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OK. So we have-- you can show the answer-- 57%, 1 and 3. So this is where we were at the end of last lecture, so if you didn't get this written down in your notes, you want to look at these definitions. So here in this, we have NH_3 with its lone pair, and it is acting as the Lewis base, and BF_3 is acting as the Lewis acid. So a Lewis base donates its lone pair electrons, and the acid accepts them.

So what was the other definition? We had the Bronsted-Lowry. Does someone want to tell me what that definition was? Can we have a volunteer tell us the Bronsted-Lowry definition? Pull out your notes. There you go.

AUDIENCE: So a Bronsted-Lowry--

CATHERINE DRENNAN: I'm not sure it's on. It is on? OK.

DRENNAN:

AUDIENCE: Bronsted-Lowry base is something that accepts H plus or H_3O plus.

CATHERINE DRENNAN: Yeah. So the base accepts the hydrogen ion or the proton, and the acid donates. So is this

DRENNAN: definition now incompatible, or are they completely opposite, or could they be worked together? Does anyone have an opinion on that?

So how many people think that these could be consistent definitions in some way? OK. Oh, quite a few people. One could make that argument. So if you know now that you can make the argument, does someone want to try to make it? We have a bag.

All right. Our mic went away, but can we get up over there? So where did our mic runner go? I should have explained they were needed twice. All right. All right, who said they might try this?

AUDIENCE: All Bronsted-Lowry bases are Lewis bases, but not all Lewis bases are Bronsted-Lowry bases because you can act as a proxy with no hydrogen ion if it connects directly. But if a hydrogen ion bonds, an example H_3O , to the lone pair of oxygen, then it's the same thing as a Lewis base.

**CATHERINE
DRENNAN:**

OK. so that's right. The Lewis acid-Lewis base definition works if you don't have a hydrogen ion involved or a proton involved, but the other still applies. So remember that when you're accepting H plus, you're taking it without an electron, so the base is taking H plus, so it's donating its electrons to form the bond. And the acid, when it gives up H plus, it's keeping all of the electrons. It's accepting the electrons, so when the acid gives up H plus, it's accepting, and when the base takes H plus, it's using its electrons.

So the Lewis definitions are lone-pair or electron-centric definitions, whereas, Bronsted-Lowry is thinking about the proton, but they both work together. So that's a good review of the definitions we talked about last time, and now I'll try to get my mic back on.

All right. So today we're going to continue talking about acid base, and we're going to continue talking about acid base quite a bit. So we have a number more lectures on acid base. So we're going to start today with the relationship between pK_w , pH, and pOH, we're going to talk about the strengths of acids and bases, and then we're going to start doing equilibrium acid-base problems. And we're going to do a bunch of those. And students who are doing this and learning it for the first time are telling me there is an infinite number of different kinds of problems. It's really not true. There are only five different types, and today hopefully, we'll get through two of them.

So first we need to think about water because water is a really important solvent when we're talking about acid base equilibrium. And we mentioned last time that water can act as an acid and a base. Does anyone remember what that's called, something that can act as an acid and a base?

AUDIENCE:

Amphoteric.

**CATHERINE
DRENNAN:**

Amphoteric. So here we can think about two waters going to H_3O^+ plus and OH^- . We could also write it out instead of just saying two waters, one water that's acting as the acid and another water is acting as a base. So the water that's acting as an acid gives up its hydrogen ion, or proton to the base. The base accepts it forming a conjugate acid, and this water that lost H plus becomes OH^- . So water plus water, H_2O plus H_2O , forms an acid and a base.

So we can ask the question, then how much water is in a glass of water, or how much H_2O is actually in this water container that I have here? So when I'm asking that question, how much H_2O in a container of water and if I know the ΔG^0 of that process, I'm asking about what is

the relationship between our products here, our ionized acid and our base over here, the ionized products of water, compared to water at equilibrium. Because this container, the water, is in equilibrium.

So if I'm asking about the ratio of products to react as an equilibrium, what am I asking about? Just yell it out.

AUDIENCE: Equilibrium.

CATHERINE DRENNAN: I'm asking about the equilibrium constant. Exactly. I'm asking what is K. What is the equilibrium constant. So let's just briefly review the relationship. We're given here a ΔG^0 , and we're asking about K. What is the relationship, how do you calculate K if you know ΔG^0 ?

