MITOCW | MIT3_091F18_lec30_300k

Thank you for that.

Happy Monday after Thanksgiving.

I hope you guys had a good break.

We are back in action here on acids and bases, second lecture on acid and bases.

Before we get into that, there's a celebration happening on Friday.

You know that's why we're all feeling like a little extra goodness right now.

Because we know there's a third midterm coming up Friday in this class, right?

And that's nothing but excitement.

Here's a concept map that I made for exam 3.

This helps you sort of see the kinds of problems, right, that you could get, the topics that we've covered, and how it's all flowed and connected together, all right?

So here are the quizzes.

There's the optional one.

Oh, you could still do it.

So please, if you want to do the 8.5 and get that graded-- and then if it's higher than your lowest quiz, we'll swap it in for that.

And you still get to drop one-- please do that and hand it in by tomorrow's recitation.

There's that one.

There's quiz-- there's the other two quizzes.

And here are the psets, and the goody bags.

All of this is the kinds of problems-- this has the practice problems and the topics that we want you to know about for Friday.

Oh, and also, on Wednesday, instead of introducing a new topic that you won't be tested on, I'd prefer just to talk about these topics.

So that's what I'm going to do on Wednesday.

And we'll do some additional problems and kind of review.

OK, but we still got to get through the rest of acids and bases.

Now this is where we left off.

This way of thinking about acids and bases is where we left off on Wednesday, all right?

And this is the Arrhenius-- oh, Svante.

Oh, Svante-- he was the first-- remember, [INAUDIBLE] over thousands of years would say, this is bitter.

Put it in that category.

This is sour.

Mmm, right?

Acid-- no, base.

Acid.

And there are ways that people categorized acids and bases for a long time.

But nobody really started to try to understand them at the atomic scale until Svante came along and said, look, it's all about the H+ and the OH-, right?

So acids-- and this is what we talked about on Wednesday.

Acids are donating H+ ions, protons, into solution, into water.

And bases are donating OH-, adding more of those types of ions into the water.

And that's what makes them have the properties they have, OK?

Now, so if you have-- so let's see.

So H-- I'm going to just write this one on the board.

Because we're going to talk about these two reactions-- plus H2O-- and remember, the water is just a kind of spectator there.

It's a liquid.

It's everywhere.

It's the thing you're dissolving in.

It's the solvent.

That's the water.

So you don't include it in these-- if it's just the liquid that you're dissolving something, it's not what you include in things like equilibrium constant expressions.

OK, so that's going to be-- oh, but these guys, these are dissolved in the liquid, all right?

These are the ions-- plus CL-.

OK, now in this particular case, you can write the equilibrium constant, right?

And that's the acid.

It's also [INAUDIBLE] the acid dissociation constant.

Remember, we had the solubility product constant.

And we also can write the Ka, right?

So Ka for this would be-- following our normal K equilibrium constant procedure, it would look something like this, right?

Now here's the thing-- for this particular reaction of-- for this particular dissociation of HCl, of this acid-- and remember, it's an acid because it's giving us H+.

Oh, H+ doesn't stay H+.

Remember, that's also what we talked about Wednesday.

H+ is actually H3O+.

But the H+ is not stable in pure water.

But sometimes-- actually, very often-- you'll see it-- in textbooks as well-- you'll see it just written as H+.

But we know that it's the same thing.

H3O-- if it's in water, it's H3O+, right, OK, so don't let-- but this is something like 10 to the sixth.

That's huge.

That equilibrium is so far over.

That equilibrium is so far over given this enormous equilibrium constant, that we often will say that this reaction has gone to completion.

Now we know that it's not infinite.

So it hasn't-- technically there's still a little arrow going that way, all right?

But-- and so, but since it's so much almost to completion, you'll often also see it written with just one arrow, all right?

So I wanted to make that clear.

But we know that it's an equilibrium.

This has found an equilibrium.

So there's still a little bit happening the other way, but just so little, all right?

So you'll often see that going to completion.

OK, good-- getting us back into the mood.

Now for the base, we can have NaOH-- let's do that one-- plus H2O goes to Na+ in solution plus OH-.

