Introduction to Oxidation and Reduction

OXIDATION and REDUCTION - REDOX REACTIONS

OXIDATION - definition and examples	REDUCTION - definition and examples
(a) The gain or addition of oxygen by an atom, molecule or ion eg ... **(1) S ==> SO_2** [burning sulphur - oxidised] **(2) CH_4 ==> CO_2 + H_2O** [burning methane to water and carbon dioxide, C and H gain O] **(3) NO ==> NO_2** [nitrogen monoxide oxidised to nitrogen dioxide] **(4) SO_3^{2-} ==> SO_4^{2-}** [oxidising the sulphite ion to the sulphate ion]	**(b) The loss or removal of oxygen from a compound etc.** eg ... **(1) CuO ==> Cu** [loss of oxygen from copper(II) oxide to form copper atoms] **(2) Fe_2O_3 ==> Fe** [iron(III) oxide reduced to iron] **(3) NO ==> N_2** [nitrogen monoxide reduced to nitrogen] **(4) SO_3 ==> SO_2** [sulphur trioxide reduced to sulphur dioxide]
(c) The loss or removal of electrons from an atom, ion or molecule eg **(1) Fe ==> Fe^{2+} + $2e^-$** [iron atom loses 2 electrons to form the iron(II) ion] **(2) Fe^{2+} ==> Fe^{3+} + e^-** [the iron(II) ion loses 1 electron to form the iron(III) ion] **(3) $2Cl^-$ ==> Cl_2 + $2e^-$** [the loss of electrons by chloride ions to form chlorine molecules]	**(d) The gain or addition of electrons by an atom, ion or molecule** eg ... **(1) Cu^{2+} + $2e^-$** ==> Cu [the copper(II) ion gains 2 electrons to form neutral copper atoms) **(2) Fe^{3+} + e^-** ==> Fe^{2+} [the iron(III) ion gains an electron and is reduced to the iron(II) ion] **(3) $2H^+$ + $2e^-$** ==> H_2 [hydrogen ions gain electrons to form neutral hydrogen molecules]
(e) An oxidising agent is the species that gives the oxygen or removes the electrons	**(f) A reducing agent is the species that removes the oxygen or acts as the electron donor**

REDOX REACTIONS - in a reaction overall, oxidation and reduction must go together

(g) Redox reaction analysis based on the oxygen definitions

- **(1) copper(II) oxide + hydrogen ==> copper + water**
  - **$CuO_{(s)}$ + $H_{2(g)}$ => $Cu_{(s)}$ + $H_2O_{(g)}$**
  - copper oxide reduced to copper, hydrogen is oxidised to water
  - hydrogen is the reducing agent (removes O from CuO)
  - copper oxide is the oxidising agent (donates O to hydrogen)
- **(2) iron(III) oxide + carbon monoxide ==> iron + carbon dioxide**
  - **$Fe_2O_{3(s)}$ + $3CO_{(g)}$ => $2Fe_{(l)}$ + $3CO_{2(g)}$**
  - the iron(III) oxide is reduced to iron, the carbon monoxide is oxidised to carbon dioxide
  - CO is the reducing agent (O remover from Fe_2O_3)
  - the Fe_2O_3 is the oxidising agent (O donator to CO)]
- **(3) nitrogen monoxide + carbon monoxide ==> nitrogen + carbon dioxide**
  - **$2NO_{(g)}$ + $2CO_{(g)}$ ==> $N_{2(g)}$ + $2CO_{2(g)}$**
  - nitrogen monoxide is reduced to nitrogen
  - carbon monoxide is oxidised to carbon dioxide
  - CO is the reducing agent and NO is the oxidising agent

- **(4)** iron(III) oxide + aluminium ==> aluminium oxide + iron (the thermit reaction)
  - $\mathbf{Fe_2O_{3(s)} + 2Al_{(s)} ==> Al_2O_{3(s)} + 2Fe_{(s)}}$
  - iron(III) oxide is reduced and is the oxidising agent
  - aluminium is oxidised and is the reducing agent