And we have some equations that will be given to you. Sorry. Not quite yet. There's going to be a clicker question coming. We have some equations here. So we have ΔG^0 equals minus RT natural log of K, which we can rearrange to solve for K. And let's just have a little reminder of what these terms are.

So we'll put this up here. T is temperature, and we're going to be talking about room temperature. This isn't room temperature now. And in fact, almost all acid-based problems are going to be at room temperature. There might be a few differences, but we have a lot at room temperature. We have a constant, R, and so 8.314 joules per Kelvin per mole and ΔG . So we were already told the value for ΔG here.

Now, thinking about this room temperature, this constant, and this value for ΔG , do you expect a large or small value for K if you have this ΔG^0 value? And that's our clicker question. All right. Let's just do 10 more seconds.

Yep. So K, you would expect to be a small value less than 1. And you could think about that mathematically from these expressions, or you could think about it in terms of the value of ΔG . So if we solve this and we put in our value for ΔG^0 , which is a positive value-- so the forward direction of the reaction is non-spontaneous.

So we put in our positive ΔG^0 and our other values here, and this is, again, joules per Kelvin per mole. And we have to make sure that we convert kilojoules to joules and our units will cancel out. And if we do that, we get this value for natural log of K and this value for K of 1.0×10^{-14} at room temperature. And you will find that you end up

memorizing this value if you do enough acid-base problems. In fact, you might have it memorized already.

So this is, in fact, a small value. And that indicates that only a small proportion of water molecules are ionized. About 1 molecule in every 200 million are ionized, so it's a very, very small value. So there is a lot of H₂O in a glass of water. So in here it's mostly H₂O. Very few molecules are ionized in that.

Now, part of the reason why you're likely to have this value of K memorized is because it has a special name. This K is called K_w for water. So that's easy to remember. And K_w, the equilibrium constant for water, is equal to the hydronium ion concentration times the hydroxide ion concentration.

And since this is an equilibrium constant, if you're at constant temperature at room temperature, this will always equal 1.0 times 10 to the minus 14th. This product is always going to be this at room temperature. And that is why it's going to turn out to be important because you're going to use that information in solving a lot of problems.

All right. So if this is an equilibrium constant, we want to think again about expressions for equilibrium and just make a note that we did not include water squared on the bottom of this expression. So K_w equals hydronium ion concentration times hydroxide ion concentration. You don't put it over water.

And the reason for this is because this is a pretty pure solvent, and you don't include pure solvents in your equilibrium expression. You also don't include solids. So we talked about that when we were talking about solubility. So you have to remember that solvents like water is nearly pure, doesn't go in. Pure other liquids, pure other solids are not included in your equilibrium expression.

All right. So we now know about water. We know about K_w. Let's do a few more definitions and then think about how all of these terms are related to each other. So a definition that many people already know, probably if you've taken any kind of science before, you're probably aware that pH equals minus log of the hydronium ion concentration of H₃O⁺. If you don't know it, it will be on your equation sheet, so it doesn't really matter, but there it is. And there's also pOH. pOH equals minus log of the hydroxide ion concentration, OH⁻ minus.

Now, let's think about the relationships of pK_w, pH, and pOH. So relationships. So I just told

you and maybe you've already memorized that K_w is going to be equal to the hydronium ion concentration times the hydroxide ion concentration, and that's going to be equal to 1.0×10^{-14} at room temperature. So now, if we take this expression and we take the log of all of these terms and multiply by a negative value, then we will get this expression.

So minus the log of K_w is pK_w . And minus log of the hydronium ion concentration is what? pH minus log of the hydroxide ion concentration is what?

AUDIENCE: pOH .

CATHERINE DRENNAN: pOH . So pK_w equals pH plus pOH , and that is equal to 14.00 at room temperature. And this is a very useful expression, the fact that pH and pOH are equal to 14. If you know one of these, then the other one by doing a simple subtraction. So you will find yourself using this expression quite often in the problem sets.

And so I'll say that one, the problem set that that's due on Friday has just a couple of acid-based problems at the end. The next problem set will be 100% acid-base problems, so there's a lot to look forward to.

