Amphoteric Metal Hydroxides Page [1 of 2]

We've talked about amphoteric compounds as compounds that can act as both acids and bases. And we talked about it in the context of polyprotic acids, for instance, carbonic acid when it's deprotonated by the first proton becomes bicarbonate, and bicarbonate is a compound that is amphoteric, meaning that it can react with hydroxide to form carbonated water, but it can also react with protons to go back to carbonic acids. So it's in the middle between carbonic acid and carbonate where carbonic acid is a polyprotic acid. Remove one proton, you get the bicarbonate, remove another proton and you get the carbonate.

Well, it turns out that a great number of metal hydroxides are also amphoteric and an example of one is lead(II) hydroxide (Pb(OH)\(_2\)) which reacts with protons to form Pb\(^+\) and water so it dissolves. This is an insoluble salt, but when you add protons, it forms Pb\(^+\) in water and also the same salt reacts with hydroxide to form the complex ion with three hydroxides attached to the lead. I'm going to show you these two equilibria right here. What I have are two beakers and these two beakers contain some lead (II) hydroxide that I've prepared and there's a little bit of water because, of course, lead (II) hydroxide is insoluble so what we have here are suspensions, and to one of them I am going to add some acid and to one of them I'm going to add some base. So this one I'm going to add some acid and we'll see the top equilibrium and then afterwards we'll look at the bottom equilibrium. So again, they're milky and if I add lots of acid—it's going to take more acid than that. So we'll put some more hydrochloric acid in there and you can see that it went clear. We've driven the equilibrium by adding a lot of protons. We've driven the equilibrium from lead hydroxide, which is insoluble to lead cations in water. And now, let's look at the one with base. We'll add some sodium hydroxide and you can see that that thing is going clear. There's a little bit of lead hydroxide still at the bottom that hasn't finished reacting, but I think you will be convinced that it went from a milky white solid suspension to a solution.

So we can calculate the equilibrium constant for this last reaction. In other words, a measure of how acidic lead (II) hydroxide is, how much it reacts with hydroxide ions by calculating the equilibrium constant for that reaction. And how do we do that? Well, we know a number of equilibria involving lead hydroxide. For instance, we know its solubility product. We can look it up and its solubility product, when we look it up, is 2.8 \times 10^{-18}. And we could also look up the formation constant for lead trihydroxide anion, this complex ion, and we find that it's 4 \times 10^{-16}. So now when we add these two equilibria, what we get is the reaction of lead hydroxide plus more hydroxide to form the complex ion and what would we do when we add equilibrium expressions? We multiply the equilibrium constants and so the equilibrium constant product of these two is 0.1. It's not huge, but what it says is if we jam enough hydroxide in the solution, if we pour in a bunch of hydroxide, we can drive this equilibrium to the right hand side by Le Châtelier's principle. Dissolve the lead hydroxide solid and make the complex ion.

Now I said that there were a number of other examples of amphoteric metal hydroxides. Things like Al(OH)\(_3\), Cr(OH)\(_3\), Sn(OH)\(_2\) and Sn(OH)\(_4\). So that's just an example. And you will recall that we talked about bicarbonate as being in the middle in a series of multiple deprotonations and we can also think of these amphoteric compounds. For instance, zinc hydroxide is another amphoteric compound that's in the middle of a series of deprotonations. So for instance, when we first put \(^+\) into water we formed the tetraaquo complex and then we deprotonate this by 2 protons and we get the insoluble dihydroxide salt and then when we add more hydroxide so we're adding hydroxide to get to here and when we add more hydroxide we make the complex ion with 4 hydroxides on every \(^+\). So once again, what we're looking at is a series of deprotonations where the middle species is the one that's amphoteric and that makes reasonable sense.

Now an application of this idea of amphotericism and the fact that the product of adding a bunch of hydroxides as a complex ion is that you can separate \(^+\) from \(^-\) because \(^-\), as I showed on the previous slide, forms a complex ion. The Zn(OH)\(^-\). Whereas magnesium doesn't do that. And so in the absence of the bottom reaction so let's forget about the bottom reaction for right now. If we had a solution of \(^+\) and \(^-\) and we started adding hydroxide, then the first thing that's going to happen is what? Well, if we look at the solubility products for these two and recognizing that the stoichiometry is the same for both, we could calculate the solubility of each one and what we'd find out is that the solubility of the zinc is lower than the solubility of the magnesium and so when we first start adding hydroxide to the solutions of \(^+\) and \(^-\), the first thing that's going to happen is we're going to start precipitating out zinc hydroxide. But then as we add more and more hydroxide, what's ultimately going to happen is that this zinc hydroxide is going to go to the complex ion, the Zn(OH)\(^-\). The magnesium hydroxide, though, can never pick up more than two hydroxides. It'll pick up two and then it'll just form an insoluble salt. And I should have labeled this as a solid and remind you that that's a solid.
So in other words, we have our solution of \( +^+ \) and \( +^+ \) and we start adding hydroxide and the first thing that happens is that we precipitate out zinc hydroxide. But as we add more and more hydroxide, what ultimately happens is the zinc hydroxide gets converted to the, once again, soluble zincate anion and the magnesium hydroxide precipitates out and it just stays precipitated out. So we can separate these two by filtration or centrifugation and what we've affected is the separation of the magnesium cations from the zinc cations.

So still, another example of aqueous equilibria, in this case illustrating the amphotericism of some metal hydroxides, but still utilizing the same source of concepts that we've used up until now that we have equilibrium constants for specifically named reactions like solubility product, like formation constant and we can use them to calculate things that we don't know anything about using the machinery of equilibria.