

Balancing redox reactions in basic solution

So, you're stuck balancing redox reactions, huh? Not to worry – though they're not as much fun as the balancing you've done so far, they're far from impossible. Let's see for ourselves, via the magic of tutorials.

A quick recap:

For those of you who don't know much about redox reactions, let's review some terms that we'll be using:

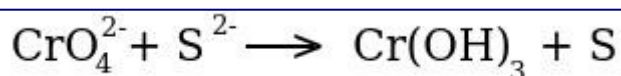
- **oxidation:** When something loses electrons, giving it a more positive oxidation state.
- **reduction:** When something gains electrons, giving it a more negative oxidation state.
- **oxidation state:** The positive or negative charge we assign to an element. This may seem weird, given that in many cases the element doesn't have any visible charge in the formula, but it helps us to keep track of where the electrons are. Some general rules: Pure elements have an oxidation state of 0, main block elements frequently have oxidation states corresponding to their distance from the nearest noble gas, and the oxidation states of transition metals depends on what reaction is going on.
- **redox reaction:** A chemical reaction in which one thing gains electrons (is reduced) and another loses electrons (is oxidized). As you might guess, you need to have both processes taking place simultaneously, as it's impossible for something to gain electrons if nothing has lost electrons in the first place.

Now that we've gone through all of that, let's get rolling!

An example that will show you what to do:

As the header above suggests, here's an example to show you the method you should use to balance redox reactions in a basic solution:

Question: Balance this redox reaction:



Follow these easy steps, and you'll be a pro:

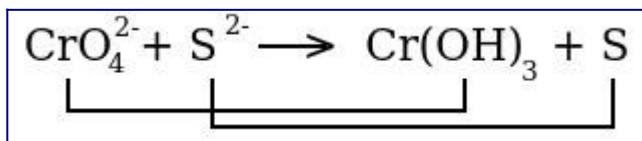
Step 1: Determine which elements have changed their oxidation states

Chromium changes from a +6 oxidation state to a +3 oxidation state on the right, meaning that it has been reduced by gaining three electrons.

Sulfur changes from a -2 oxidation state to a 0 oxidation state, meaning it has been oxidized by losing two electrons.

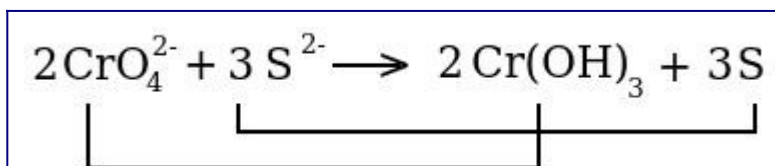
Step 2: Make some lines on the equation that link the elements that have been oxidized and reduced.

In this case, we connect the compounds that contain chromium to each other, and the compounds that contain sulfur to each other:



Step 3: Change the coefficients in front of each pair of compounds to that the number of electrons gained is equal to the number of electrons lost.

In our example, we can see that chromium gains a total of 3 electrons, and that sulfur loses a total of 2 electrons. In order to make these electrons equal each other, we put the coefficient “2” in front of the pair containing chromium, and a “3” in front of the pair containing sulfur:



Step 4: Figure out the amount of charge on both sides of the equation.

In this equation, we have two charged compounds on the reagent side:

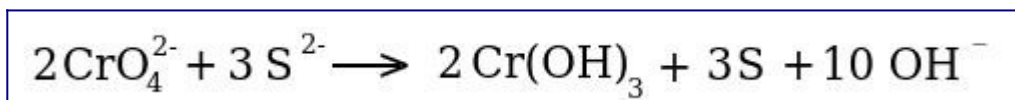
- Each chromate ion has a -2 charge. Because there are two of them, the total is -4.
- The sulfide ion has a -2 charge. Because there are three of them, the total is -6.

The total on the reagent side, then is -10.

On the product side, there's nothing with any charge. As a result, it's a nice round 0.

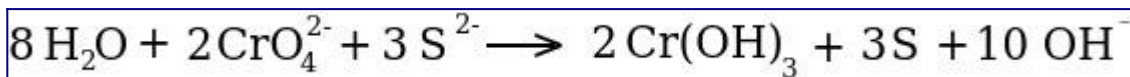
Step 5: Add hydroxide ions until the charge on each side of the equation is the same.

Given that there is a -10 charge on the reagent side and no charge at all on the product side, let's put ten hydroxide ions on the left so that both side have a total of -10.



Step 6: Add water to balance out the elements.

It's not hard to see that in this equation there's hydrogen on the product side and no hydrogen on the reagent side. Given that there are 16 hydrogen atoms on the right, let's put 8 water molecules on the left so we have 16 hydrogen atoms over there, too:



Step 7: Double-check to see if it's right

Basically, just check to make sure that you have the same number of atoms of each element on the left and right, and that the total charges of everything on the left is the same as that on the right. In this case, we're in good shape, so that's our answer!