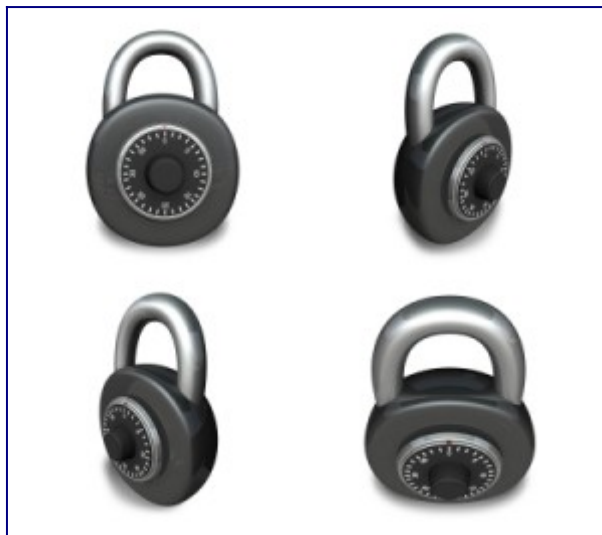


An introduction to redox reactions

Your teacher has started talking about redox reactions, and though the term sounds not unlike things you're familiar with (Maalox, Hydrox, combination locks, and the like), you're probably a little less comfortable with the redox phenomenon. Instead of worry about it, let's explain this so you enjoy it just as much as the other items with "ox" sounds above:



Get ready for some fun!

Before we get started: The joys of redox terminology

Here are some handy terms you'll need to know before you start learning about redox reactions. They may not make much sense now, but we'll be discussing all of them later, so you can just refer to this section if you don't get it.

Oxidation state: The formal charge (either positive or negative) we assign to an element. This may not make a lot of sense, but maybe this will help:

The oxidation state of elements in their pure form is zero. If you see Al or F₂, both aluminum and fluorine have an oxidation state of zero.

The oxidation state of a monatomic ion is equal to the charge on that ion. For example, the sodium ion, Na⁺, has an oxidation state of +1. Likewise, the fluoride ion (F⁻) has an oxidation state of -1.

Rules for neutral molecules: The oxidation state of elements in a neutral compound is bound by the following rule: In a neutral compound, the sum of the oxidation states of all of the elements must be zero. However, this doesn't necessarily tell us what these states are, so follow these rules:

- Assume the most electronegative element has the oxidation state you'd expect it to have when it's acting like an ionic compound. For example, in NaF, fluorine is in the form of the fluoride ion (F⁻) with a -1 charge. In oxygen difluoride (OF₂), we would also assume fluorine have a -1 charge.
- Assume the other elements have the oxidation that's needed to bring the overall charge of the molecule to zero. In NaF, Na must have a +1 oxidation state to balance

out the -1 of the fluoride. In OF_2 , oxygen must have a +2 oxidation state to balance out the two -1 charges (one for each fluoride atom).

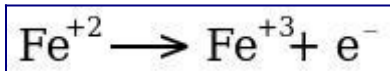
Rules for polyatomic ions: Same as that for neutral molecules, except that the overall charge has to add up to the overall charge of the ion. For example, in the SO_4^{2-} ion, each oxygen has a -2 oxidation state (it's the most electronegative element) for a total negative charge of -8. For sulfur to bring that negative charge back to -2, it has to have an oxidation state of +6.

Oxidation: Oxidation occurs when an element loses electrons, giving it a more positive charge. For example, if iron goes from a 0 to a +2 oxidation state, it is said to be oxidized because it has lost electrons.

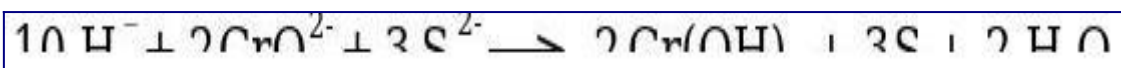
Reduction: Reduction occurs when an element gains electrons, giving it a more negative charge. For example, if iron went from a +3 to a +2 oxidation state, it is said to be reduced because it has gained electrons.

Redox reaction: A redox reaction is a reaction that contains both an oxidation process and a reduction process. Actually, this pretty much describes any electrochemical process, because you can't reduce something (i.e. give it electrons) unless something has given these electrons to it by being oxidized. As a result, in redox reactions, the thing that gets reduced is called an "oxidizing agent" because it takes electrons from something else, and the thing that gets oxidized is called a "reducing agent" because it gives electrons to something else.