And this is the Arrhenius acid-base definition-- gesundheit.

And you could write a similar expression for the base, Na+, OH-.

And that's going to go over NaOH.

Now this is also very large.

Now these are both strong.

And we're going to come back to this at the end of the lecture.

I want to talk about what makes an acid or a base strong or weak.

These are both strong.

And just as a little preview to where we're getting, that's because they dissociate so much.

So this also has a very high Kb-- Kb, because we're talking about a base-- Kb, Ka, Kb, right?

Oh, but some people don't like these large exponents, these large or small 10 to the minus 5th, 10 to the positive fifth.

And so we take the log.

You can take the log of these.

And then you'll get, like, a pKb, right, or pKb, which is the power of that.

So log of Kb would be low, right, would be low.

If this is going to be high, then minus the log of the power of that is going to be low.

So you can talk about pKb, Kb, Ka, pKa.

It's all just thinking about these equilibrium constants.

And we'll come back to this in a second.

If I had a question for you-- if I said, OK, how would you use this-- if I said that I've got point-- I'm going to give you a problem now-- 0.05 moles per liter of NaOH, and I asked you, what's the pH, well, you can do it now, all right?

Why?

Because I've just told you it's a almost fully dissociated base.

So the equilibrium constant for that dissociation of that base is very, very high.

It means it lies very far over to the right.

So basically, what does that mean?

If I've got 0.05 moles per liter starting out, it's mostly dissociated.

[SNEEZE]

Gesundheit.

So I've got 0.05 moles per liter of Na+ and OH-, all right?

OK, but the OH- is what matters.

So full dissociation-- oh, we're going to do this because I want to be precise here-- put that in quotes.

It's almost full-- means that I have 0.05 moles per liter of OH--- by the way, and Na+.

But Na+ is a spectator in terms of the pH.

It's not going to change the pH, right, the Na+?

Oh, but the OH will, all right?

And so now we go back to these Ks.

We know-- right, we know-- that the equilibrium constant of pure water is equal to the concentration of H+ in solution, which is H3O+, times OH-, and that that's equal to 10 to the minus 14th, all right?

That's another thing that we talked about on Wednesday.

So if that's true, if the-- which it is at 25 C-- the equilibrium constant for water, which is-- remember, water can be either one.

It's this times this.

And we know what this is.

Now you know what that is, all right?

And if you know what H+-- if the concentration of H+ is known, then the pH-- power of-- the p is the minus lower-- the pH of that is something you can calculate.

And so the pH winds up, here, being 12.7, all rigth?

And that's the kind of thing that you can have fun with.

You dissolve some base, and you can measure a pH, or a pOH, if you wanted-- all stuff we talked about last week.

OK, but now we move on.

And now we're going to introduce the next definition of acids and bases by first talking about what happens when you mix them together, all right?

These are not mixing together yet.

They're just-- a base dissolves in water.

An acid dissolves in water.

But what if I had an acid and a base, and I mix them together.

OK, well, so if I did that, then I would have HCl plus NaOH.

Now I'm just going to-- I'm not going to include the water.

It's there on the left, and it's going to be there on the right.

I'm just going to write all these ions down that happen, right?

So you're going to get-- oh, you're going to get H+-- oh, we know it's H3O+-- plus Cl- plus Na+.

OK, these are all dissolved in water.

I'm emphasizing that with the little aqueous subscripts.

And you can see what's going to happen now, right?

So if I have these there, well, look at this.

The Na+ and the Cl-, boy, do-- they see each other, and they're like, well, maybe we could form some salt. Maybe we could form some salt.

And then the OH and the H, and those form water.

And so you can get H_2O plus NaCl.

And so what happens is what's written there, is that when you mix an acid with a base, you will get water and some salt. Now remember, salt is a general term, all right, for these cations and anions, right, coming together in a solid.

So OK, now here's the thing though, all right?

So this was the acid and I'm writing this explicitly.

Because you'll see-- and this was the base.

Because what is neutralization?

Well, you think about it in terms of pH.

I had something with maybe a low pH, or maybe a high pH, and I want it to become neutral.

