Oxidation and Reduction Definitions

Oxidation is **gain** of oxygen
Reduction is **loss** of oxygen

Oxidation is **loss** of electrons
Reduction is **gain** of electrons

Oxidation is **increase** in oxidation number
Reduction is **decrease** in oxidation number

Oxidation and Reduction

Oxidation is **<u>loss</u>** of electrons
Reduction is **<u>gain</u>** of electrons

Loss of **E**lectrons is **O**xidation
Gain of **E**lectrons is **R**eduction

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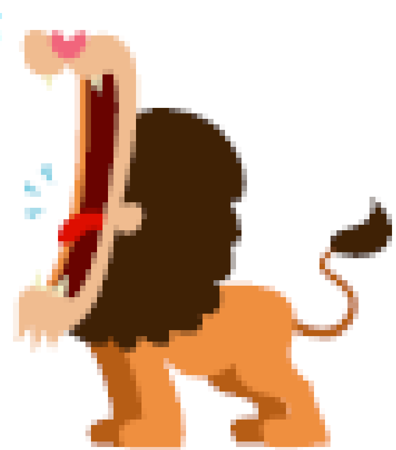

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Recognising oxidation and reduction

For example:

Rust formation – iron is converted into iron oxide

Iron (Fe) is oxidised as it gained oxygen to form iron oxide (Fe_2O_3)

$4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$

Formation of copper metal by the addition of magnesium to a copper nitrate solution

Copper ions are reduced as they gained electrons to form copper metal

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

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pg 6, 7 , 8

Do now:

What are the three different ways we can define redox reactions?

Oxidation is : gain in oxygen
loss of electrons
increase in oxidation number

Use your definitions to decide what species is getting oxidised and what species is getting reduced in the following equations?

$Zn + Mg^{2+} \rightarrow Zn^{2+} + Mg$

Zn is getting oxidised
Mg^{2+} is getting reduced

$Cu + Cl_2 \rightarrow CuCl_2$

Cu is getting oxidised
Cl_2 is getting reduced

Oxidants and reductants

Two definitions:

An **oxidant** gets **reduced** itself. It oxidises other things.

An **reductant** gets **oxidised** itself. It reduces other things.

Recognising oxidants and reductants

For example:
Rust formation – iron is converted into iron oxide

Iron is oxidised as it gained oxygen. Iron is the reductant as it got oxidised (and it reduced oxygen in the process)

Oxygen is reduced. Oxygen is the oxidant as it got reduced (and it oxidised iron in the process)

Formation of copper metal by the addition of magnesium to a copper nitrate solution

Copper is reduced as it gained electrons. Copper is the oxidant. Magnesium is oxidised as it lost electrons. Magnesium is the reductant.

Calculating oxidation numbers

We can work out oxidation numbers for all elements in a compound.

We then follow elements through a reaction and if the oxidation number changes then a redox reaction has taken place.

Increase in ON is: oxidation
Decrease in ON is: reduction

Calculating oxidation numbers

There are steps to follow when calculating oxidation numbers

- All elements by themselves are 0

  eg. Zn The oxidation number of Zn is 0

  eg. Cl_2 The oxidation number of Cl is 0

- For all <u>mono</u>atomic ions the oxidation number is the charge on the ion

  eg. Zn^{2+} The oxidation number of Zn^{2+} is +2

  eg. Cl^- The oxidation number of Cl^- is -1

- For all <u>poly</u>atomic ions or compounds the oxidation number of all the elements in the ion or compound add to the charge on the ion

  eg. SO_4^{2-} the sum of the oxidation numbers of O and S have to equal -2

Calculating oxidation numbers

- H always has an oxidation number of +1
- O always has an oxidation number of -2
  (apart from peroxide, H_2O_2, where it is -1)

Using these rules we can work out the oxidation number of elements in polyatomic molecules and ions.

For example: Calculate the oxidation number of the underlined elements.

$\underline{Cu}$	$\underline{Br}^-$	$\underline{Mg}^{2+}$	$\underline{N}H_3$	$\underline{S}O_2$	$\underline{N}O_3^-$	$\underline{Fe}_2O_3$	$\underline{S}_2O_3^{2-}$
0	**-1**	**+2**	**-3**	**+4**	**+5**	**+3**	**+2**

Workbook pg 9

Half equations

Show the transfer of electrons.

$2\ Mg + O_2 \rightarrow 2\ MgO$ ionic compound

$2\ Mg + O_2 \rightarrow 2\ Mg^{2+} + 2\ O^{2-}$

$Mg \rightarrow Mg^{2+}$ **+ 2 e** Oxidation (loss of electrons)

O_2 **+ 4 e** $\rightarrow 2\ O^{2-}$ Reduction (gain in electrons)

Metals will be oxidised to form their ions.
Non metals will be reduced to form their ions.

Balancing more complicated half equations

Write down reactant and product   (IO_3^- example on board)

Balance all ions that aren't O or H

Add water (H_2O) to balance O

Add hydrogen ions (H^+) to balance H

Balance charge by adding electrons to the side that is most positive so that the charges are the same

Do now:

Complete the following half equations:

Fe^{2+} → Fe^{3+} **+ e** **Oxidation:** **loss of e**

+2 **+3** **increase in ON**

NO_3^- **+ 2 H^+** **+ e** → NO_2 **+ H_2O**

Reduction: **gain of e**

+5 **+4** **decrease in ON**

Write down the oxidation number of Fe and N in each species

Putting half equations into a full equation

Write down oxidation and reduction half equations (IO_3^- and SO_2 example on board)

Multiply one or both equations so the number of electrons in each equation are the same

Write new half equations with the multipliers

Combine the two equations together

Cancel electrons and any other elements/compounds that are both products and reactants

Workbook pg 14, 15

Revision today

Today you need to:

Complete pg 10 on oxidation numbers

Complete pg 14 and 15 on half equations

When you have completed these work on the questions on pg 20, 21

Do now:

Complete the following half equations:

Reduction: gain of e decrease in ON

$Cr_2O_7^{2-}$ **$+ 14\ H^+ + 6\ e$** $\rightarrow$ **2** Cr^{3+} **$+ 7\ H_2O$**

+6 **+3**

Oxidation: loss of e increase in ON

$H_2O_2 \rightarrow O_2$ **$+ 2\ H^+ + 2\ e$**

-1 **0**

Write down the oxidation number of Cr and O in each species