Chem1311Ch4Ep5Transcript

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Hello and welcome to the fifth episode of reactions in aqueous solutions.

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Previously in reactions in aqueous solutions.

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We learned that there are three types of double displacement reaction.

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Precipitation, neutralization and neutralization with gas evolution.

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The precipitation reaction is characterized by the formation of a solid product.

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Called a precipitate.

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This is evident in both the molecular and the net ionic equations.

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The neutralization reaction is characterized by having an acid and a base as reactants.

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And by having water as one of its products.

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The difference between a neutralization reaction with a strong acid and with a weak acid is not apparent in the molecular equation.

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But it is quite obvious in the net Ionic equation.

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The neutralization reaction with gas evolution is characterized as an acid base neutralization.

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In which the base is a carbonate or bicarbonate.

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And the products include equal amounts of water and carbon dioxide.

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You must know how and when to use the solubility table.

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And which seven acids are strong.

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Index cards could be very helpful in the study of this chapter.

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In today's episode.

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We will learn important terms related to redox reactions.

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We will learn to assign oxidation numbers.

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And we will learn to identify oxidizing agents reducing agents.

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And federal agents.

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Wait, no, strike that last one.

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Redox reactions are chemical reactions in which electrons are transferred from one element to another.

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Oxidation is defined as the loss of electrons.

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And reduction is defined as the gain of electrons.

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Of course, they both must occur simultaneously, so the reaction is called oxidation-reduction.

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Or redox for short.

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These photographs illustrate two different redox reactions.

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In the image on the left, we see that zinc metal replaces the copper in the copper sulfate solution and as a result, solid copper forms.

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During this process, zinc metal.

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Becomes zinc ion.

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And loses 2 electrons in the process.

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Simultaneously, copper ion forms copper metal.

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And gains 2 electrons.

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In this example, zinc was oxidized because it lost electrons.

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This makes zinc the reducing agent because the reducing agent.

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Is the reactant that oxidizes.

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That also means that Zinc’s oxidation number increased.

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These three items always go together.

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If one is true, they are all true.

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In this same example.

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Copper (II) was reduced because it gained electrons.

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That makes copper (II) sulfate the oxidizing agent because it was reduced.

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And Copper (II)'s oxidation number decreased.

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What is this oxidation number that I speak of?

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I'll show you.

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Zinc’s oxidation number went from zero to positive 2.

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Copper (II)’s oxidation number went from positive two to zero in this reaction.

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There are rules for assigning oxidation numbers to the elements in a compound or in a reaction like we just did.

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You don't have to memorize them, but you must know how to apply them.

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These oxidation rules are part of your reference materials.

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The oxidation number is the charge an atom would have if electrons were completely transferred.

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This is a lie, of course, but it's a useful lie.

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Free elements are assigned an oxidation number of zero that is all uncombined elements.

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This includes elements which form molecules like hydrogen, oxygen and phosphorus.

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The charge of a monatomic ion is the oxidation number of that ion.

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Oxygen's oxidation state.

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Is usually negative two.

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But not in peroxides or superoxides.

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Hydrogen peroxide is the only one you will need to remember for this course, so I wrote it down for you.

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See, I'm kind of nice. Sometimes.

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The oxidation number of hydrogen is plus one.

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Except for metal hydrides in which it is negative one.

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Here are a couple of examples of metal hydrides.

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Rule 5 repeats the monoatomic ion rule, but it is specific to groups one and two in the periodic table.

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It's also kind of unnecessary.

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The sum of the oxidation numbers of the atoms forming an ion is equal to the charge of the ion.

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If that molecule has no charge, the sum will be 0.

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This is one of the rules that you will use the most.

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Oxidation numbers don't have to be integers.

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However, they usually are.

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Let's practice assigning the oxidation number to each element in the following species.

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Pause the video.

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Write down your answers.

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And then come right back.

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Welcome back.

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The first example consists of two.

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Monatomic ion species, so the charge is the oxidation number.

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The second example, nitric acid, is a little more complicated.

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Hydrogen is usually plus one.

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And oxygen is negative 2.

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But what about the nitrogen?

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How did we know

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It was +5?

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Well, we know that the sum of the oxidation states must be 0, so this is what we know.

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1 - 6 is negative 5.

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So the question is.

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What minus five is 0?

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I know it sounds dumb when I put it that way, but that's OK because I'm dumb.

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And that is where the positive 5 for nitrogen comes from.

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The third example.

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Has a dichromate.

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The oxygen, as usual is negative 2.

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And the sum must be equal to the charge of the dichromate.

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So we ask.

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What minus 14 is negative 2.

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And the answer is 12 and therefore the oxidation state of chromium is positive 6.

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Your reference materials include this table with the most common oxidation states for most elements.

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Oxidation states not listed here are still possible.

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These are the most common, but they're not all inclusive.

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And that's all there is.

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There isn't any more.