So let's talk about the strengths of acids and bases. So if you have a pH of pure water, it should be equal to minus log times 1.0×10^{-7} , or the pH equals 7, which is a neutral pH . So if we have a scale here, pH minus 1 to pH 14, 7 is a neutral value.

If we're talking about things that are acidic, the pH of an acidic solution is less than 7, and so this is the acidic part over here. And the pH of a basic solution is greater than 7, so that is down in this part. The EPA defines waste as corrosive if the pH is lower than 3, so that would be corrosive, or higher than 12.5, also corrosive. So living creatures like to be more in the neutral range. When you get to things that are very low pH or very high pH , that is less good.

So let's now check out a few pH 's, and I'd like to have five TAs come down to help me with this. And I'll just show you what we're going to do. So we'll have some volunteers. We have pH paper. The pH paper has an indicator on. You put the paper in the solution and then try to see which thing matches best. So this is a very quick and dirty way to estimate pH . There's pH meters and indicator dyes that work better, but we'll have five people measure five things.

We have ammonia, a cleaner. We talked about the importance of cleaning bathrooms last lecture. We have soda. We have vinegar, which is often used in salad dressings. Before I came here I just went to a random sink and got some MIT water to measure.

And then this, this is special. This is a prescription medicine. So this is a solution of iron 2 sulfate. And this is prescribed if have an iron deficiency. An iron deficiency is pretty bad, and a lot of kids have iron deficiencies. It's really common in infants.

So my daughter had an iron deficiency, and this was her medicine. In fact, this box is completely full of this medicine. So kids cannot swallow pills very easily. How many of you actually cannot swallow a pill? A lot of adults are pill challenged. How many are happy swallowing pills? Let's do a more positive thing. OK.

So swallowing pills is not everyone's favorite. When you're a kid, it's really impossible to tell someone to put something in their throat and go back and try to get it down. They choke, spit up, it's awful. So kids have to take medicine in solution. So this medicine is 300 milligrams in 5 milliliters. It's very concentrated, so my daughter had a pretty severe iron deficiency.

So when she tried taking it, and I tasted it, it was horrific beyond belief. And after she had taken it a couple of times, she started to have sores on her tongue and in her mouth. And I as a chemist said, let's measure the pH. So you will measure the pH of this prescription medicine for a four-year-old and tell me whether the EPA would prove it or not.

Anyway, so here are the TAs, and they're going to come around. Raise your hand if you're willing to measure a pH. TAs, just grab a solution. I should have had you pour it, but OK, we're doing it. And we'll go around, and we'll measure, and tell me what you find, and I'll write it on the board. Yeah, OK.

All right. So we have pH. All right. Do we have an answer? Did we get one? What do you have?

AUDIENCE: Ammonia.

CATHERINE DRENNAN: Ammonia is 12. So almost corrosive, but not quite. You can still have clean, but if you want an excuse not to clean, you're like, sorry, too close. Too Corrosive Do we have another one?

AUDIENCE: MIT water is 7.

CATHERINE DRENNAN: MIT water is 7?

AUDIENCE: MIT water is 7.

CATHERINE DRENNAN: Awesome. So often, actually, water that you get from a tap is not 7 because often there are ions that are dissolved in it. But a little bit around 7, a little bit of ions never hurt anyone. We have lots of ions in our body. Soda?

AUDIENCE: 3 and 1/2.

CATHERINE DRENNAN: Soda at 3 and 1/2. So soda is definitely getting into our corrosive range.

AUDIENCE: Vinegar at 2.

CATHERINE DRENNAN: And we have vinegar at 2. And that is corrosive. People often use pure vinegar to clean out a coffee pot because it is pretty corrosive. It really does a good job of cleaning it. When you're having it in salad dressing, you don't want to have it pure vinegar. That would not be good to drink. And what about medicine?

AUDIENCE: Medicine has a pH of 2.

CATHERINE DRENNAN: 2. So we also have in the corrosive category the medicine that my four-year-old daughter was asked to take. So when she started getting these sores in her mouth from taking her medicine, I brought some pH paper home and measured the pH and discovered that what she was taking was corrosive. So I went to the doctor and said, can you prescribe something different. This medicine is corrosive. And the doctor looked at me and it's like the pH, the pH is like 2 or lower. And she looked at me, and she's like, what do you mean pH?