(h) Redox reaction analysis based on the electron definitions

- **(1) magnesium + iron(II) sulphate ==> magnesium sulphate + iron**
  - $\mathbf{Mg_{(s)} + FeSO_{4(aq)} => MgSO_{4(aq)} + Fe_{(s)}}$
  - this is the 'ordinary molecular' **equation for a typical metal displacement reaction**, but this does not really show what happens in terms of atoms, ions and electrons, so we use **ionic equations** like the one shown below.
  - The sulphate ion $\mathbf{SO_4}^{2-}{}_{(aq)}$ is called a **spectator ion**, because it doesn't change in the reaction and can be omitted from the ionic equation. No electrons show up in the full equations because electrons lost by x = electrons gained by y!!
  - $\mathbf{Mg_{(s)} + Fe^{2+}{}_{(aq)} ==> Mg^{2+}{}_{(aq)} + Fe_{(s)}}$
  - the magnesium atom loses 2 electrons (oxidation) to form the magnesium ion, the iron(II) ion gains 2 electrons (reduced) to form iron atoms.
  - Mg is the reducing agent (electron donor) and the Fe^{2+} is the oxidising agent (electron remover or acceptor)
  - **Displacement reactions** involving metals and metal ions **are electron transfer reactions.**
- **(2) zinc + hydrochloric acid ==> zinc chloride + hydrogen**
  - $\mathbf{Zn_{(s)} + 2HCl_{(aq)} => ZnCl_{2(aq)} + H_{2(g)}}$
  - the chloride ion Cl^- is the spectator ion
  - $\mathbf{Zn_{(s)} + 2H^+{}_{(aq)} ==> Zn^{2+}{}_{(aq)} + H_{2(g)}}$
  - Zinc atoms are oxidised to zinc ions by electron loss, so zinc is the reducing agent (electron donor)
  - hydrogen ions are the oxidising agent (gaining the electrons) and are reduced to form hydrogen molecules
- **(3) copper + silver nitrate ==> silver + copper(II) nitrate**
  - $\mathbf{Cu_{(s)} + 2AgNO_{3(aq)} ==> 2Ag + Cu(NO_3)_{2(aq)}}$
  - the nitrate ion NO_3^- is the spectator ion
  - $\mathbf{Cu_{(s)} + 2Ag^+{}_{(aq)} ==> 2Ag_{(s)} + Cu^{2+}{}_{(aq)}}$
  - copper atoms are oxidised by the silver ion by electron loss
  - electrons are transferred from the copper atoms to the silver ions, which are reduced
  - the silver ions are the oxidising agent and the copper atoms are the reducing agent
- **(4) iron(II) chloride + chlorine ==> iron(III) chloride**
- **(5) halogen** (more reactive) **+ halide salt** (of less reactive halogen) **==> halide salt** (of more reactive halogen) **+ halogen** (less reactive)
  - $\mathbf{X}_{2(aq)} + 2K\mathbf{Y}_{(aq)} ==> 2K\mathbf{X}_{(aq)} + \mathbf{Y}_{2(aq)}$
  - $\mathbf{X}_{2(aq)} + 2\mathbf{Y}^-{}_{(aq)} ==> 2\mathbf{X}^-{}_{(aq)} + \mathbf{Y}_{2(aq)}$
  - where halogen **X** is more reactive than halogen **Y**, F > Cl > Br > I)
  - **X** is the **oxidising agent** (electron acceptor)
  - K**Y** is the **reducing agent** (electron donor)
- **(6) Electrode reactions in electrolysis are electron transfer redox changes**
  - **at the negative cathode** positive ions are attracted:
    - metal ions are reduced to the metal by electron gain:
    - $\mathbf{M^{n+} + n\ e^- ==> M}$
    - **n** = the numerical charge of the ion and the number of electrons transferred
    - or $\mathbf{2H^+{}_{(aq)} + 2e^- ==> H_{2(g)}}$ (for the discharge of hydrogen)

- **at the positive anode** negative ions are attracted:
  - negative non-metal ions are oxidised by electron loss eg
  - for oxide ions: $\mathbf{2O^{2-} - 4e^- ==> O_2}$ or $\mathbf{2O^{2-} ==> O_2 + 4e^-}$
  - for hydroxide ion: $\mathbf{4OH^- - 4e^- ==> O_2 + 2H_2O}$ or $\mathbf{4OH^- ==> O_2 + 2H_2O + 4e^-}$
  - for halide ions (X = F, Cl, Br, I): $\mathbf{2X^- - 2e^- ==> X_2}$ or $\mathbf{2X^- ==> X_2 + 2e^-}$