Neutral would be-- if I'm in pure water, K_w is 10 to the minus 14th.

Neutral would be-- it's the same acid and base ratio.

So these are each 10 to the minus seventh, pH 7, neutral, all right?

There's no excess of protons or OH^-, all right?

So if there's no excess, it's neutral.

So neutralization is basically the act of making it neutral.

So I had something that gave a lot of H^+.

And now I got something that gives me a lot of OH^-.

If I have the same amount of them, all right-- so they're equivalent-- then you can make it neutral.

That's neutralization, OK?

But see, the thing is I can take HCl, and I can mix Windex, or ammonia, with it and neutralize it.

And back then, they were like, but that doesn't make any sense.

But first of all, because we don't have Windex yet.

But second of all, because Windex doesn't have any OH^-.

Gesundheit.

And so that got them really thinking, all right?

So if I take-- so let's take NH_3.

So if I take NH_3-- dah.

If I take NH_3 and I go like this, NH_3 plus HCl-- now this is not-- I'm not writing the liquid dissociation of the thing.

I'm mixing it, acid and something else that I know neutralizes it.

And in fact, what ends up happening is you get-- gesundheit.

You get this.

And so what we know is that this is neutralizing.

It's taking-- what's happening in this reaction?

What's actually happening?

What did the-- what did that base do effectively to neutralize HCl?

What it did is it took H^+ from the solution.

That's what it did, all right?

So NH_3-- in this reaction, NH_3 takes.

It takes H^+ from the solution.

That's exactly-- HCl was like, I want to make it acidic-- H^+, H^+, H^+.

And NH_3 is like, no, give me the H^+.

You cannot make this.

I'm neutralizing you.

I'm taking your H+ back.

If it takes that, then that's neutralizing.

That's going to neutralize the HCl as well.

But see, we have a problem.

Because this is not an OH-donating molecule.

But it still neutralizes.

And so we need a broader definition.

And that came from-- that came from Bronsted and Lowry.

And so this is another way of thinking, a more general way of thinking of acids and bases-- more general and more correct, because the Arrhenius definition does not cover classes of molecules that could be a base.

And so in this way, in the Bronsted-Lowry definition-- in the Bronsted-Lowry definition-- it's all about the proton.

So let me write that down here.

So in Bronstein and Lowry, their definition is that acid-- and he did it by looking at acid-base neutralization reactions like I just described.

Acid-base neutralization-- I'm not going to write it, because there's no room-- reactions is all about the proton transfer.

These are proton transfer reactions.

This is how they saw acids and bases, as exchanging a proton.

Taking it or giving it, that's what it's about, all right?

And so in our sort of general-- you know, in the general HA lingo that we introduced, you would have HA plus B. And that's going to go to BH+ plus A-.

That's an acid-base reaction.

And so you can see that these are-- this is a proton donor.

This is a proton donor.

And this is an acceptor.

It's a donor, H+ donor, H+ acceptor.

That defines them, according to Bronsted and Lowry, as acids and bases.

OK.

OK, well we can also-- OK, so now let's see.

Let's now go back.

That's a broader definition.

Let's now go back to the ammonia, OK?

And what I want to do now is, instead of mixing it with an acid to think about whether it trades a proton, I want to just dissolve it in water.

So now we know it's a base, OK?

So what happens?

So if I dissolve NH3 in H2O-- I want to point something out here, that if I do this, and I get NH4+ plus OH-, that's the-- now I'm taking a base.

So there we had-- there's the acid.

Here's the base.

That's the Arrhenius base.

But now I'm taking a Bronsted-Lowry base, right?

It took an H+.

Well, that means that this is a base.

But look, this-- OK, this gave a proton.

So this must be an acid according to Bronsted-- according to Bronsted and Lowry.

And similarly, over here, we know that this is going to give one.

And I'm going to write that in a second.

And this would love to accept one.

So that's a base, right?

And remember, we have a name for this.

In the same reaction, we have a name for this.

This would be the conjugate, all right?

Those are conjugates.

That's why I'm writing those, all right?

So there's a conjugate base-acid, conjugate acid-base.

