

LESSON PLAN: Redox Reactions (Class 1 of 2)

Introduction:

Subject: Oxidation-Reduction Reactions (Redox)

Learner Outcomes/Objectives: Students will learn about real life occurrences of oxidation-reduction reactions. They will learn how these redox reactions are behind many known phenomena, from paper becoming yellow to apples becoming brown to fire to the most destructive of explosions. They will review how oxidation-reduction reactions involve electron transfer between atoms. Students will learn the origin of the words "oxidation" and "reduction". Finally, students will receive practical advice on fire safety.

Rationale: A spectacular sequence from the movie *Daylight* and two YouTube segments provide an effective starting point to introduce or review oxidation-reduction reactions in chemistry and to help students to relate theory learned in class to real life, including fire safety recommendations.

Description of the Film Clip: A car being chased by the police drives into a tunnel from New York to New Jersey. As the speeds through the tunnel, weaving in and out between the other cars, it finally crashes and ends up flying into a truck loaded with flammable materials. The snippet shows how the subsequent explosion sets off a shock wave of flames through the entire tunnel. One YouTube segment shows a shock wave when a rocket fuel plant explodes. The other shows an example of spontaneous combustion of a large pile of hay.

Helpful Background:

Explosions, like the one shown in *Daylight* and many other movies, are nothing more than oxidation-reduction reactions which happen very rapidly. Slower oxidation-reduction reactions are all around us, like the rusting of some metals or a bitten apple getting brown if left in the air. The basis of oxidation-reduction (in short: "redox") reactions is the transfer of electrons between atoms. The atoms losing electrons are being oxidized while those receiving electrons are being reduced.

The name oxidation comes from the fact that forming a bond with oxygen is equivalent to losing electrons to the oxygen atoms, because oxygen is very electronegative. This means that oxygen exerts a strong attraction over the electrons of other atoms and basically attaches them to itself, thus gaining negative charges (it becomes a negative ion). This process leaves the other atom positively charged (as a positive ion). The charge difference is then the basis for an electrostatic attraction between both, known as ionic bonding.

Recall that in the other type of bonding between atoms, covalent bonding, both atoms involved have a similar tendency to attract electrons (they both have electronegativity) and end up sharing electrons. This is a much stronger bond than the ionic bonding of redox reactions. See Figure 1.