Miscellaneous Redox Notes

- **Redox changes can often be observed as significant colour changes** eg
  - **iron + copper(II) sulphate ==> iron(II) sulphate + iron**
    - $\mathbf{Mg_{(s)} + FeSO_{4(aq)} => MgSO_{4(aq)} + Fe_{(s)}}$
    - $\mathbf{Mg_{(s)} + Fe^{2+}_{(aq)} ==> Mg^{2+}_{(aq)} + Fe_{(s)}}$
    - Sulphate, $SO_4^{2-}{}_{(aq)}$, is colourless BUT **a blue to pale green colour change** is observed in the solution as the **blue copper(II) ion** is replaced by the **pale green iron(II) ion**.
  - **Potassium manganate(VII)** is a powerful oxidising agent and an **intense purple colour** in water due to the MnO_4^- ion. In acidified solution it changes to an almost colourless* manganese(II) ion, Mn^{2+} when it oxidises something (* which actually is a very pale pink transition metal ion).
  - **Potassium dichromate(VI)** is another strong oxidising agent and is **orange due to the dichromate(VI) ion, $Cr_2O_7^{2-}$ ion**. When it oxidises something it changes to the **green chromium(III) ion, Cr^{3+}**.
  - **Potassium iodide** is a colourless salt dissolving in water to form a colourless solution. If it is oxidised eg with chlorine a **yellow==>orange==>brown colour** develops as iodine is formed from the colourless iodide ion.
- **The use of Roman Numerals in names**:
  - **This indicates** what is called **the oxidation state of an atom** in a molecule or ion.
  - It is easy to follow for simple metal ions because it equals the charge on the ion
    - eg the **oxidation state** of copper in the copper(II) ion is referred to as **+2**
      - the more electrons removed from the atom or ion by oxidation, the higher its oxidation state eg $Fe^{2+} - e^- ==> Fe^{3+}$, gives iron the oxidation state of +3 in the iron(III) ion (via a suitable oxidising agent).
    - **but** for more complex ions things are not so simple.
      - in manganate(VII) ion, the Mn is in the +7 oxidation state
      - in dichromate(VI) ion, the Cr is in the +6 oxidation state
  - This topic is dealt with at AS-A2 advanced level chemistry (there is an **introduction** on another web page) but not for KS4-GCSE!

Oxidation Number

Oxidation numbers are a useful tool for determining whether a substance has been oxidized or reduced. An element that undergoes a change in oxidation number in the course of a reaction has been oxidized or reduced. Let's learn how to assign oxidation numbers.

Assigning Oxidation Numbers

1. At atom in its elemental state has an oxidation number of 0.

Na H_2 Cl_2 S Xe

Each atom in these elements has an oxidation number of 0.

2. An atom in a monoatomic ion has an oxidation number identical to its charge.

Na^+ Ba^{2+} Al^{3+} Br^- S^{2-}
+1 +2 +3 –1 –2

The oxidation number is equal to the charge on the monoatomic ion.

3. An atom in a polyatomic ion or a molecular compound usually has the same oxidation number it would have if it were in a monoatomic ion.

a. Elements to the left on the periodic table are "cationlike" and have positive oxidation numbers.

b. Elements to the right on the periodic table are "anionlike" and have negative oxidation numbers.

Consider NH_3.

N has an oxidation number of –3; each H has an oxidation number of +1.

c. Hydrogen

has a +1 oxidation number when bonded to nonmetals, and

has a –1 oxidation number when bonded to a metal.

NaH (H –1 oxidation number)

H_2O (H +1 oxidation number)

d. Oxygen

often has a –2 oxidation number, but

can have a –1 oxidation number in the peroxide ion, O_2^{2-}.

H_2O (O –2 oxidation number)

HOOH (O –1 oxidation number)

e. Halogens

usually have an oxidation number of –1,

Unless bonded to oxygen, when they have a positive oxidation number.

