often, that this gets left out, because it's constant.

It doesn't change.

OK, good.

Now we get to that, because this is a constant.

OK, so what does that mean?

Well, it means that if I dissolve this thing in water, then the equilibrium, the equilibrium is a fixed number.

This equilibrium product is fixed.

And that has a very important meaning.

It's the green curve.

You see, it says equilibrium.

And so what that means is that if I'm at point B, well, there I go.

I've got my-- we said K_{sp} .

K_{sp} for this one equals, what was it, 1.7 times 10 to the minus 10th.

Units are coming.

And so that's the K_{sp} for silver chloride and H_2O .

Oh, and you have to say at some temperature.

It will depend on temperature.

You often will just be at room temperature, 25C, for example.

So often these get quoted just at room temperature.

OK, so once I know that, then I know that that's equilibrium no matter what, right?

And so that means that if-- and then we solve this.

You said, well OK, if I dissolve silver chloride, and I had nothing else, then the K_{sp} would equal the concentration of silver times the concentration of chlorine ions.

But those are equal.

And so we let each one of them be x , then it's equal to x squared.

And so we know that the concentration of each of them is 1.3×10^{-5} let me write it exactly-- 1.3×10^{-5} .

Now, it's a concentration, so you know that it's got to be capital M, moles per liter.

And since that's the case, since that's where you know what it is, then now you can back up.

Because this is the concentration of Ag ions and it's also the concentration of chlorine ions, then you know that the units of K_{sp} must be M squared, all right?

Units in this case, moles per liter squared.

Must be, because otherwise the concentration isn't right, the units of concentration.

OK, now we're going to play with this.

OK, so I'm at B. I'm at point B. You see that?

There I am.

And I've got my concentration of chlorine, concentration of silver.

And I'm here.

And so that's 1.3×10^{-5} .

OK, now if I didn't have as much, if I weren't at the saturation point, then I could add more.

I'd be here.

And I could add more solid silver chloride.

And eventually, I'd reach this equilibrium.

But if I added more than that, so now I'm adding way more than I can, then it's just going to be a solid.

And that's kind of boring.

So now I'm up here.

But maybe I add ions in.

So these are the ions.

So maybe I add a whole bunch of ions.

Well, they're just going to precipitate.

I can't be here and be in equilibrium.

That's the point of the green curve.

This K doesn't change.

This K is this K. That's what the green curve is.

It is that being constant.

That's very important.

That is equilibrium.

Now, there is a way though to move around on it, because if I did, imagine now-- look at point C.

Imagine now that I've found some source of chlorine ions and I add a whole bunch of them, right, a whole bunch of them into the solution.

I don't want to add silver and chlorine together.

I just add chlorine.

Well, you can see what's going to happen.

I have to get to equilibrium.

And so in order to stay in equilibrium for silver chloride in solution, then I must lower my

concentration of silver, because I added all this chlorine.

The only way to do that is to eat up some silver and precipitate.

That's the only way to do that.

And that has a name.

And that's what I want to talk about next.

That is called the common ion effect.

Now, this is an example.

This is an example of a broader principle which I mentioned Monday, which is Le Chatelier's principle.

OK.

And that is that position of equilibrium will move.

It's still equilibrium, but it's moving on that green curve to counteract change.

This is a very general principle.

It applies to many, many things, not just concentrations.

It applies to changes in pressure and other changes that you make in a system.

Here we care about concentration.

And in this particular case, we're going to see it for what happens when you add ions of one particular type and not the other.

That's the common ion effect, but it's a very general effect, that it resists the change.

Now, let's see how that works.

If I want a source of ions, let's say I want a source of chlorine, one way would be to add a whole bunch of sodium chloride.

So let's suppose that I have-- OK, so as an example, I'm going to add to this nice equilibrium that I reached here-- so I'm at point B, OK?

And I'm going to add 0.1 moles per liter of sodium chloride.

And this nearly-- and so we're going to say it does-- it nearly fully disassociates.

And we'll be talking about that later today.

And so that means that NaCl basically goes to Na^+ plus Cl^- , both in solution.