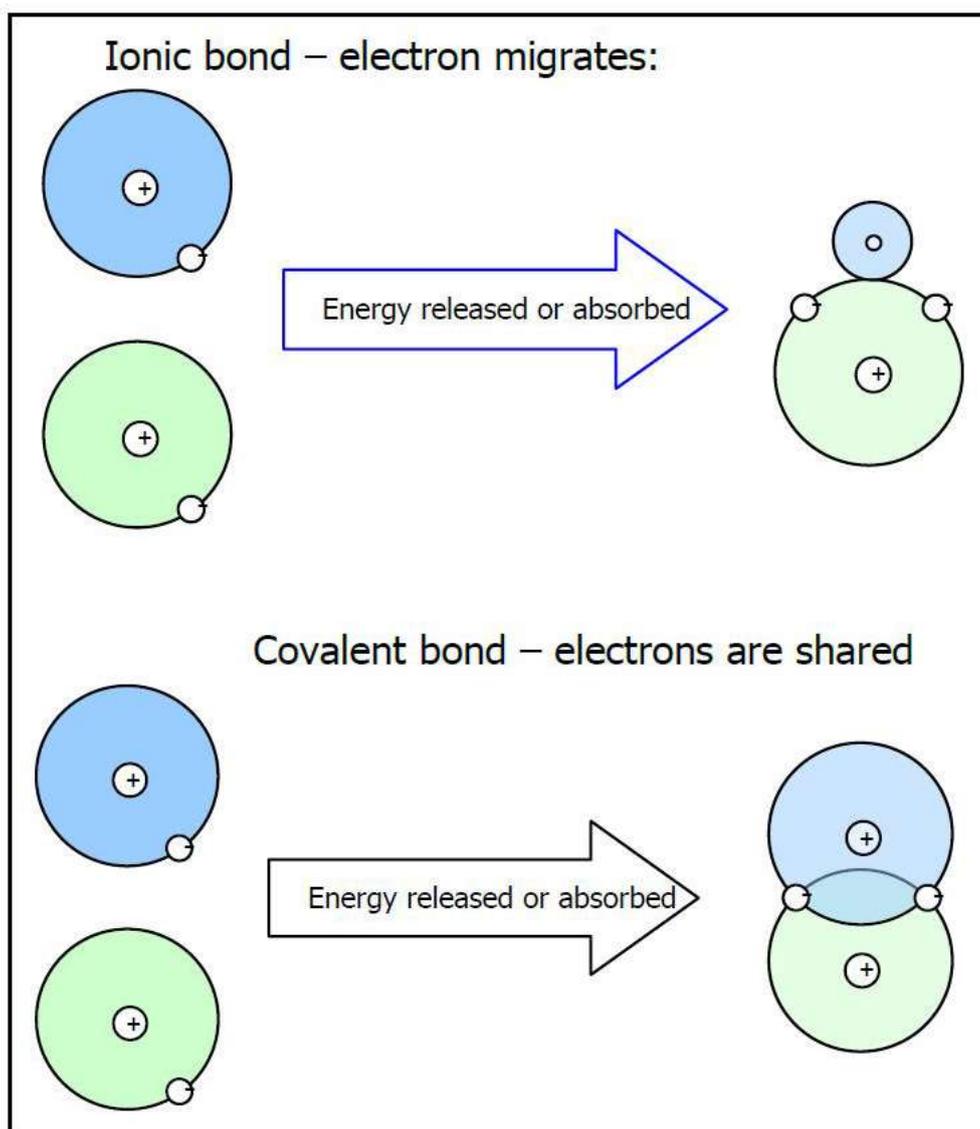


Figure 1. Atomic Bonding: Ionic and covalent bonds with one electron

So, the term oxidation can refer to the bonding with oxygen or the loss of electrons. These are two different ways to describe the same process. There is a third meaning of the term oxidation, which is the loss of whole hydrogen atoms by a molecule. This is related to both previous meanings in that oxidizing agents remove hydrogen atoms from (generally organic) compounds when they are available.

Scientists have devised a way to unify all three meanings in one, defined as "the increase in oxidation number." This also explains the origin of the word reduction, as it then means "reducing the oxidation number." For a detailed explanation of oxidation numbers.

Let us now review the oxidation reactions in terms of their speed and release of heat and light. Old paper becoming yellow, or potatoes and apples becoming brown, do not release either. In the rusting of iron, however, it is known that heat is released very slowly. It is not noticeable, but it can be measured.

The next step up in oxidation is combustion. A steady release of heat and light is produced. A frequent example is organic compounds made of carbon and hydrogen that are oxidized by gaseous oxygen. This is the case of fossil fuels in their different forms used to power engines, or provide heat in fires or kitchens. Combustion can happen to gases, liquids and solids. A particular case is the combustion of metals, which release light of different colors and is used in fireworks.

Combustion needs the fuel to reach a certain temperature in order to start. When this temperature is very low for a particular material, it can happen that it ignites spontaneously. What happens in the clip from *Daylight* is more violent still: what we know as an explosion is a very rapid, practically instantaneous oxidation, with intense heat and light release, plus an increased pressure due to rapid release of gases (products of the redox reaction). This makes a loud noise and creates a shock wave. If this shock wave is subsonic what really drives the propagation of the fire in the explosive blast is the rapid heat transfer to the materials above the temperature they need to ignite. This is called a deflagration. When the shock wave is supersonic, the high overpressures that are created when it heats materials above their autoignition temperature (kindling point) drives the explosion. This is called detonation and is the most violent oxidation mechanism.

Advice on how to deal with a person in flames: As seen in the final sequence of the segment, a person who is in flames can be helped by making them roll on the floor and by wrapping them in a blanket or anything that can stop oxygen from feeding the flame. Usually people in this situation are in shock or panic, and they need to be helped, or even forced to let themselves be helped in this way. In some places there are special fire blankets placed beside fire extinguishers.