HCl (Cl –1 oxidation number)

HOCl (Cl +1 oxidation number)

4. The sum of the oxidation numbers is 0 for a neutral compound and is equal to the net charge for a polyatomic ion.

Oxidizing and Reducing agents

Oxidation and reduction always occur together. Whenever one atom loses electrons (is oxidized), another atom must gain those electrons (be reduced). The reactants can be classified as either a **reducing agent** or an **oxidizing agent**.

Reducing agent

- causes reduction
- loses one or more electrons
- is oxidized
- oxidation number of atom increases

Oxidizing agent

- causes oxidation
- gains one or more electrons
- is reduced
- oxidation number of atom decreases

The Activity Series of the Elements

TABLE 4.3 A Partial Activity Series of the Elements

	Oxidation Reaction	
Strongly reducing	$Li \rightarrow Li^{+} + e^{-}$	These elements react rapidly with aqueous H^{+} ions (acid) or with liquid H_2O to release H_2 gas.
	$K \rightarrow K^{+} + e^{-}$	
	$Ba \rightarrow Ba^{2+} + 2e^{-}$	
	$Ca \rightarrow Ca^{2+} + 2e^{-}$	
	$Na \rightarrow Na^{+} + e^{-}$	
	$Mg \rightarrow Mg^{2+} + 2e^{-}$	These elements react with aqueous H^{+} ions or with steam to release H_2 gas.
	$Al \rightarrow Al^{3+} + 3e^{-}$	
	$Mn \rightarrow Mn^{2+} + 2e^{-}$	
	$Zn \rightarrow Zn^{2+} + 2e^{-}$	
	$Cr \rightarrow Cr^{3+} + 3e^{-}$	
	$Fe \rightarrow Fe^{2+} + 2e^{-}$	
	$Co \rightarrow Co^{2+} + 2e^{-}$	These elements react with aqueous H^{+} ions to release H_2 gas.
	$Ni \rightarrow Ni^{2+} + 2e^{-}$	
	$Sn \rightarrow Sn^{2+} + 2e^{-}$	
	$H_2 \rightarrow 2H^{+} + 2e^{-}$	
	$Cu \rightarrow Cu^{2+} + 2e^{-}$	These elements do not react with aqueous H^{+} ions to release H_2.
	$Ag \rightarrow Ag^{+} + e^{-}$	
	$Hg \rightarrow Hg^{2+} + 2e^{-}$	
Weakly reducing	$Pt \rightarrow Pt^{2+} + 2e^{-}$	
	$Au \rightarrow Au^{3+} + 3e^{-}$	

The elements at the top of the table readily give up electrons and are stronger reducing agents. The elements at the bottom give up electrons less readily and are weaker reducing agents. Any element higher in the activity series will react with the ion of any element lower in the activity series.

Some Applications of Redox Reactions

A vast number of redox reactions occur in industrial and biological processes. A few are summarized here.

1. Combustion is the burning of fuel by oxidation with oxygen in air. Fuels include natural gas, wood, paper, and other organic substances composed of carbon and hydrogen. Some metals also burn in air.

$$CH_4(q) + 2\,O_2(g)\quad CO_2(g) + 2\,H_2O(l)$$

$$2Mg(s) + 2\,O_2(g)\quad 2\,MgO(s)$$

2. Bleaching is the use of redox reactions to decolorize or lighten colored materials. Oxidizing agents used in bleaching include hydrogen peroxide (H_2O_2) and sodium hypochlorite (NaClO).

3. Batteries are all based on redox reactions.

4. Metallurgy is the science of extracting and purifying metals from their ores.

5. Undesirable oxidation reactions are termed corrosion. The rusting of iron in moist air is a familiar process with enormous economic impact.

$$4\ Fe(s) + 3\ O_2(g)\quad Fe_2O_3H_2O(s)$$

6. Respiration is the process of breathing and using oxygen for the many biological redox reactions that occur in living organisms.

$$C_6H_{12}O_6(s) + 6\ O_2(g)\quad 6\ CO_2(g) + 6\ H_2O(l) + energy$$

<u>DONE</u>