So if I put 0.1 moles per liter of sodium chloride into this container of the silver chloride, I'm basically going to get 0.1, because of the stoichiometric coefficient here, 1, 1, 1, right?

So if I put 0.1 moles per liter of this, I'm going to have 0.1 moles per liter of chlorine ions that I just dumped into the container.

But the sodium is not going to do anything.

But look at the chlorine.

That's going to mess with my equilibrium.

But K is a constant.

So let's see what happens.

We can do it visually.

But we can also do it back when we use the ICE table, right?

So now we're going to go ICE again, AgCl.

And here's Ag^+ .

And here's Cl^- .

So now I've added the-- OK, so my initial is that this is solid.

And remember, this is the solid that I'm putting in there.

And then my initial is that I started in that nice equilibrium at point B.

So I've got 1.3×10^{-5} .

And over here, I've got the same.

But now the change, now the change.

OK, so we're going to add some chlorine ions, 0.1.

So I've added 0.1 moles per liter of chlorine ions from the salt, from sodium chloride.

Yeah, but now some of that is going to react, because I just said, Le Chatelier's principle tells me that now it's got to counter that.

And you can see it right here too.

I'm adding a whole bunch of chlorine.

Look at how far I'm going out.

0.1, that's literally all the way out to here.

And so the amount of silver has got to go down.

And what's going to happen is it's going to consume some of the silver and form precipitate to counter all this addition.

It's going to run the reaction the other way.

So it's going to lose a little bit there from the 0.1.

That's going to react with the silver ions.

So that's going to also lose, and this is going to gain.

That's how much that's going to gain.

So those are the variables, right?

This is the ICE table.

That's how much it's going to lo-- these are going to react to give me some precipitate.

And so the equilibrium is more solid forms.

That's not a good E. More solid forms.

And over here, it's $1.3 \times 10^{-5} - x$.

And over here, it's $0.1 - x$.

That's the equilibrium condition.

I've added in a common ion.

This is the common ion effect.

Yeah, so now if you do the math, then what happens?

Well again, K_{sp} is the same.

It's the equilibrium constant.

So it's the constant.

So that is going to be the same value.

It's going to be 1.7×10^{-10} .

But you see, K_{sp} is also equal to the concentration in equilibrium of the silver ions times the concentration of the chlorine ions.

So it's going to be $0.1 - x$ times $1.3 \times 10^{-5} - x$.

And then we don't like doing all this math, and so we simplify.

Then we can simplify, because look, you started with silver ions in a very small concentration.

You can't take away something you don't have.

And so x has got to be somehow this or less, right?

And so x is a small number.

And this is a really big number in comparison.

And so we like simplifying our lives, right?

And so we like saying that that's like 0.1 .

We know that.

That'll make the math a lot easier.

It'll make the math easier.

And so now we can say that this is equal to 0.1 times 1.3 times 10 to the minus fifth minus x.

And then we get that x equals 1.7 times 10 to the minus ninth.

That is the common ion effect.

What I've done, I've taken this thing in equilibrium, and I've added one of the ions.

And it changes the solubility.

Remember, the solubility is what we care about.

That's what the solubility constant helps us determine.

And I've literally just changed it by orders of magnitude, because now the amount that this-- so this thing precipitates like crazy.

I added in a little bit of some other salt that had chlorine ions.

And all of a sudden, silver chloride precipitates out.

So here's a video of this.

This is kind of cool.

So watch.

This is an equilibrium container.

And what I'm adding is salt. And look at that.

What is that?

That's silver chloride precipitate, right, because the thing has to reach equilibrium.

And so in order to do so, what you do is you suppress the solubility.

Literally, by adding chlorine, the common ion effect, you suppress the solubility of the silver chloride.

It's pretty cool stuff.

That is the common ion effect.

Now, why does this matter?

Let me give you one-- no, I'll get to why this matters in a second.

One more question.

Now without even doing the math, check this one out.

This is cool.

Without even doing the math, I can take barium sulfate, and from the common ion effect-- so barium sulfate is going to go.

So what does that one look like?

Well, it's going to go to Ba^{2+} plus SO_4^{2-} , both in solution, in solution.