This is why I spend many lectures on pH in this class and acids and bases. Some of you will be doctors, and some of you will not encourage parents to buy this much corrosive thing to feed their child when they're iron deficient. So this is what I'm talking about. And for those of you who are not going to be doctors, many of you will probably have children with an iron deficiency. It's super common. So you will know that you need to measure the pH of the solution before giving it to your child.

So what did I do? So what I did was I bought adult pills, same stuff, iron sulfate, and I made sure that this told me the number of milligrams that were in there. I looked at what was prescribed. So I got the pills with the correct number of milligrams, and then I crushed them, and that stuff also tastes really nasty, but a little tip for the future, there is one and only one

that I discovered, taste that can cover the taste of iron 2 sulfate, and that is Nutella. So don't give your child corrosive medicine. Give them crushed pills in Nutella. pH is important. pH is important.

So let's talk about the strengths of acids. And you might want to reconsider huge soda consumption too. I don't really know, but that's something you might want to think about.

Strength of acids. So here we have some acetic acid, CH_3COOH aqueous, dissolved in our solvent, which is liquid water. And this is an acid, so it will give up a hydrogen ion or a proton to the water, which acts as a base and forms our hydronium ions, H_3O^+ plus, and also the conjugate base of the acid. So it's the acid missing H plus, CH_3COO^- minus.

So if we're talking about the strength of the acid, what we really want to know is how much of that acid forms ions, how much of it ionizes. That is going to determine how strong it is because the amount that ionizes is equal to the amount of the hydronium ions you get, and pH equals minus the concentration of hydronium ions. So the pH depends on the extent to which it ionizes.

And again, we're talking about at equilibrium, so we're talking about a equilibrium constant, and this one also gets a fancy name. In acids and bases, all the equilibrium constants get their own little names, and this is the acid ionization constant, K_a . So it's the equilibrium constant for an acid, K_a . So that's at least very easy to remember.

So now based on what you know about equilibrium constants, why don't you tell me the answer to this? All right 10 more seconds.

All right. So this one is wrong. It's always products over reactants for an equilibrium expression. Water is a solvent and is nearly pure, so it should not be there. This should be in the expression because it's aqueous. If it were solid, it wouldn't be included, but it's aqueous, so its concentration is going to change, so it is in our expression. See answer to 3 for that one. Enough information, and it's not correct.

So we can write the expression for K_a . So that's equal to the products, our hydronium ion concentration, our conjugate base over are conjugate acid. And so we can look at what the value is for this expression, and we can look it up. There's lots and lots of tables of these in your book. And it is 1.76×10^{-5} or 10^{-3} . I don't have my glasses, whatever it says there. I think it's 5 at room temperature, which is small.

So when K_a is small, it means that very little of it is ionizing. And so only a small number of our acetic acid molecules are donating their proton when they're dissolved in water. So that's the definition of a weak acid. When it has a very small K_a , it's not ionizing very much.

Now, in doing these units, a lot of people get hung up on the names of the different acids and one of the first steps in solving a problem is to write the equilibrium expression out, but people get hung up with that. So don't get hung up with that. You can use generic expressions for acids and water. If you don't want to write out the whole name of the acid, you can always just say HA, aqueous, plus the solvent, water, goes to hydronium ions plus A minus, the conjugate of the weak acid.

So this is an acid in water, and the acid is HA here. And an acid expression should be forming hydronium ions. If it's acidic, you should have acidic pH, and so you need to have some H_3O^+ plus around. You can also write this expression as BH plus plus water. BH plus can give up the H plus and become B and also generate hydronium ions.

So here are the acid is BHA. And often when you're looking at these kinds of problems, a weak acid will be HA, but sometimes you have a problem involving the conjugate acid of a weak base, and that's often expressed as BH plus. So both of these expressions are valid for an acid in water.

So now let's think about strong acids versus weaker acids, and here's our definition in this class. A strong acid has a K_a greater than 1, a lot of strong acids have a K_a really, really, really greater than 1, and the acid ionizes almost completely.