DEFINITIONS OF OXIDATION AND REDUCTION (REDOX)

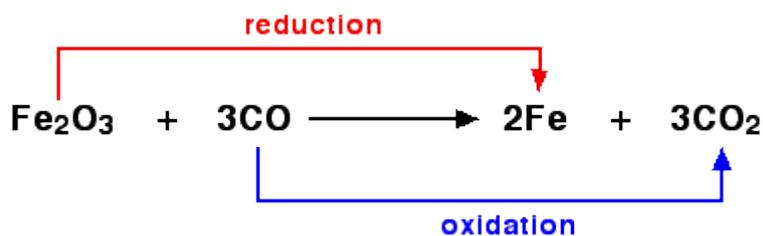
This section looks at the various definitions of oxidation and reduction (redox) in terms of the transfer of oxygen, hydrogen and electrons. It also explains the terms oxidising agent and reducing agent.

Oxidation and reduction in terms of oxygen transfer

Definitions

- Oxidation is gain of oxygen.
- Reduction is loss of oxygen.

For example, in the extraction of iron from its ore:



Because both reduction and oxidation are going on side-by-side, this is known as a redox reaction.

Oxidising and reducing agents

An oxidising agent is substance which oxidises something else. In the above example, the iron(III) oxide is the oxidising agent.

A reducing agent reduces something else. In the equation, the carbon monoxide is the reducing agent.

- Oxidising agents give oxygen to another substance.
- Reducing agents remove oxygen from another substance.

Oxidation and reduction in terms of hydrogen transfer

These are old definitions which aren't used very much nowadays. The most likely place you will come across them is in organic chemistry.

Definitions

- Oxidation is loss of hydrogen.
- Reduction is gain of hydrogen.

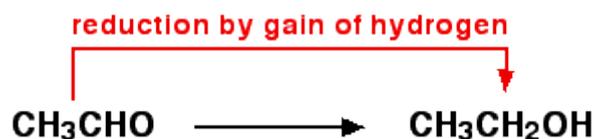
Notice that these are exactly the opposite of the oxygen definitions.

For example, ethanol can be oxidised to ethanal:



You would need to use an oxidising agent to remove the hydrogen from the ethanol. A commonly used oxidising agent is potassium dichromate(VI) solution acidified with dilute sulphuric acid.

Ethanal can also be reduced back to ethanol again by adding hydrogen to it. A possible reducing agent is sodium tetrahydridoborate, NaBH₄. Again the equation is too complicated to be worth bothering about at this point.



An update on oxidising and reducing agents

- Oxidising agents give oxygen to another substance or remove hydrogen from it.
- Reducing agents remove oxygen from another substance or give hydrogen to it.

Oxidation and reduction in terms of electron transfer

This is easily the most important use of the terms oxidation and reduction at A level.

Definitions

- Oxidation is loss of electrons.
- Reduction is gain of electrons.

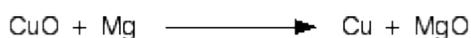
It is essential that you remember these definitions. There is a very easy way to do this. As long as you remember that you are talking about electron transfer:

OIL RIG

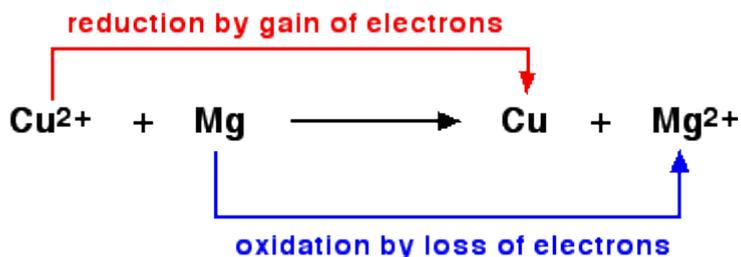
oxidation is loss color: red;">reduction is gain

A simple example

The equation shows a simple redox reaction which can obviously be described in terms of oxygen transfer.



Copper(II) oxide and magnesium oxide are both ionic. The metals obviously aren't. If you rewrite this as an ionic equation, it turns out that the oxide ions are spectator ions and you are left with:



A last comment on oxidising and reducing agents

If you look at the equation above, the magnesium is reducing the copper(II) ions by giving them electrons to neutralise the charge. Magnesium is a reducing agent.

Looking at it the other way round, the copper(II) ions are removing electrons from the magnesium to create the magnesium ions. The copper(II) ions are acting as an oxidising agent.

Warning!

