# Redox

Redox is the transfer of electrons from one species to another.

Oxidation: Loss of electrons

Reduction: Gain of electrons

This is remembered using OIL RIG or Oxidation Is Loss Reduction Is Gain.

or LEO the lion says GER. Loss of Electrons is Oxidation and Gain of Electrons is Reduction.

The following is a redox reaction:

$Mg + O_2 \Rightarrow 2MgO$

In this case you can clearly see oxygen has been added to magnesium. The magnesium has been oxidised.

On a deeper level the magnesium has changed from a neutral Mg to an $Mg^{2+}$ ion.

Oxidation: $Mg \Rightarrow Mg^{2+} + 2e^-$. The magnesium has lost 2 electrons.

The oxygen has changed from a neutral $O_2$ molecule to an $O^{2-}$ ion.

Reduction: $O_2 + 4e^- \Rightarrow 2O^{2-}$. The oxygen molecule has gained 2 electrons.

Redox assessment comes in 5 main parts:

1. You are to calculate the oxidation number of different atoms
2. You are to balance half and full redox equations
3. You must understand and draw galvanic cells
4. You must understand and draw electrolytic cells
5. Tell if a reaction will or not occur

<u>Oxidation numbers</u>:

An easier method than figuring out what has gained or lost electrons is the use of the idea of “redox numbers”. Oxidation numbers follow the following rules:

1. Hydrogen is +1 (except when a metal hydride, then it is -1)
2. Oxygen is -2 (except when a peroxide, then it is -1)
3. The redox numbers of a molecule add up to the overall charge of the molecule/ion.
4. Any substance just made of one uncharged atom is 0

Calculate the redox numbers of each atom and work out which molecule is being reduced and oxidised (if any! – Not all reactions are redox reactions).

1. $Mg + HCl \Rightarrow$
2. $4Fe + 3O_2 + 2H_2O \Rightarrow 4FeO(OH)$
3. $C_6H_{12}O_6 + 6O_2 \Rightarrow 6CO_2 + 6H_2O$ + energy
4. $C_4H_{10} + O_2 \Rightarrow$
5. $2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \Rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$
6. $H_2O_{2(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$

Find the oxidation numbers of the following:

1. $Cr_2O_7^{2-}$
2. $ClO_3^-$
3. NaH
4. $K_2O_2$

Balancing redox reactions:

Basic Example: $2Na_{(s)}+Cl_{2(g)}$ =>$2NaCl_{(s)}$

Oxidation: $Na_{(s)}$=>$Na^{+}+e^{-}$ (sodium has lost an electron, or redox number has increased)

Reduction: $Cl_{2(g)}+2e^{-}$=>$2Cl^{-}$

**Must** balance the electrons.

Final equation is $2Na_{(s)}+ Cl_{2(g)}$ =>$2NaCl_{(s)}$as expected.

Advanced Example: Hydrogen peroxide is mixed with acidified permanganate.

$MnO_4^{-}$ =>$Mn^{2+}$

Step 1: Balance main atoms (already done)
Step 2: Add waters to balance oxygens. $MnO_4^{-}$ =>$Mn^{2+}+4H_2O$
Step 3: Add hydrogen ions to balance hydrogens. $MnO_4^{-}+8H^{+}$ =>$Mn^{2+}+4H_2O$
Step 4: Add electrons to balance charge. $MnO_4^{-}+8H^{+}5e^{-}$=>$Mn^{2+}+4H_2O$

One half-equation is now balanced.(remember the check!)

Now balance the half equation for $H_2O_2$ => $O_2$ then add them together.

Try balancing the half equations below:

1. $Cr_2O_7^{2-}$ =>$Cr^{3+}$

2. $IO_3^{-}$ => $I_2$

3. $H_2C_2O_4$ => $CO_2$

4. $SO_4^{2-}$ => S

Reactions:

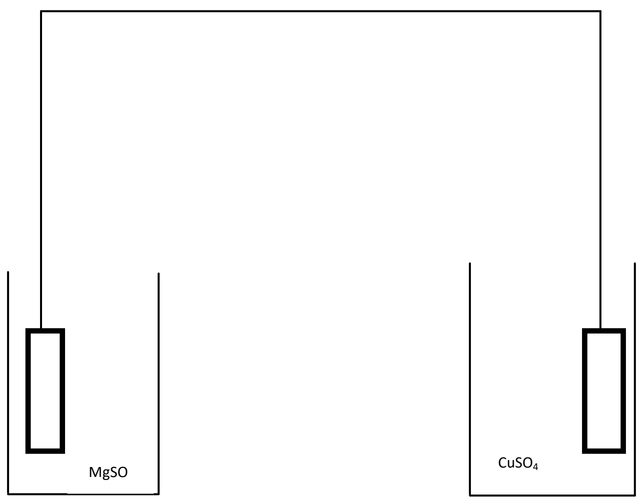
1. Write the net ionic equation for the following reactions and what observations you could make
    a. Copper solid is placed in hydrochloric acid

    b. Steel Wool is placed in a bowl of copper(II) sulfate

    c. Chlorine gas is bubbled through a solution of sodium bromide.

    d. Bromine liquid is added to sodium chloride solution.

Electrochemical cells:




For the following cells find the:

a) Anode and cathode location and reactions
b) Direction of electron flow
c) Direction of ion migration
d) Voltage

1. $Mg/Mg^{2+}$ / $Cu/Cu^{2+}$ with an ammonia nitrate salt bridge
2. Bubbling hydrogen and chlorine gas into water

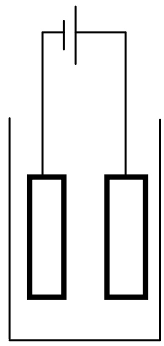
Electrolytic cells:

Electrolytic cells do not happen spontaneously so require an external energy source to force them to happen.

Example: 2 platinum electrodes are placed in a copper Iodide solution.

$Cu^{2+} + 2e \rightarrow Cu \quad +0.34$

$2I^{-} \rightarrow I_2 + 2e \qquad -0.54$



Try creating an electrolytic cell for molten sodium chloride. (when molten both elements exist as ions)