Now we have a Bronsted-Lowry understanding of what conjugate acid-base means.

It's just a proton transfer, all right?

It's just a proton transfer.

Now notice I could take this-- so I've taken this base, and I've put it into a solution of water.

I could take this acid and do the same thing.

So if I do that, I've got NH4+.

And I'm going to put that in water now, OK?

So I've taken this out.

And now I'm putting that acid into water.

And if I do that, look at what happens.

I get NH_3 plus H_3O^+.

Notice what happened.

Well, I've switched acid.

That's my acid.

But this took the proton, so it must be a base, right?

And over here I've got the base.

And this must be an acid.

Proton transfer, that's what's happening here-- conjugates.

I took the base, and I got an acid.

So this tells us that water-- and it tells us what we said on Wednesday.

Water can be both.

Notice water is an acid in the one hand and a base on the other.

That makes it amphoteric.

So H_2O can be both.

It can take or give protons.

And so it is amphoteric.

Just for fun, if you take the conjugate-- you know, if you take any acid-base pair, and they're conjugates, you can also write out the Ka's and Kb's, all right?

And so if I take a conjugate acid and base, well, you know that you could write Ka, all right?

OK, so let's write the acid dissociation again.

So for that one, here it is.

Ka is NH-- is H3O+.

OK, I'm going to take the bottom one, H3O+, times NH3, right, concentrations, divided by NH4+.

But if I took the base dissociation constant from the conjugate-- I'm taking the conjugates now, right-- then I would get that it's-- OK, it's OH- and NH4+ divided by NH3.

Like I'm saying, just for fun, you can now look at this.

You go, Ka times Kb-- Ka times Kb.

That's reassuring that Ka times Kb is equal to H3O+ times OH-.

And I mean concentrations, which is Kw.

That is fun, for the Bronsted-Lowry acid-base conjugate.

NH4+, NH3, right?

OK.

OK, that's Bronsted-Lowry Lowry acids and bases.

Here's a little recap.

Here's a little recap.

I'm not going to read through this in detail, but I just wanted to give you-- OK, what have we done so far?

We've talked about amphoteric.

We've talked about conjugate acids and bases.

These are now written in the Bronsted-Lowry form, right?

They accept or give a proton.

We've talked about the conjugate acid-base pair, what I just drew an example of right there, neutralization reactions, and salt. OK, good.

These are some concepts that we have covered that I would like you to know about.

Now but there's a few more concepts that we have to come back to-- a few more concepts that we have to come back to.

And it has to do with what I started with.

In the beginning, I said, you know, these Ks are huge.

So the equilibrium-- if an equilibrium constant lies-- is very large, then you know you're mostly making product, right?

And so it's essentially to completion.

And I also mentioned that those are strong, strong acids, strong bases-- strong acids, strong bases.

Why is an acid or base strong or weak?

That's what I want to talk about next, all right?

And actually, there are-- there really aren't that many strong acids or strong bases.

In fact, they're mostly here.

These are really, pretty much most of them.

Minor [INAUDIBLE] I don't want to go into.

But these are really-- if you think about the general category of strong acids, strong bases, this is it.

Why?

Why is that?

Well, it has to do with what we've been talking about, all right?

So those are strong acids.

Those are strong bases.

It has to do with dissociation.

It has to do with dissociation.

So if I take HCl and I put it in water, it almost completely dissociates.

That's why it's a strong acid.

Now there's confusion around this.

And I want to make that-- I want to make it very clear.

There is a big distinction I want to make.

Because if something is a weak acid, meaning it's not one of these, what that means is it doesn't dissociate very much.

That's what it means.

It doesn't have anything to do with concentration, all right?

So let's take a weak acid.

Let's take vinegar, all right?

So if I take a weak acid-- we see vinegar as CH_3COOH.

You know what's nice about writing them out this way?

You say, why didn't you put the H in there, and make it CH_4, and CO_2, or put the two Cs together?

Oftentimes you'll see molecules written out this way, because this tells us a little bit of a hint about connectivity.

And for acids and bases, it's actually useful.

Because you see the H there?

That's the one that's going to trade.