Dissolved ions, this is a solid.

Now, here's the question.

Which of the following will be required in the least amount to dissolve the same amount of BaSO_4 ?

I don't need any constants or math.

I can use this same principle, because if I had some 0.1 molar or 0.01 moles per liter of either of these, what are they doing?

They're serving up ions, right?

In the one case, you're serving up-- so in case A, you're going to give me-- what do I have there?

H_2SO_4 , so let's do A first.

H_2SO_4 is going to give me a whole bunch of H^+ but also SO_4^{2-} .

And in case B, you have BaCl_2 -- OK, so you have BaCl_2 .

And if I put that in, I've got this in water, right?

But if this dissolves in water a little bit, then it's going to serve up some Ba^{2+} plus 2Cl^- .

Either way, now you know what's going to happen, because if I'm trying to dissolve something and I add any of the ions that it's dissolving into from some external source, it's going to drive it this way.

It's going to precipitate.

So the answer has to be just water.

Stick with pure water in this case.

Otherwise you're going to have more trouble dissolving, just like we just showed.

You're going to have more trouble dissolving, not less.

That's the common ion effect.

OK, so now why does this matter?

Why does this matter?

We go back to the pteropod.

And by the way, I didn't have this link, and I should have when I showed you the-- and this is your goody bag, et cetera.

There's some really nice articles here that you can find related to these experiments and other things about ocean acidity in case you're interested.

But see, what I did was that this was my why this matters on Monday.

And I wanted to tell you about the goody bag and about how things dissolve, because Monday was about dissolving and finding a saturation point.

Now we can get to the next place, which is, why does that matter for the pteropod's shell?

What is the chemistry that matters there?

OK, so I made the ocean a little more acidic.

Why does that matter?

But you see, now we're armed with the knowledge we need to answer that question.

Now we're armed with it.

All we need to do, as always with everything in life, is look at the chemistry.

That's it.

Say that at the Thanksgiving table.

You'll be very popular.

We said CO_2 plus H_2O .

This goes-- oh, find some equilibrium to H_2CO_3 .

Now, this is called carbonic acid.

This is called carbonic acid.

There's the pteropod up there.

And there's the reaction that's really relevant that you'll see from the ones we're about to write down.

Why?

Because the thing is that, what happens to carbonic acid?

Well, carbonic acid also goes through a dissolution reaction.

So this is CO_2 dissolving in water.

So carbonic acid goes like this.

OK, it goes into HCO_3^- plus H^+ .

Now, here's the thing.

OK, what is the shell made of?

Well, the shell that's dissolving, the core material is calcium carbonate.

So that's CaCO_3 .

That's the shell.

And the shell also has an equilibrium reaction that happens.

The shell of a sea creature is in dynamic equilibrium with the ocean.

And so it's going like this.

It's going to, well, OK, Ca^{2+} and CO_3^{2-} .

OK, but the thing is, it has an equilibrium constant.

All these have equilibrium constants.

So for example, for this one, the K_{sp} , the solubility product constant-- because this is a solid.

This is the solid shell, solid, going to ions in aqueous solution.

So the K_{sp} for that is somewhere around 5 times 10 to the minus ninth, 5 times 10 to the minus ninth.

Yeah, but here's the thing.

We just went through this.

If I consume one of these or change the concentration of one of these and not the other-- we just did this.

If I could change the concentration, if I consume any of these, then I might drive the reaction-- consume or produce.

If I change any of these independently, I'm going to drive the reaction, because that's how we keep to our K . That's what Le Chatelier's principle tells us.

And so what ends up happening is you've got the extra H^+ ions.

So these are ions in solution.

Where did they come from?

They came from the CO₂ giving us carbonic acid, which then gave us H⁺.

Those are what I'm talking about.

Well, they react with the CO₃²⁻.

This lowers the CO₃²⁻ concentration near the shell-- [STUDENT SNEEZES]

--right?

Gesundheit.

And if I lower this, because I've taken some of this now and I've reacted it, so now I've got less of it.

And because of what we just saw, if I've got less of this, you're going to drive this way, which is going to dissolve more of the shell.

