Looking In-Depth at Redox Reactions

Oxidation-Reduction Reactions

Balancing Redox Reactions by the Oxidation Number Method Page [1 of 2]

We just spent some time learning how to assign oxidation numbers or oxidation states to chemical compounds. Why did we bother doing that? The answer is, because by knowing oxidation states, it helps us in our bookkeeping when we have to balance complex chemical reactions. So let’s put that in practice now. Let’s see how this actually works.

To begin with, let’s remind ourselves what a redox reaction is. Remember that this is a process where one species is transferring electrons to another species. Now, what’s oxidation, what’s reduction? Just a reminder. Oxidation is losing electrons and reduction is gaining electrons. Keep that in the back of your mind. Now, before we go into this problem I want to remind you about the difference between oxidation states or oxidation numbers and the formal charge of something, which we’ve talked about previously. Let’s just use this anion as an example just to highlight the differences. These are both ways of bookkeeping, but they differ in two very important ways. In formal charge, what we’re looking at is kind of an averaging of electron density. So, for instance, the oxygen here will have a formal charge of -1. Remember, that’s the effective nuclear charge of +6 minus the electrons that are not bound, minus half of the electrons that are bound, so it has a formal charge of -1 and the chlorine here has got a formal charge of +1. Again, it’s effective nuclear charge, minus the electrons not bound, minus half the electrons that are being shared.

Compare now that to oxidation states. In this case we’re assigning oxygen as the full 2 minus charge and the chlorine here—each oxygen has a 2 minus charge, the overall charge for the chlorine has an oxidation state of +3. Notice the differences. In this bookkeeping system we’re taking shared electrons and splitting them when we assign charge. In this case we’re taking the shared electrons and giving them entirely to the more electronegative atom. That’s the difference. So 3+ versus 1+, 2- versus 1-. So our focus now has to be on oxidation state. This is a redox reaction. We’re going to focus only on the oxidation states of compounds, not formal charge.

So let’s start with a trivial example, and then we’ll expand to something a bit more robust. Our method is going to be, in this case, using the oxidation state method to balance equations. These are the five steps that we’re going to use. The first thing we’re going to do is assign oxidation states to all the species. We’re then going to balance the redox atoms just to make sure we have correct numbers of atoms on both sides. We’re going to identify the electrons that are transferred based on oxidation states. We’ll then balance the electrons, so this is our bookkeeping that I’m referring to, and then finally we’ll balance all the other atoms so that we have at the end of the day a completely balanced chemical reaction consistent with transfer of electrons from one species to another.

So our example is something that we’ve actually seen in a different context—copper metal reacting with silver plus to generate copper 2+ and silver as a solid. So the first step is to assign oxidation states of each of the species. So remember, elemental copper has an oxidation state of zero. Silver plus, oxidation state of 1, copper 2+, that’s its oxidation state, and again, silver metal is zero. Next, we need to identify the electrons that are transferred. In this case copper is going from copper zero to copper 2+. That’s a change of two electrons. Remember that this is our oxidations step. This is our oxidation piece. Meanwhile, silver is gaining an electron to go to silver metal. That’s the number of electrons transferred in this step. This is our reduction. Now we want to balance electrons. The number of electrons given up has to equal the number gained. So for every two electrons copper transfers, it will take two silvers to be reduced so that we have a balance of electrons. So I multiply this part of the equation by 2 so that I have a transfer again of 2 electrons.

So our final formula then, after I’ve multiplied the silver by 2, is copper plus 2 silver goes to copper 2+ and silver is solid. Good enough. Now, there are a couple of steps that we left out here because this was a very simple equation. Let’s look at a more complex chemical reaction where we get a full blown treatment of this method. Now we have something a bit more meaty here. We’ve got dichromate plus chlorite goes to chromic plus, dichlorohephtoxide. Now, let’s going on here. Our first step is to assign oxidation states to each of these species. Let’s look at what’s going on here. This is more complicated. Remember, each oxygen is -. That means that the total is a 2- charge, the chromium is a 6+. Each chromium is a 6+. In this case, by the same method, we have a 3+ oxidation state for the chloride. If that doesn’t make sense to you, stop now and make sure that you can assign this oxidation state properly. That gives us chromic. That’s straightforward. And, we now have an oxidation state of 7+ per chlorine atom.

Now, we want to balance the redox atoms. This is something we didn’t have to worry about before but we do in this case, because this species has two chromiums in it, whereas this has only one. So I need now to include this 2 because for every one of this species I’m going to get two chromic ions. Likewise, I’m going to need a 2 here because I have two to give me only one. So I’ve balanced my redox atoms.
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Now, I need to identify the electrons that are actually transferred. In this case we're going from—now, watch carefully. This is a place that's so easy to get confused with. I have two chromiums. Each one is a 6+ oxidation state going to a 3+ oxidation state, so we have a total of 6 electrons then, 3 for each chromium that are transferred in that process. Correspondingly, I have two chloride atoms here, each one is a 3+ and it goes to . That's a 7+ oxidation state, but there are two of the chlorine atoms, so we're going through a net change of 4 electrons per chlorine atom, or 8 electrons that are transferred.

Now, notice there are 6 electrons involved in this step, 8 electrons involved in this step. Our next step is to balance the electrons so that the electrons given are the electrons received. Now remember, 6 electrons and 8 electrons. So I need to multiply such that my electron count is the same. So if I multiply this by four, 24 electrons will be transferred here. And I multiply this by 3, it would give me 24 electrons that are involved, in this case, in the oxidation step. So notice again, here's our reduction step and here's our oxidation step. 24 electrons that are added and 24 electrons that are taken away, so everything balances. That means I'm going to need to multiply these guys by 4, and I need to multiply these guys by 3. So if we do that, there's the 4 and this is 4 times 2, and here's the 3, and this was 3 times 2. So now we have a balanced equation except for the fact that we haven't accounted for our oxygens and our hydrogens yet. Now, we don't have any hydrogens to worry about here, but in some cases we will have hydrogen. So we have to balance the rest of the atoms now.

Now, we have a choice. Do we want to balance the rest of the equation in acidic solution or basic solution. Let's look at acidic solution first. Our strategy here for acidic solution is to balance oxygen atoms using water, then to balance the hydrogen atoms using H+. So look at our oxygen atoms here. What we're going to end up needing to do in this case is have 19 water molecules on the right so that it balances, and that will require 38 hydrogens on the left. And so our final equation now doesn't quite fit in our box, but let's look at what's on this side—38 hydrogens, 4 dichromates, 6 chlorites, goes to 8 chromics, 3 and 19 waters. And you can prove to yourself—you should, in fact—that this is completely balanced. All atoms on each side should add up and the charges on each side must balance.

Now, there's one other thing that we can do and that is to balance this in base. If we want to balance this in base I'll show you in the box over here how we would do that. We're going to now instead of adding acid in water we're going to use two hydroxides on the side that needs oxygen, and a water on the side that doesn't. So for every two hydroxides we put on one side we'll put a water on the other side, and that will give us a balanced equation in base, as you can see. Now, alternatively, let's go back to our acid equation for a moment. This was balanced in acid. Imagine if I added 38 hydroxides. This would then become 38 waters over here, and if I put that same 38 hydroxides on this side, then some of my water will cancel on each side and I'll end up with a balanced equation in base. So whichever method works best for you—we're showing you both methods in the box over here—whichever method makes more sense to you, you can do. Makes no difference which way you want to do it, but what's important is that you're able to balance in both acid solution and in base solution.