That's the Bronsted-Lowry transfer H, all right?

So often, for acids and bases, you'll see that written out in this way.

Makes it easier to think about.

And if I mix this into water, then what do I get?

Well, I'm going to get H3O-- I'm going to write this down here-- H3O+-- that's an acid, right-- plus CH3COO-.

There's the H that came off-- proton transfer.

So why is it weak?

Well, it's weak because here, the acid equilibrium constant is something like 10 to the minus fifth.

So even if I have a lot, you do even if I had like 1 mole per liter of this stuff in solution, it wouldn't-- most of it wouldn't dissolve.

In fact, much less than a percent of vinegar actually gives me H3O+, all right?

But you think, oh, just add more vinegar.

No.

No, because we're talking about concentrations, right?

It's the concentration in that water.

I'm locked into that by this.

This is an equilibrium constant.

And it's talking about the concentration that I get in the water.

I'm stuck.

I can't be-- that's how much you dissociate, right?

That's how much you associate as a concentration.

So it's more acidic than water for sure, all right?

It's 100 times more acidic than water, all right, 10 to the minus fifth.

But it's never going to be a strong acid, because I don't give very much of the protons.

So this is the key that I want to make clear.

This is the key, OK?

It's very important.

It gets its own board.

The strength of an acid, OK?

So acid concentration is not equal to acid strength.

More, and more, and more vinegar-- no.

It doesn't matter, because it's about association.

See, this is about-- this is a function of the solubility.

So how much can I get in there?

We talked about all that a lot already.

That's a different thing, solubility, the Ksp.

How much of this can I dissolve?

Great.

I can dissolve a lot of vinegar in water.

And it doesn't crash out, right?

That's wonderful.

Doesn't make it a strong acid.

This is about dissociation, all right?

And so if you think about what makes an acid a strong acid-- or for that-- or a weak acid, or what the relative strengths of acids are-- and the same holds for bases, just the other way around-- it has to do with how easy it is to take that H off or add it on, right?

That's what it has to do with.

And so if you look at-- for example, if you look at just a little series, all right-- if you look at a series there, even the strong acids have differences.

They're all strong.

They go mostly to completion.

But if you think about it in terms of what we just talked about, the acid strength is going to go roughly often opposite trend as bond strength.

That's of the proton transfer of the H, all right?

There is an H-- HCl, HBr, HI.

Which will be the strongest acid of those three, right?

Well, we now know.

This must be the strongest, right?

Opposite-- because this is going to be easier to get the hydrogens off in solution.

And you can think about the weak acids in the same exact way.

These are the strong acids.

This holds for all-- you know, it's a proton transfer.

It's all about the strength of the proton.

Neutralization is a proton transfer.

Whether you make an acid strong or how strong it is depends on how many of those protons go into solution.

And so you can look up charts like this.

And I love looking at this stuff, because it makes you think about the chemistry, all right?

So here you go.

You've got, OK, acid, different acids, molecular formula, all right, structural formula, conjugate-- notice they're putting that H there in blue.

Oh, that's so helpful.

And you just look at the blue H, and you think about it as dissociating, just like Bronsten and Lowry did.

You think about that as coming off.

And then you think about the conjugate base, all right, that's formed on the other side.

And then you go over here, and you think, oh, this tells me how much it's coming up, how often.

This tells me the equilibrium of putting this into solution.

You say, this is-- which one of these is going to be weaker or stronger, all right?

Well, and what does that mean about the bond strength itself?

What does that mean about how easy it is to dissociate?

We now know.

And we can think about that, and understand that.

These are not-- these may look similar, these Hs attached to carbon or oxygen. What's the difference?

Big difference when you put them in water.

A lot more or a lot less might come off.

Changing the acid strength-- the more that comes off, the more the dissociation is to the right, all right?

Oh, and the lower, right, the higher this number, because the equilibrium moves over to the right, all right?

And the less that comes off, the lower this equilibrium constant is, Ka, right-- the lower it is.

The higher the pK-- we don't like these small, tiny numbers.

You take a minus log, and you talk about pKs.

That's fine, all right?

OK, good.