So if you say it's a strong acid in water, whatever concentration of that acid you put in is going to be equal to the concentration of hydronium ions. It basically goes straight. You form lots and lots and lots of products at equilibrium. A weak acid has a K_a of less than 1, and a weak acid does not form many ionized species in solution. Not very much H_3O^+ plus is formed.

So what about pKa? pKa's are really important in organic chemistry, in biology, in a lot of areas. And every year the organic chemistry faculty in 512 talk about pKa's in organic, and the students there say, no, we never learned about that. They're like, you took freshman chem, right? It's like a GIR or you pass the test. You should know about pKa's, and they're like, no.

So they contact me, and I'm like they did not take 5.111 if they do not know what a pKa is. There are other courses that sometimes people decide to take that are not 5.111, but in

5.111, you know about pKa's. So what I want you to do is later in life when you're in a class and pKa's come up and everyone else in the class is like, I don't know. You're like I took 5.111. I can answer that. That makes me very happy.

So pKa's. so pKa is minus the log of the Ka. That's easy to remember. We already talked about the relationship of Ka with strong acids or weak acids. The lower the value of the Ka the higher the value of the pKa. That's just out of that expression. So the higher the pKa means what about the acid? Think about what you know about Ka to answer this question.

All right. 10 more seconds. All right. So that is correct. So we have a weaker acid.

So if we have a low Ka, that means that it is a weak acid, and then the higher value of Ka. So the higher the Ka, that's going to mean the lower the value or the weaker of the acid. So just remember the relationship from the equation, and think about it, and you can think about the value of Ka, products over reactants. You might have a lot of products. That means a lot of ionization and a stronger acid and fewer products, that means it's weaker.

So let's look at some tables, and I just put the values that are with the arrows in your notes. I didn't put the entire table. The tables are in the book, and this is page one of the table, so there's a lot of values here.

So this has the acid. It tells you what the HA term is, what the A minus, so what the weak acid is, what the conjugate base is. It gives you the Ka and the pKa. And so you can see the relationship between Ka and pKa. So up here, HI is at the top of this table. So that probably means it's the strongest or the weakest. Which do you think? Is this the strongest or the weakest? You just yell it out.

AUDIENCE: Strongest.

CATHERINE DRENNAN: It is the strongest. Right. So it has Ka value that is much, much greater than 1. And it has a very, very, very negative tiny, tiny pKa value. So this is a super strong acid. In fact, do not use that. There's no reason really you would ever want to be using it. It is not a good thing to play with.

All right. So HCl is used more often in pH-ing things. Is this a strong or weak acid?

AUDIENCE: Strong.

CATHERINE It's also strong. It's K or K_a is also greater than 1, but not as strong as this. We have 10 to the seventh, and that one, gosh, I should bring my glasses, 10 to the 11th. That's super strong.
DRENNAN: That's still pretty strong.

All right. So let's look down here now. Tell me, is this a strong or weak acid?

AUDIENCE: Weak.

CATHERINE That is weak acid. The K_a is less than 1 here, and I think that's a minus 2. And the pK_a value,
DRENNAN: now we're out of the negative number, so it's on the bigger side. And down here, is this strong or weak?

AUDIENCE: [INAUDIBLE].

CATHERINE Also weak. And it's even weaker than this one. It's like 10 to the minus 4, and pK_a value is
DRENNAN: 3.75. So you see you can look at the K_a values and think about whether it's a strong or weak acid. You can also look at the pK_a values. This is really, really small negative value. This is a bigger value down here.

And this is not been the weakest acid. There's a whole other page of tables for acids, lots and lots of acids.

Less bases. There are a few, and you will see this space quite often, NH_3 . So we have a base in water. The base accepts a hydrogen ion or proton from the water, forming NH_4^+ plus. And then after water loses its hydrogen ion, it forms OH^- minus.

So the equilibrium expression or the K for this base in water problem is called the base ionization constant, or K_b . And we can write that expression here. We have the ammonium ion concentration times the hydroxide ion concentration over the concentration of ammonia. Again, water is not in the expression. It stays pretty much pure throughout this entire thing. It's the solvent. And so this is our expression for K_b .