This is potentially very confusing if you try to learn both what oxidation and reduction mean in terms of electron transfer, and also learn definitions of oxidising and reducing agents in the same terms.

Personally, I would recommend that you work it out if you need it. The argument (going on inside your head) would go like this if you wanted to know, for example, what an oxidising agent did in terms of electrons:

- An oxidising agent oxidises something else.
- Oxidation is loss of electrons (OIL RIG).
- That means that an oxidising agent takes electrons from that other substance.
- So an oxidising agent must gain electrons.
- Or you could think it out like this:
- An oxidising agent oxidises something else.
- That means that the oxidising agent must be being reduced.
- Reduction is gain of electrons (OIL RIG).
- So an oxidising agent must gain electrons.
- Understanding is a lot safer than thoughtless learning!

WRITING IONIC EQUATIONS FOR REDOX REACTIONS

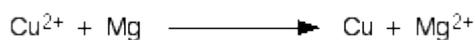
This section explains how to work out electron-half-reactions for oxidation and reduction processes, and then how to combine them to give the overall ionic equation for a redox reaction. This is an important skill in inorganic chemistry.

Don't worry if it seems to take you a long time in the early stages. It is a fairly slow process even with experience. Take your time and practice as much as you can.

Electron-half-equations

What is an electron-half-equation?

When magnesium reduces hot copper(II) oxide to copper, the ionic equation for the reaction is:



You can split the ionic equation into two parts, and look at it from the point of view of the magnesium and of the copper(II) ions separately. This shows clearly that the magnesium has lost two electrons, and the copper(II) ions have gained them.



These two equations are described as "electron-half-equations" or "half-equations" or "ionic-half-equations" or "half-reactions" - lots of variations all meaning exactly the same thing!

Any redox reaction is made up of two half-reactions: in one of them electrons are being lost (an oxidation process) and in the other one those electrons are being gained (a reduction process).

Working out electron-half-equations and using them to build ionic equations

In the example above, we've got at the electron-half-equations by starting from the ionic equation and extracting the individual half-reactions from it. That's doing everything entirely the wrong way round!

In reality, you almost always start from the electron-half-equations and use them to build the ionic equation.

Example 1: The reaction between chlorine and iron(II) ions

Chlorine gas oxidises iron(II) ions to iron(III) ions. In the process, the chlorine is reduced to chloride ions.

You would have to know this, or be told it by an examiner. In building equations, there is quite a lot that you can work out as you go along, but you have to have somewhere to start from!

You start by writing down what you know for each of the half-reactions. In the chlorine case, you know that chlorine (as molecules) turns into chloride ions:



The first thing to do is to balance the atoms that you have got as far as you possibly can:



ALWAYS check that you have the existing atoms balanced before you do anything else. If you forget to do this, everything else that you do afterwards is a complete waste of time!

Now you have to add things to the half-equation in order to make it balance completely.

All you are allowed to add are:

- electrons
- water
- hydrogen ions (unless the reaction is being done under alkaline conditions - in which case, you can add hydroxide ions instead)

In the chlorine case, all that is wrong with the existing equation that we've produced so far is that the charges don't balance. The left-hand side of the equation has no charge, but the right-hand side carries 2 negative charges. That's easily put right by adding two electrons to the left-hand side. The final version of the half-reaction is:



Now you repeat this for the iron(II) ions. You know (or are told) that they are oxidised to iron(III) ions. Write this down:



The atoms balance, but the charges don't. There are 3 positive charges on the right-hand side, but only 2 on the left.

You need to reduce the number of positive charges on the right-hand side. That's easily done by adding an electron to that side:



Combining the half-reactions to make the ionic equation for the reaction

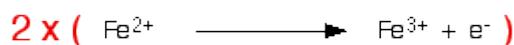
What we've got at the moment is this:

This reaction needs 2 electrons . . .



. . . but this is only producing 1 electron

It is obvious that the iron reaction will have to happen twice for every chlorine molecule that reacts. Allow for that, and then add the two half-equations together.



But don't stop there!! Check that everything balances - atoms and charges. It is very easy to make small mistakes, especially if you are trying to multiply and add up more complicated equations.

You will notice that I haven't bothered to include the electrons in the added-up version. If you think about it, there are bound to be the same number on each side of the final equation, and so they will cancel out. If you aren't happy with this, write them down and then cross them out afterwards!