That's why this works.

Well, that's why this happens.

Works sounds like a positive thing.

So this lowers the concentration of CO₃²⁻, and that drives more dissolution.

That is what's happening.

And we now can understand it in terms of the concepts that we've just learned.

Historically, I mentioned 50 million years.

Actually, by some accounts, it's 300 million.

It depends on which studies you read.

But for at least 50 and maybe as much as 300 million years, the ocean has had a pH of 8.18.

Now, where's the-- oh.

So let's just say last, oh, 50 to 300 millionish years, the ocean pH was 8.18.

And today, it's 8.07.

And the prediction is that in 2100, it will be 7.8.

Now, as we will see in a little bit-- and you say, what's pH?

And many of you probably already know, but I will tell you what it is in a little bit.

But because this is a logarithmic thing, this is a lot, right?

Today the ocean is 25% more acidic than it's been in 300 millionish years.

And in 2100, it will be 126% more acidic.

That's why they use 7.8 in the experiments of the pteropods.

OK, right.

Why does this happen?

Why is this molecule an acid?

We're talking about acidification of oceans.

Why is this an acid in the first place?

What does it mean to be an acid?

And that is the next topic.

That is what I want to talk about next.

What is an acid?

And an acid is something that is very specific.

It has a very definite meaning to it.

And I want to talk about that today, and then we'll continue after the break.

So this molecule, carbonic acid, is called an acid because of that proton.

And you could feel it.

You could feel it in this whole thing.

The proton is the thing that caused the problem, right?

An acid, this is an acid because of the proton.

So an acid, it's a dissolution reaction, dissolution.

It's what we've been talking about.

It's a dissolution reaction that gives an H^+ .

Now if I think about it as a generic case, generic acid A, then the reaction looks something like this.

I'm going to call out the acid and the proton that it gives separately.

OK, so this is HA.

So we do this because it highlights that it's an acid, H^+ in solution-- and these can all be in solution-- plus A^- in solution.

All right, H^+ , A^- .

Right, so I've taken a proton off.

You see it's right there.

That's HA, HA, where A is HCO_3^- and H is H. And then what I've done is I've transferred the proton into solution.

That's an acid.

Now the thing is, you will see a lot of places and people writing H^+ .

And we will do that too, because so many textbooks use H^+ .

But what I want you to know is that H^+ is never a thing.

We just write it that way.

H^+ is not stable in water.

So the proper-- ah!

The proper way to write this general reaction would be HA plus H₂O is going to go to-- and now I'm going to write the full, so the full goes both ways-- is going to go to H₃O⁺ plus A⁻.

[STUDENT SNEEZES]

Gesundheit.

This is what happens, because H⁺ is not stable in water.

But you will see H⁺ written all over the place.

When you see H⁺, and if it's in water, know that it's H₃O⁺.

That is what is stable.

OK, H⁺ in water is not.

Now, there's something here that's important in terms of terminology.

And that is that these are-- notice that these are related.

The HA and the A⁻ seem very related.

And the H₂O and the H₃O⁺ seem very related.

Well, they are, because a proton is the only difference between them.

And so those are called conjugate pairs.

Conjugate, related, conjugate pairs.

OK, those are called conjugate pairs.

Now, back in the day, there was a lot of interest.

People knew about these kind of liquids, mostly liquids, for a long time.

And there was a lot of experimentation before this.

Let's see, well, there's these general classes of materials.

We're going to call them A and B. Well, A seems to always taste kind of sour.

B is bitter.

A reacts with carbonates to make CO₂.

B reacts with fats and soap to make soaps.

One reacts with metals.

The other doesn't.

And by the way, if we mix them together, we seem to always get salt and water.

This had been going on for centuries, this investigation into properties of acids and bases.

But who was it that came along?

Who was it that came along?

Svante!

Is there anything you can't do, Svante?

I ask you again.

In the late 1800s, it was Svante to the rescue who first proposed that what is happening here is a dissolution reaction.

That is what an acid or a base is.

It is that on the one hand, you're given an H⁺, which is what I just said.

But on the other hand, for a base you're given a OH⁻.