Half-reaction: This is a part of a chemical reaction that shows only the oxidation step or the reduction step. For example, if iron were to go from the +2 to the +3 oxidation state, a half-reaction for this might be:



Please note that in every half-reaction (and in every reaction at all, for that matter), the sum of the charges on the product side must be the same as the sum of the charges on the reagent side. In this case, both sides of the equation have a net +2 charge.

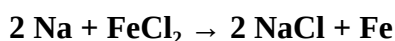


Protip: Your teacher won't think that this "half reaction" is funny if you write it on an exam.

Redox reaction: A chemical reaction in which one (or more) of the elements present is oxidized, and one (or more) of the elements is reduced. We'll be seeing a lot of these.

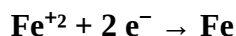
The joys of redox reactions:

The main idea behind a redox reaction is this: The oxidation states of at least two elements changes. One of the elements give electrons to the other – this causes the electron that gives away the electrons to get oxidized (it has more + charge) and the one that gets the electrons to get reduced (it has more – charge). One example of such a reaction is this:



In this example, sodium is oxidized, as it goes from an oxidation state of zero to an oxidation state of +1. Iron, on the other hand, is reduced, because it goes from an oxidation state of +2 to an

oxidation state of zero. If we were to break this down into two half reactions, it might look something like this:



You may have noticed that chlorine doesn't show up anywhere in either equation. That's because it doesn't change its oxidation state at all, so the number of electrons it has remains the same. Since moving electrons is the whole point of this reaction, we ignore it. It's just loafing around and looking at the action, which is why it's called a *spectator ion*.

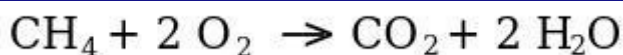
Another feature is that the first half reaction with sodium generates one electron, while the second half reaction with iron requires two electrons. This isn't anything particularly weird – it just means that two sodium atoms need to lose one electron each to give away a total of two electrons. Given the complete equation for this process, that's not really all that surprising because there are two sodium atoms involved.

Where you're likely to see redox reactions

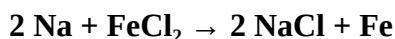
There are a lot of redox reactions going on around you that you're probably not even familiar with. These include things from apples going brown in air, the rusting of an old car's bumper, and fresh new batteries becoming useless old dead batteries.

As far as the six types of reaction go, some of these are redox reactions as well:

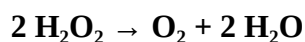
- **Combustion reactions** take place when oxygen reacts with an organic molecule to make carbon dioxide and water. A familiar example of this is the combustion of methane:



- In this reaction, carbon starts with a -4 oxidation state (hydrogen has +1) and ends up with a +4 oxidation state (oxygen has -2). It is oxidized in this reaction.
 - Oxygen starts with a 0 oxidation state (it's a pure element) and ends up with a -2 oxidation state. It is reduced in this reaction.
 - Hydrogen is a spectator and doesn't do anything. It orders pizza and waits for things to settle down.
- **Single displacement reactions** involve the oxidation of a pure element (usually a metal) and the reduction of an element in an ionic compound (also a metal). The example from above shows this, where sodium is oxidized from 0 to +1, and iron is reduced from +2 to 0.



- **Synthesis and decomposition reactions** typically involve changes in oxidation state. For example, in the synthesis reaction $2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O}$, hydrogen is oxidized from 0 to +1, and oxygen is reduced from 0 to -2. Similarly, if we decompose hydrogen peroxide into oxygen and hydrogen gas via the equation:



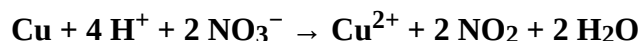
We see an interesting case where hydrogen doesn't change its oxidation state at all (it's +1) but oxygen is both oxidized *and* reduced. Initially, the oxidation state of oxygen is -1, but it ends up as 0 in O₂ (representing an oxidation) and -2 in water (representing an reduction). Pretty cool, huh?



James only became cool after learning about oxidation and reduction.

On the other hand, some types of reactions are nearly never redox reactions: Double displacement and acid-base reactions. This makes sense, because the ions in both types of reactions just switch around.

However, just because acid-base reactions aren't another example of redox reactions doesn't mean that acids won't be involved in redox reactions. For example, nitric acid is a very strong oxidizing agent, and will undergo redox reactions such as the one below:



It should be noted, however, that just because nitric acid is involved doesn't mean that this is an acid-base reaction. Given the process that takes place, this is more of a decomposition kind of thing.

We'll do more redox fun in a while, but for now, relax and practice being cool!

Image credits:

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