What is a redox reaction?

A redox (or oxidation-reduction) reaction is a type of chemical reaction that involves a transfer of electrons between two species.

What is a "species"?

Great question! A chemical species is a term that refers to a set of atoms, molecules, or ions with the same chemical formula.

For example, in the thermite reaction:

\[ \text{Fe}_2\text{O}_3 + 2\text{Al} \rightarrow 2\text{Fe} + \text{Al}_2\text{O}_3 \]

the four species involved in this reaction are the reactants \( \text{Fe}_2\text{O}_3 \) and \( \text{Al} \), as well as the products \( \text{Fe} \) and \( \text{Al}_2\text{O}_3 \).

We can tell there has been a transfer of electrons if there is any change in the oxidation number between the reactants and the products.

Redox reactions are everywhere! Your body uses redox reactions to convert food and oxygen to energy plus water and \( \text{CO}_2 \), which we then exhale. The batteries in your electronics also rely on redox reactions, which you will hear more about when we learn electrochemistry. Can you find other examples of redox reactions happening around you?

An example and important terms

Redox reactions have some associated terms you should be comfortable using. We will go over these terms using the following example reaction:

\[ 2\text{Fe}_2\text{O}_3(\text{s}) + 3\text{C}(\text{s}) \rightarrow 4\text{Fe}(\text{s}) + 3\text{CO}_2(\text{g}) \]

Here are some questions we want to be able to answer:
1. Is this reaction a redox reaction, and how do we know?
2. If this is a redox reaction, what is being reduced and oxidized?

\[ \text{Remember LEO goes GER! and Oil Rig!} \]
\[ \text{OIL RIG: "Oxidation Is Loss" and "Reduction Is Gain"} \]
\[ \text{LEO the lion says GER: "Loss of Electron is Oxidation" and "Gain of Electron is Reduction"} \]

3. What is the reducing agent in this reaction?
4. What is the oxidizing agent in this reaction?

Question 1:
Based on the title of this article, we can make an educated guess about the first part of this
question. Yes, this is *probably* a redox reaction, but how do we know that for sure? We need to show there is an electron transfer occurring, and we can do that by checking if any oxidation numbers change from the reactants to the products.

If we find the oxidation numbers for each atom in the reactants and products, we get the following:

\[
2\text{FeO}_3(s) + 3\text{C(s)} \rightarrow 4\text{Fe(s)} + 3\text{CO}_2(g)
\]

\[
\begin{array}{cccccc}
& +3 & -2 & 0 & 0 & +4, -2 \\
\downarrow & \downarrow & \downarrow & \downarrow & \downarrow & \downarrow & \text{(Oxidation numbers)}
\end{array}
\]

What the heck are oxidation numbers?
C(s) and Fe(s): The oxidation number for pure elements is 0, so we know the oxidation number for C(s) and Fe(s) is 0.

FeO₃(s): The oxidation number for O is −2 and the charge on the compound is neutral, so the oxidation state of the iron atoms is +3.

CO₂(g): The oxidation of O is also −2 and the charge on the compound is neutral, so the oxidation number of the C +4.

We can use the oxidation numbers to answer the second part of question 1, because we can show that the oxidation numbers for carbon and iron are changing during the reaction from a transfer of electrons.

**Question 2:**
Carbon being oxidized because it is *losing electrons* as the oxidation number increases from 0 to +4.
Iron is being reduced because it is *gaining electrons* when the oxidation number decreases from +3 to 0.

**Question 3:**
The *reducing agent* is the reactant that is being oxidized (and thus causing something else to be reduced), so C(s) is the reducing agent.

**Question 4:**
The *oxidizing agent* is reactant that is being reduced (and thus causing something else to be oxidized), so FeO₃(s) is the oxidizing agent.