And if you're writing an expression for K_b , you should always make sure, check yourself that you're putting hydroxide ion concentration in there. If it is a base in water, it should be forming hydroxide ion concentration. If you start writing a K_b and you have hydronium ion concentration in there, you want to stop and rethink what you're doing.

So in this case, we also have a weak base. 1.8×10^{-5} is a small number for

K_b . It's a small equilibrium number, so that means you're not a lot of products. Not a lot ionized when you put this weak base in water. So only a tiny amount of NH_3 ionizes to NH_4^+ plus and OH^- in solution, so this is what they call moderately weak base.

So as we saw before, you can have generic expressions for a base in water. We can write $\text{B} + \text{H}_2\text{O} \rightleftharpoons \text{BH}^+ + \text{OH}^-$. The base accepts a hydrogen ion or proton, and the water loses one forming OH^- . We can also write the expression that often exists from the conjugate of a weak acid, and this is a weak base. So $\text{A}^- + \text{H}_2\text{O} \rightleftharpoons \text{HA} + \text{OH}^-$. So either of these are generic expressions that you can write for a base in water.

So in terms of the definitions, it's the same. You would say a strong base is something that ionizes almost completely to give OH^- . There aren't a lot of examples of strong bases. Most of the ones you'll see in your class are things like sodium hydroxide. Yes, that is hydroxide. That's a strong base, and if you put it in water, you should definitely form a lot of hydroxide. So they're not a whole lot of examples there. So they're not all those tables. It's not like the acids.

But there are some terms that you still need to know. So $\text{p}K_b = -\log K_b$. And again, the larger the value for K_b , the stronger the base. Again, it tells you you have a lot that have ionized. And because of this relationship of this equation, the larger the $\text{p}K_b$ the weaker the base. Because if you have a big number here, you're going to have a small number there.

So now, let's talk about conjugate acids and their bases. This is super important for buffers that, hopefully, we'll get to in class on Friday. So the stronger the acid, the weaker its conjugate base. The stronger the base, the weaker its conjugate acid. And here we have HCl , which we determined was a strong acid a little while ago, giving up its hydrogen ion to water, forming hydronium ion concentration and Cl^- . Cl^- is not a very good conjugate base. It's not a good base.

So if we look at this table when we have a strong acid, its conjugate it is going to be ineffective as a base. You're basically not pushing that equilibrium back at all, the other direction. It's completely ionized, and it stays that way. But if you have a moderately weak acid, you're going to have a very weak base, a very weak acid, you'll have a moderately weak conjugate base, a strong base and you'll have something that's ineffective as an acid. So they are inversely related to each other.

So let's think about why this would be true. And this relationship holds $K_a \times K_b = K_w$.

And we know that K_w is 1.0×10^{-14} at room temperature. And if we put logs and then minus logs by everything, we can derive this expression, which is the pK_a plus the pK_b equals the pK_w equals 14.00.

So you can't have an acid and its conjugate base both be strong. You can't have a base in its conjugate acid both be strong. The pK 's need to add up to 14. So if you have something that's a good acid, then its conjugate base is not going to be particularly good.

The pK_a and the pK_b must add up to 14. If you have something that's a good base, its conjugate acid is not going to be that good. So these are connected to each other.

So we can think about, then, the strong acids and the strong bases again. So we talk about equilibrium, but really this is pushing the equilibrium pretty much to completion. If it's a strong acid in water, however much strong acid you put in is how much hydronium ion concentration you get out. A strong base is going to give you that amount of hydroxide ion concentration.

Strong acids and bases push the equilibrium pretty much completely toward ionization. We have these numbers of 10^{-11} , 10^{-7} . These are hugely over here. We're forming almost ionizing completely.

Now, for weak acids and bases, it's very different. Here the equilibrium you are going back and forth. You have this dynamic equilibrium. The acid in water is forming hydronium ions and an A^- , but a A^- is also a somewhat good, weak base pushing the equilibrium back the other way. So in this case, we have a back and forth between the forward and the reverse.

So if you have very weak, you have moderate, you have moderate, you have very weak, and this is what's important for forming buffers, and buffers are really important.

Let's see if we can get to one of these types of problems because they're five we need to get to in this unit. And there are only five. And people will say salt and water is another, but salt and water just breaks down into weak acid and weak bases. So everyone can learn how to do five types of problems, so this is very doable.