Example 2: The reaction between hydrogen peroxide and manganate(VII) ions

The first example was a simple bit of chemistry which you may well have come across. The technique works just as well for more complicated (and perhaps unfamiliar) chemistry.

Manganate(VII) ions, MnO_4^- , oxidise hydrogen peroxide, H_2O_2 , to oxygen gas. The reaction is done with potassium manganate(VII) solution and hydrogen peroxide solution acidified with dilute sulphuric acid.

During the reaction, the manganate(VII) ions are reduced to manganese(II) ions.

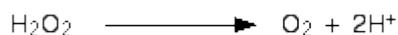
Let's start with the hydrogen peroxide half-equation. What we know is:



The oxygen is already balanced. What about the hydrogen?

All you are allowed to add to this equation are water, hydrogen ions and electrons. If you add water to supply the extra hydrogen atoms needed on the right-hand side, you will mess up the oxygens again - that's obviously wrong!

Add two hydrogen ions to the right-hand side.

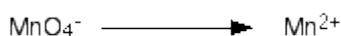


Now all you need to do is balance the charges. You would have to add 2 electrons to the right-hand side to make the overall charge on both sides zero.

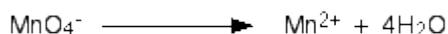


Now for the manganate(VII) half-equation:

You know (or are told) that the manganate(VII) ions turn into manganese(II) ions. Write that down.



The manganese balances, but you need four oxygens on the right-hand side. These can only come from water - that's the only oxygen-containing thing you are allowed to write into one of these equations in acid conditions.



By doing this, we've introduced some hydrogens. To balance these, you will need 8 hydrogen ions on the left-hand side.



Now that all the atoms are balanced, all you need to do is balance the charges. At the moment there are a net 7+ charges on the left-hand side (1- and 8+), but only 2+ on the right. Add 5 electrons to the left-hand side to reduce the 7+ to 2+.

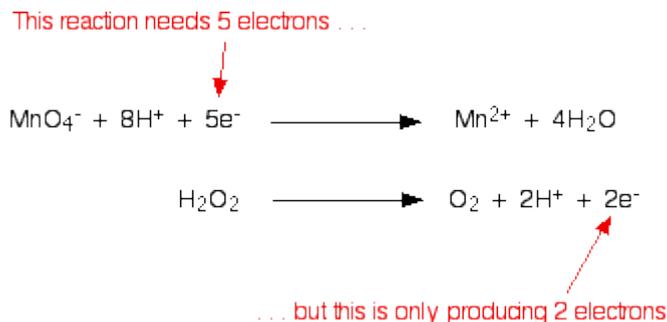


This is the typical sort of half-equation which you will have to be able to work out. The sequence is usually:

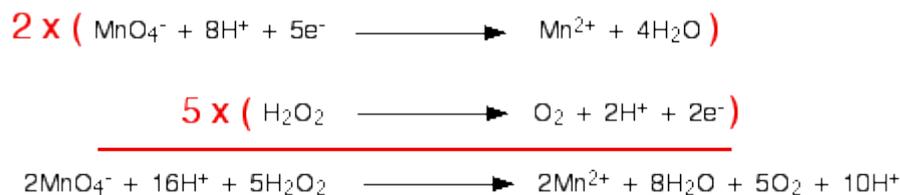
- Balance the atoms apart from oxygen and hydrogen.
- Balance the oxygens by adding water molecules.
- Balance the hydrogens by adding hydrogen ions.
- Balance the charges by adding electrons.

Combining the half-reactions to make the ionic equation for the reaction

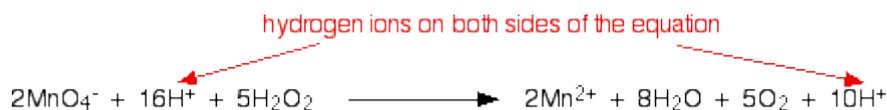
The two half-equations we've produced are:



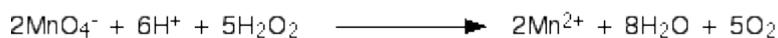
You have to multiply the equations so that the same number of electrons are involved in both. In this case, everything would work out well if you transferred 10 electrons.