And it was Svante Arrhenius who first put that down and first conceived of acids and ba-- are we snapping?

Are we quiet snapping for Svante?

I love it.

I've just recently learned what that means, which makes me happy.

Quiet snapping.

It's instead of clapping.

But you don't make noise, but you still want to give props.

I got it.

I'm there.

I'm right there.

Now, OK.

OK, now here's the thing.

So we'll talk about it.

So Arrhenius says bases give OH ions in solution, and oh, acids give H⁺.

But we know that that's H₃O⁺.

OK, we know that.

Yeah, but the thing is there's disillusion reactions.

These things lead to really small concentrations.

And so Soren Sorensen.

That is a beer bottle.

By the way, his research was funded by a beer company, and it was all about the taste of beer.

He called it looking at proteins, but it was about beer.

And what he realized, he said, I've been doing all these dissolution reactions of ions.

I think Arrhenius was right.

Ions are really important, and we're trying to measure these things.

But I'm sick of all these zeros.

I don't like writing 0.0001 or whatever all the time.

And oh, look at that, 1.3 times 10 to the minus fifth.

That's not efficient.

And so what Sorensen did is he talked about how you've got the power.

Now, what do I mean by power?

So many things, right?

But if I look at H^+ , OK, H_3O^+ , and I say it-- let's suppose the concentration is 0.00, oh boy, 00001, that's 10 to the minus 7.

But Sorensen wasn't happy with that.

You simplified it.

Wait, two, four-- no, it's not.

OK, maybe it is now.

He said, well, 10 to the minus seventh is still four characters I got to write, four.

So he went to the math department at his place and he said, what else can be-- I want to talk about the power of these ions.

And they're like, power, well, that's the logarithm.

And so if you take the log, it's the power of hydrogen. Power of hydrogen is the pH.

And that's simply minus the log.

They didn't want negative numbers either.

Oh, but you can have negative pH.

You can.

Why?

Because this is the definition of pH.

Or you could do pOH, if you're going with Arrhenius.

So the power of hydrogen is the pH.

The pOH would be minus log of the concentration of OH⁻.

That's the pOH.

But pH is the one that we most often see to describe whether something is acidic or basic.

And so now you see, well, OK, well, this is a pH.

Now this is easy, right?

It's a pH of 7.

And if the concentration were 0.1, it would be a pH of 1, minus log, the power of hydrogen, the log, literally the log.

OK, now if you look at scales of pH, and it's a really fun thing to do, then what you find is that usually, almost always actually, you'll see the scale ranging from 0 to 14.

That's what Sorensen originally decided.

I'm going to talk about these ions in solution.

I don't want to write it out every time.

I'm just going to have a simple scale.

We'll take the minus log, and there you go.

Yeah, but you can see that if the concentration is more than one here, then you could have negative pHs.

There's nothing special about zero here.

It just so happens that most things, certainly that he was playing with at the time, were in this scale.

And here is a few things that we know.

There it is, right?

OK, don't play with that one, but play with lemon juice, vinegar.

By the way, vinegar-- OK, wine is around 3 and 1/2.

Vinegar is around 2.

The word vinegar is [NON-ENGLISH SPEECH]..

It's eager wine.

That's it, right?

But it's because of the acidity, the acidity.

Well, it's more than that.

But it's the acidity is changing by orders of magnitude, power.

And then you can go back and you see, but OK, there's coffee.

I like that a lot.

And blood.

By the way, by the way, talk about ocean acidity, blood has a very narrow range of pH.

And if you change the pH of your blood by more than 0.2, it is very likely to lead to death.

So just think about that.

That's what we're doing to the ocean.

OK, anyway, that was an aside.

Seawater, oh, pH 8.07 currently and dropping.

OK, baking soda, ammonia.

So this is very powerful, no pun intended actually.

Now, these are dissociation reactions.

These are dissolution reactions.

I'm taking an acid, and I'm dissociating it in water.

And so you have the equilibrium constant for that dissociation, right?

And so if you go back to that, that's still going to matter here.

And that's going to be important in thinking about what an acid is.

Because, again, why is this it?

Well, because there's these protons.