All right. So let's just look at equilibrium of weak acids and we'll save weak bases for last time. So a weak acid is vitamin C, and I brought some vitamin C here. I tried to do a demo once and discovered that vitamin C, it's really coated well, so it does not dissolve until it hits the acid of your stomach, so I could not dissolve it. So I'll just hold this up right here. But if we had a 500

milligram tablet, which I think is what this is and if I had 100 mL of water and if this wasn't coated so there's no way you can dissolve it, then I could calculate the pH.

So how am I going to do this? I'm given the K_a , I'm told the number of milligrams, and I'm told the volume of water. So the first thing you want to do is calculate molarity. So you want to convert grams to moles or milligrams to grams to moles. And then you want to use your volume, and you have 100 mL, but we're going to do molarity. So we want to convert milliliters to liters. And we can calculate 0.0284 molar moles per liter.

Then we can write our expression, and if you want to, you can just write HA here and A minus. You don't have to write out the whole thing if you don't want to.

All right. So here is the expression again. Now, just as we saw with chemical equilibrium, we can write this table talking about the initial molarity. So we have a weak acid we have no hydronium ion concentration. We have no weak conjugate base over here, so the change then is going to be the constant, the molarity that we had, and these are change in molarities, so don't put moles or other things here, molarity here, minus x , that's what you have in equilibrium, plus x plus x .

So we can write our expression now for K_a . We're given the value K_a , and we know how to write an expression. We have our hydronium ion concentration, our conjugate base over here over our conjugate acid. This is equal to x squared over 0.0284 minus x .

Now, if you want to with these problems, you can make an assumption that x is going to be small and then not have to use the quadratic equation, but you do need to check your assumption. But I'll show you how to do it with an assumption, so we're going to assume x is small, and this value, the x , will drop out. So we can rewrite this expression as x squared over just the concentration we started with. And then we can solve for x . And x , in that case, is point 0.00151, but really two significant figures. Our K_a is just 2.

Now, we can check the assumption. And we're going to check the assumption that this is, in fact, small. And if it's less than 5%, then we can use that. So here this is the x we calculated. Here is the amount times 100% gives us 5.3%, which is more than 5. 5 is the magic number for the course, and so then we have to use the quadratic equation.

Now, a lot of you probably have calculators that can work with quadratic equation, so you don't really care, but if you want to check the assumption, you can, and sometimes we'll need to

check it. So this value, this percent here, is sometimes called percent ionization because it's x over the amount you started with. It's the percent that ionized or percent deprotonated. If this is an acid, it's the amount of hydronium, the amount that deprotonated over what you started with. So sometimes you're asked to do this even if you're not checking any assumptions.

All right. So we got to end with if a quicker question. We get here this is really two significant figures. So why don't you tell me how many significant figures are in your answer for pH. All right. 10 more seconds.

So let's take a look at the answer here. So there's the answer. So this had two significant figures, which means you get two significant figures after the decimal point. So maybe we'll have one of these in the clicker competition on Friday, and we'll do our weak bases next time.

So again, we've been studying acid base. And there are only five types of acid-base problems, weak acid in water, which we already did, check, weak base in water, which we're going to do right now, and as soon as we're done doing weak base in water, I'm going to explain to you how salt and water are just weak acid in water or weak base in water problems. So you already know how to do them as soon as they teach you about weak bases.

And then we're going to move on and do buffers. And then next week we're going to do strong acids and strong bases, and then you'll have all five. You already have enough that after today's lecture, you can do, I think, all the problems set 7 except for the last two questions, I believe, or at least a large fraction of problem set 7. I will warn you that acid-base problems take a lot of writing and a lot of time to work.

One problem often has five or six parts to it, and they're all time-consuming part pretty much. So don't leave problem set 7 to the last minute. That would be a mistake on this particular problem set.

Weak bases. Acid-base type problem 2. So in this example, we have ammonia, NH_3 , in water, and we measured the pH of some solutions before and saw they were basic. We had a pH of, I think, 12 last time. And so this conjugate base goes to a conjugate acid. So the base accepts a hydrogen ion or proton from water, becomes NH_4^+ , and the water that lost its hydrogen ion becomes OH^- , and that is a base in water problem because you're forming hydroxide ions.