**Balancing a simple redox reaction using the half-reaction method**

Redox reactions can be split into reduction and oxidation *half-reactions*. Chemists use half-reactions to make it easier to see the electron transfer, and it also helps when balancing redox reactions. Let's write the half-reactions for another example reaction:
Al(s) + Cu^{2+}(aq) → Al^{3+}(aq) + Cu(s)

Is the above reaction balanced? Our atoms appear to be balanced: we have 1Al atom and 1Cu atom on each side of the arrow. However, when we add up the charges on the reactant side we get a 2+ which is not the same as the 3+ charge on the product side. We need to make sure both the atoms and the charges are balanced! We will use the *half-reaction method* to balance the reaction.

**Reduction half-reaction:** The reduction half-reaction shows the reactants and products participating in the reduction step. Since Cu^{2+} is being reduced to Cu(s) we might start by writing out that step:

\[ \text{Cu}^{2+}(aq) \rightarrow \text{Cu}(s) \]

However, this is *not* the correct half-reaction, because it is not charge-balanced. There is a net charge of 2+ on the reactant side and 0 on the product side. We can balance the charges by including the electrons being transferred, and then we will get our reduction half-reaction:

\[ \text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s) \]  
**Reduction half-reaction**

The balanced half-reaction tells us that Cu^{2+} is gaining 2e^- per copper atom to form Cu^0. So where are those electrons coming from? We can follow the trail of electrons to the oxidation half-reaction.

**Oxidation half-reaction:** The oxidation half-reaction shows the reactants and products participating in the oxidation step. This reaction will include the oxidation of Al(s) to Al^{3+} and we will also want to make sure the half-reaction is charge-balanced:

\[ \text{Al}(s) \rightarrow \text{Al}^{3+}(aq) + 3e^- \]  
**Oxidation half-reaction**

The oxidation half reaction tells us that each atom of Al(s) is losing 3e^- to form Al^{3+}.

We will combine the balanced half-reactions to get the balanced overall equation, but there is one more thing to check. *The electrons must cancel out in the overall equation.* Another way to think of this that we want to make sure that any electrons that are released in the oxidation half-reaction get used up in the reduction half-reaction. Otherwise we would have stray electrons floating around! That means we need the number of electrons being transferred in each half-reaction to be equal.

We can multiply the reduction half-reaction by 3 and multiply the oxidation half-reaction by 2 so both reactions involve the transfer of 6 electrons:

\[ 3 \times [\text{Cu}^{2+}(aq)+2e^- \rightarrow \text{Cu}(s)] \]  
3 × *reduction half-reaction*

\[ 2 \times [\text{Al}(s) \rightarrow \text{Al}^{3+}(aq)+3e^-] \]  
2 × *oxidation half-reaction*
Now that we have the same number of electrons in each half-reaction, we can add them together to get our overall balanced equation:

\[
\begin{align*}
6e^- + 3Cu^{2+}(aq) &\rightarrow 3Cu(s) \quad &\text{3\times\text{reduction half-reaction}} \\
2Al(s) &\rightarrow 2Al^{3+}(aq) + 6e^- \quad &\text{2\times\text{oxidation half-reaction}} \\
2Al(s) + 3Cu^{2+}(aq) &\rightarrow 2Al^{3+}(aq) + 3Cu(s) \quad &\text{Overall balanced reaction}
\end{align*}
\]

Lastly, we can check to see if any reactants and products appear on both sides. Since that is not the case here, we are done! Our reaction is balanced for both mass and charge.

**Summary**

We can identify redox reactions by checking for changes in oxidation number. Redox reactions can be split into oxidation and reduction half-reactions. We can use the half-reaction method to balance redox reactions, which requires that both mass and charge are balanced. Three common types of redox reactions are combustion, disproportionation, and single replacement reactions.

**Attributions:**

This article was adapted from the following articles:

1. "Oxidation-Reduction Reactions" from Beginning Chemistry, CC-BY-NC-SA 3.0
2. "Balancing Redox Reactions" from UC Davis ChemWiki, CC-BY-NC-SA 3.0
3. "The importance of redox reactions in metabolism" from UC Davis ChemWiki, CC-BY-NC-SA 3.0

The modified article is licensed under a CC-BY-NC-SA 4.0 license.

**Additional References:**