This is the kind of fundamental understanding of acids and bases that I want you guys to have.

And so if we go back to neutralization, now you say, well, OK, this makes sense, right?

We had a strong acid and a strong base.

If you had them equivalent, you'd get a pH of 7.

Now you know, if you had a strong acid, weak base, weak acid, strong base, it's going to-- this is fairly self-explanatory, less than 7 or greater than 7.

And now you know also, if it's a weak acid and a weak base, and you're neutralizing these together, it would depend on the Ka and Kb.

Because it depends on how many of those ions you put in solution or can take out.

And that's given by the Ka and Kb, OK?

Good.

Good.

So we could answer questions-- we're now armed to answer questions.

And like I said, I want to answer-- we're going to do some problems on Wednesday.

We'll talk about the topics that are going to be on exam 3.

But we can answer questions like this.

Here we are talking about acids and bases.

What if I mix a bunch of them together, and they've got different Ka's?

So think about this, what if I mix all of these together?

Can I get my pH, right?

So here's a question.

You have 1 mole per liter of HCN with that acid dissociation constant and 5 moles per liter of HNO2 with that one.

What's-- OK, you write that dissociation reactions and calculate the pH of the solution.

That's a good question, right?

So let's see.

Now how would I answer it?

Well, oh, Bronsted-Lowry, don't want to erase you, but I have to.

I have no choice.

So let's answer that question.

The way that we think about this, it's telling us what to do.

Write the dissociation reactions for each species.

OK, so I'm going to write it down.

So I have HCN.

HCN, right, is going to go to H+ plus CN-.

Now Ka, for this, is equal to H+-- oh boy, did I drop my subscripts-- times CN- divided by HCN concentrations.

But we're back to solubility product days, right?

This is an equilibrium.

And so it's going both ways.

But you know from the coefficients here-- all right, you know from those coefficients that if I dissociate x amount of this, then that's how much of this and this that I make, all right?

So now I know, right, that-- well, OK-- Ka times-- if I just bring this over there-- HCN is equal to x squared, where x-- let's let x equal concentration of H+, which equals the concentration of CN-.

Been there, done that, right, from the solubility product problems that we've done.

And so you can get, from this, that the concentration of H+, which is also H3O+, is equal to 2.5 times 10 to the minus 5 moles per liter.

Good.

OK.

But you can keep going.

And I'm not going to do all the math.

But you could go through the same exercise with HNO2.

So for HNO2, you get that the concentration of these ions that you get from that dissociation is around 2 times 10 to the minus 3 moles per liter.

And by the way, maybe I got less than 10 to the minus seventh out of these things-- maybe.

Maybe if I did, then, if I did, then I'd also need to take into account pure water, right?

So from pure water, we know, if you just had water by itself, then you'd have, the concentration here is 10 to the minus seventh moles per liter.

But now you just look at this, and you say, OK, I got my answer.

Because everything's a lot lower than this one thing.

So this is going to dominate.

I mix it all together.

But you see, it doesn't matter.

The other things aren't able to disassociate nearly as many H+ ions-- they're just not able to-- as HNO2.

So this dominates.

And you can just calculate the pH from this.

So you're going to get, like, two, three digits down some effect.

But it's OK, right?

From this, you can get the pH.

And the pH is equal to 1.35.

The total contribution of H+ ions in that solution is roughly given by this contribution here, OK?

Might not have been, but it was in that case.

OK, how about this one?

So here's another-- OK, so we could take another example.

Let's stay over here.

OK, so in this example, I'm barely giving you anything here, it seems like.

But I give you enough.

I've got a 0.2 moles per liter solution of a weak acid.

And it's 9.4% dissociated.

That's the key.

That's the key.

Because remember, this is-- now we've learned the Bronsted-Lowry definition.

It's all about H+ in or coming out.

That's what-- there it is.

Now I erase it though.

It's all about the H+.

So if I know 9.4% dissociated, I know something about what kind of acid that is, right?

So if I have, now, HA plus H2O, and that goes to H3O+-- I've switched back-- plus A-, then how much dissociated?

Well, 0.2 moles per liter times 0.094, right?