For a base in water, you're talking about the equilibrium constant for that base in water, which

has a special name K_b , B for base. And it's important to remember the A's and B's on the equilibrium constant because it's always checking your work. Is this a base in water problem. People sometimes try to apply K_a 's when they should be doing K_b 's. So pay attention to basin water. We're thinking about K_b .

So here if we're going to calculate the pH of a solution, we have molarity of 0.15, and room temperature, pretty much everything is at room temperature. So whenever you're asked to calculate the pH of a weak base in water, you want to think about at equilibrium, what's the condition now, what will be the condition, what's the change, and what will be the condition at equilibrium. And so you can make this table. You can write out the expression.

Again, you can forget about water. Water is our solvent. It's not going to be showing up. And then put in our initial molarity, 0.15. And we put zeros in the other categories. And why don't you try the rest? Is it slowing down?

SAM: [INAUDIBLE] answers are really quick.

CATHERINE DRENNAN: All right. Let's just do 10 more seconds. Yup. Excellent. Sam told me everyone responded really quickly. So that is, in fact, what you're going to do. We'll put that up here as well. So we lose some of this minus x . So at equilibrium, we have point 1.5 minus x . And we get plus x here and plus x there.

Now, we can write our K_b for our base in water, its products, the conjugate acid, NH_4^+ , times the concentration of hydroxide ions over NH_3 . Again, water is the solvent. It's not in the equation. And then we can write it out-- this is x . That's x , so we have x^2 . And then the change in the amount of the weak base, when it's in water, is 0.15 minus x .

So at this point, you can either use the quadratic equation, or you can make an assumption that x is going to be small compared to 0.5. We don't want to make the assumption up here. We want to get to this point before we try the assumption, and you always need to check your assumptions to make sure they are correct. But if we use the assumption here just to simplify the math. And so the assumption is that this x is small, so we can drop that x minus x out, and then we can calculate that x is 0.00164. And then we can check the assumption, and that is another clicker question.

All right. 10 more seconds. All right. It is number 1. So here you're checking the assumption and the way that you check the assumption is that you put x that you've calculated by

simplifying it, divide it by the number that you're asking is it smaller than, and then times 100 because our rule is 5%. So if this value is greater than 5% of this value, you need to use the quadratic. If it's less than 5%, then you can go ahead and use your assumption.

And so this is the assumption we're talking about, that this number is under 5%. And for weak acids and bases, it very often is less than 5% because they're weak, so there's not a lot of ionization. And this is often referred to as percentage ionization as well.

All right. So our checking assumption, again, that x 0.00164 divided by 0.15 times 100% is 1%, 1.1, that's less than 5, so the assumption is OK. And then we can calculate pOH because this is a weak base in water, so the x that we calculated is the amount of hydroxide ion in solution. So we calculate pOH, and we get 2.79. And that's two significant figures, everything we had after the decimal because everything that we had was two significant figures. And so this is two significant figures, two significant figures after the decimal point.

And then you need to calculate pH because the problem asked for pH. So this is, again, at room temperature. So we can use 14 minus 2.79 is 11.21. That's our pH.

And don't make the mistake of stopping here forgetting that x is hydroxide instead of hydronium ion and tell me that the pH of my weak base in solution is pH 2. So one of the things that's really good about this unit is that there are some good checks that you've got the right answer at the end. If you're talking about a base in solution, you should have a pH above what value?

AUDIENCE: 7.

CATHERINE DRENNAN: About 7. So it should be on the basic side of neutral. pH 7 is neutral, so it should be above 7. And so if this was your answer for pH, that wouldn't make a lot of sense. So always check at the end.

And it's fine if you get to an exam and you get to the end and you did something wrong and you realize that your answer doesn't make any sense, if you write, this answer doesn't make any sense, but I don't have time to figure out what I gave wrong, you will get points for recognizing that that is not a valid answer. Because it's a weak acid problem and you have a basic pH or vice versa.

So keep that in mind. There's lots of partial credit. Grading acid-base problems on exams is a whole new fun adventure that the TAs have no idea what is going to be happening. But I'll be

right with you there for 10 hours of grading, and we'll have a good time. Write neatly, please.