Well, how many of them are there?

It depends on how it dissolved.

Where did it find its equilibrium?

And so for some of them, like I just wrote, we just talked about HCl plus H₂O.

That went to H₃O⁺ plus Cl⁻.

OK, now the thing is, if I had 0.1 moles per liter of HCl, then it's going to lead to a concentration of H₃O⁺ of around 0.1 moles per liter, because it's nearly full, nearly full dissociation.

What do I mean by that?

Well, I mean that-- yeah, OK, so you'll often see for an acid like that, one arrow.

Well, we know that in reality, there's another arrow there.

But see, here the acid dissociation constant is huge.

So now we have a thing for the equilibrium of an acid, which forms by dissociating the acid into its ions.

And this is-- well, in this case, it's 10 to the sixth.

It's enormous, which means that the equilibrium lies very, very, very, very far over.

And so oftentimes you'll see if it's going to strongly dissociate it, like sodium chloride did, then you'll

sometimes see it written as just one arrow.

But there's a little bit going back as well.

That is another equilibrium constant.

And we will talk about K_a a little later and probably pick up on it next week.

OK, so near full dissociation.

Now, the other thing that can happen that's important is water can do both.

Let's see, where should I go?

I'll go back over here.

So water, see that picture there?

They got it wrong, but it's OK.

They got it right, because we all get along.

If you want to put H^+ , fine.

But you know that it's actually H_3O^+ .

OK, but you can put H^+ , no problem.

OK, now there it is.

And you can see that what water can do is it can do both.

And that turns out to be extremely important.

So if I take water and I mix it with water, ah ha, then you can get a combination of these ions, oh, plus a whole bunch of water.

So this would be like dissociation.

Now, the word for this, water can turn into either.

It can be basic or acidic.

It can deliver OH^- or H_3O^+ .

That's called autoionization, and it also has a special word.

Water is called amphoteric, amphoteric.

Now, that means that the same thing can be both, both acid or a base.

It can act as either.

So if you take pure water-- so let's see.

If I take pure water and [GROANS] and I look at the concentrations of these, so H_3O^+ , it's all happening from the water.

There's nothing in it.

I didn't add anything to it.

It's just pure water, H_3O^+ .

So that means that if that reaction happens, it's going to generate the same number.

So the H_3O^+ would equal the concentration of the OH^- .

And in pure water, that is a value that is equal to 1 times 10 to the minus seventh moles per liter at 25C, at 25C.

And so now we have yet another.

That's neutral.

You can understand that that's neutral, because you have the same-- if an acid is H^+ or H_3O^+ in solution and a base is OH^- , and I've made the same number, then the acidity and basicity are neutralizing one other.

So that's neutral.

And in fact, you can see that that pH is 7.

Oh, and we can even write the equilibrium constant for water, which is going to be those guys, H_3O^+ times OH^- , which is equal to 10 to the negative 14th, because that's the dissociation reaction for

water.

Water plus water goes to those ions in solution.

We ignore the water, and then we're left with this same thing we've been doing, 10 to the minus 14th.

Now, that takes us-- those are neutral.

And so speaking of neutral, that's where we're going to go next.

And I think what I want to do now is-- so we're going to go next into neutralizing things, so like if you add an acid to a base or if you add a base to an acid.

And we already said on Monday, you get salt and water.

But see, to understand that, you need a broader definition.

Svante is amazing, but Svante missed something.

And so we need a different definition of acids and bases that's more general.

And that's what we're going to start with on Monday.

But wait, because this is a great place for me to work on my arm.

I got to hit-- OK, we're going to go all the way up there and all the way up there and right there and right there and, well, there and there and there.

[YELLING]

Oh, that was excitement.

Let's go back there and there.

Well, OK, and there.

Oh, that was the same direction there.

And there and there in the middle and there and there.

Oh, that was there and there and there and there.

You guys are seeing I need a stronger arm here.

[GRUNTS] I can't really get too far.

And [GRUNTS].

Oh, that side is all-- and there and there and there.

And one more, I'm going deep.

Have a great Thanksgiving, everyone.

[APPLAUSE]

See you guys on Monday.