So 0.094, that's the percent-- is that right?

Yes, 9.4%-- times 0.2 moles per liter.

And that equals 0.019-ish, right?

That's how much HA-- HA-- dissociated.

But if I know that, and I know my reaction, and I got the stoichiometric coefficients from the reaction, then I also know how much H3O+ and A- formed.

If that much dissociated, then that much formed, right?

Because they're-- right?

That's the reaction.

Those are the coefficients.

And so I can get it all.

And I can calculate Ka, which equals 2 times 10 to the minus third.

I didn't give the pH.

All I gave you was how much of it dissociated.

I didn't even tell you what acid it was.

But that's OK, because we now know that this is what counts.

This dissociation, how many of these Hs per liter, how many of those go into solution is what matters.

OK, now there's another thing that I want to talk about.

And we'll do some more-- you know, a little bit more, like I said, problems on Wednesday.

There's just one more thing I want you to know about related to acids and bases, and that is that you can have more than one.

All these examples, if you go back to this list, those all had one hydrogen that was ever going to leave, let's say.

I mean, OK, ever-- maybe if you went up to the middle of the sun or something, you'd get those.

But this is-- you know, in normal conditions, these are the ones that are going to leave, and no more.

But you can have molecules like phosphoric acid where-- look at all those protons.

So here's acetic acid.

That's the one that's willing to leave.

Those are not.

They're on the carbon.

Way too hard to get them off.

This one-- OK, in solution, sure, to a certain extent.

In phosphoric acid, I've got three hydrogens.

And so that's called a multiprotic acid.

Why?

Because protic-- proton-- how many protons, right?

It could be monoprotic, like the ones we've talked about, or it could be multiprotic, like phosphoric acid.

And so if you think about this the way we now think about acids and bases, we think about an acid as a dissociation, right, with some dissociation equilibrium constant.

And so if you think about phosphoric acid, H3PO4-- oh, let's put the water in, H2O, liquid.

And that goes to H2PO4- plus H3O+.

Well the Ka for this one is-- so if I put the Ka's in here, this one is 7.5 times 10 to the minus 3.

But see, what happens is, now, this one has two more hydrogens that could come off.

So you could write those dissociation reactions as well.

So H2PO4-, if you put that in solution, which it is, then that could give you HPO4 2-, all right?

H2, H-- it's a proton transfer.

Minus, 2-, all right-- plus H3O+.

And there's one more, plus one more.

This dissociation constant is much, much smaller, 6.2 times 10 to the minus 8.

And the third one is 4.2 times 10 to the minus 13th.

So they're smaller, but they happen.

And that's important.

And it can be very important.

It depends on the molecule.

But in this case, it's not just the parent, that acid there that's losing a proton.

But then, this one can also lose a proton.

It can also be an acid.

It just has an equilibrium that's not as far over.

So it's not going to give as many H+ into solution, OK?

It's important to know about.

Then finally, the last thing I want you to know, you will not be tested on this, but it's important.

Because here's where we are.

Here's where we are.

Arrhenius said, right before Thanksgiving, it's all about the H+ and the OH-.

That got us really far, got us through Thanksgiving.

But then today, we're like, no, it's got to be more general.

It's just proton transfer.

No, don't tell me about OH-.

It's just proton transfer.

And so then comes Lewis.

Oh, we know Lewis.

We've already talked a lot about Lewis.

Lewis said, why are you guys all obsessed about protons?

And he said it in just that way.

An acid is much more general even than that.

An acid is any species, anything that except a pair of electrons.

Remember, Lewis was all about those electrons and the octets.

And a base is any species that can donate a pair of electrons.

So Lewis correctly took a much more general view and said, no, acids and bases are not even about proton transfer.

They're about electron transfer.

And again, I'm not going to test you on Lewis, acids, and bases, but I wanted you to have that full picture.

That means, by the way, aluminum 3+-- take an aluminum atom, make it 3+.

It wants electrons.

It's an acid.

It's an acid, all right?

Doesn't have a prot-- it wants electrons though.

That makes it an acid according to Lewis.

So that concludes acids and bases.

And I'll see you guys on Wednesday.