But this time, you haven't quite finished. During the checking of the balancing, you should notice that there are hydrogen ions on both sides of the equation:



You can simplify this down by subtracting 10 hydrogen ions from both sides to leave the final version of the ionic equation - but don't forget to check the balancing of the atoms and charges!



You will often find that hydrogen ions or water molecules appear on both sides of the ionic equation in complicated cases built up in this way. Always check, and then simplify where possible.

Example 3: The oxidation of ethanol by acidified potassium dichromate(VI)

This technique can be used just as well in examples involving organic chemicals. Potassium dichromate(VI) solution acidified with dilute sulphuric acid is used to oxidise ethanol, $\text{CH}_3\text{CH}_2\text{OH}$, to ethanoic acid, CH_3COOH .

The oxidising agent is the dichromate(VI) ion, $\text{Cr}_2\text{O}_7^{2-}$. This is reduced to chromium(III) ions, Cr^{3+} .

We'll do the ethanol to ethanoic acid half-equation first. Using the same stages as before, start by writing down what you know:



Balance the oxygens by adding a water molecule to the left-hand side:



Add hydrogen ions to the right-hand side to balance the hydrogens:



And finally balance the charges by adding 4 electrons to the right-hand side to give an overall zero charge on each side:



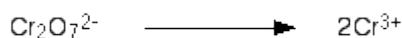
The dichromate(VI) half-equation contains a trap which lots of people fall into!

Start by writing down what you know:

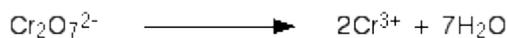


What people often forget to do at this stage is to balance the chromiums. If you don't do that, you are doomed to getting the wrong answer at the end of the process! When you come to

balance the charges you will have to write in the wrong number of electrons - which means that your multiplying factors will be wrong when you come to add the half-equations . . . A complete waste of time!



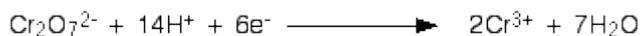
Now balance the oxygens by adding water molecules . . .



. . . and the hydrogens by adding hydrogen ions:



Now all that needs balancing is the charges. Add 6 electrons to the left-hand side to give a net 6+ on each side.



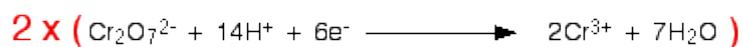
Combining the half-reactions to make the ionic equation for the reaction

What we have so far is:

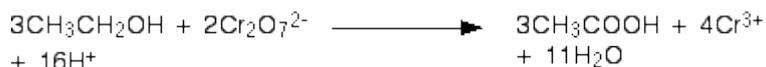


What are the multiplying factors for the equations this time? The simplest way of working this out is to find the smallest number of electrons which both 4 and 6 will divide into - in this case, 12. That means that you can multiply one equation by 3 and the other by 2.

The multiplication and addition looks like this:



Now you will find that there are water molecules and hydrogen ions occurring on both sides of the ionic equation. You can simplify this to give the final equation:



PRACTICE PROBLEMS

- $\text{H}_3\text{PO}_2 + \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{H}_3\text{PO}_4 + \text{Cr}^{3+}$
- $\text{Ag} + \text{NO}_3^- \rightarrow \text{Ag}^+ + \text{NO}$

3. $\text{Zn} + \text{NO}_3^- \rightarrow \text{Zn}^{2+} + \text{NH}_4^+$
4. $\text{Cr}_2\text{O}_7^{2-} + \text{C}_2\text{H}_4\text{O} \rightarrow \text{C}_2\text{H}_4\text{O}_2 + \text{Cr}^{3+}$
5. $\text{MnO}_4^- + \text{C}_2\text{O}_4^{2-} \rightarrow \text{MnO}_2 + \text{CO}_2$
6. $\text{ClO}^- + \text{Fe}(\text{OH})_3 \rightarrow \text{Cl}^- + \text{FeO}_4^{2-}$
7. $\text{HO}_2^- + \text{Cr}(\text{OH})_3 \rightarrow \text{CrO}_4^{2-} + \text{OH}^-$
8. $\text{N}_2\text{H}_4 + \text{Cu}(\text{OH})_2 \rightarrow \text{N}_2 + \text{Cu}$