1.7 REDOX

Oxidation and Reduction:

- Oxidation and reduction reactions can be identified by looking at the reaction in terms of electron transfer:
- Our understanding of oxidation and reduction was limited to reactions involving oxygen (and hydrogen):

$$Mg \quad + \quad {}^{1}/_{2}O2 \quad \rightarrow \quad MgO$$

- Convert these to ionic and half equations and you can see clearly how the electrons are transferred:

Magnesium:

$$Mg \quad \rightarrow \quad Mg^{2+} \quad + \quad 2e^{-}$$

- Magnesium has been **oxidised** as it has gained oxygen.
- But it has also **lost 2 electrons**:

Oxygen:

$${}^{1}/_{2}O_2 \quad + \quad 2e^{-} \quad \rightarrow \quad O^{2-}$$

- Oxygen has been **reduced** as it has lost oxygen.
- But each oxygen has also **gained 2 electrons**:

New definitions:

***O**xidation*
***I**s*
***L**oss of electrons*

***R**eduction*
***I**s*
***G**ain of electrons*

- Oxidation and reduction must occur simultaneously as all reactions involve a movement of electrons.
- These reactions are given the shorthand term of **REDOX** reactions. As they involve

REDuction and **OX**idation

Example: Identify what has been oxidised and reduced:

1) **Overall chemical equation:**

$$Na_{(s)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + {}^{1}/_{2}H_{2(g)}$$

2) **Covert to ionic equation and identify spectator ions:**

$$Na_{(s)} + H^{+}_{(aq)} + Cl^{-}_{(aq)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)} + {}^{1}/_{2}H_{2(g)}$$

3) **Remove spectator ions and identify what will be in each ½ equation:**

$$Na_{(s)} + H^{+}_{(aq)} \rightarrow Na^{+}_{(aq)} + {}^{1}/_{2}H_{2(g)}$$

4) **Write the half equation and determine REDOX using OILRIG:**

$Na_{(s)} \rightarrow Na^{+}_{(aq)} + e^{-}$ **Oxidation – lost electron**

$H^{+}_{(aq)} + e^{-} \rightarrow {}^{1}/_{2}H_{2(g)}$ **Reduction – gained electron**

Example questions:

1) Look at the following reactions and decide whether they are oxidation or reduction reactions:

   a. $Ca \rightarrow Ca^{2+} + 2e^{-}$

   b. $Cl_2 + 2e^{-} \rightarrow 2Cl^{-}$

   c. $2Br^{-} \rightarrow Br_2 + 2e^{-}$

2) Convert the following reaction into half equations, then identify the species that has been oxidised and which species has been reduced:

   Overall reaction:

   $$2KI_{(aq)} + Cl_{2(aq)} \rightarrow 2KCl_{(aq)} + I_{2(aq)}$$

   Ionic equation:

   Half equations and state which is oxidation and which is reduction:

Oxidation states (numbers)

- It is used to describe the number of electrons used to bond with another atom.
- It is also used for combining powers of atoms.
- It is a type of 'book keeping' for electrons.
- It is a number describing the movement of electrons and is found by the application of certain rules:

Rules for assigning Oxidation States

1) Oxidation state of an element = 0

Na Ar O_2 H_2 All have Ox state: 0

2) The oxidation state of a monatomic ion = the charge on the ion.

Na^{1+} = +1 Cl^{1-} = -1 Al^{3+} = 3+

3) The sum of the oxidation states in a neutral compound = 0

NaCl		Li_2O		$AlCl_3$	
+1	-1	+1	-2	+3	-1
		+1			-1
					-1
0		**0**		**0**	

4) The sum of the oxidation states in a polyatomic ion (SO_4^{2-}) = charge on the polyatomic ion

OH^{1-}		NH_4^{1+}		SO_4^{2-}	
-2	+1	-3	+1	+6	-2
			+1		-2
			+1		-2
			+1		-2
-1		**+1**		**-2**	

5) Some elements have fixed oxidation states in a compound:

Group 1	Group 2	Group 3	Hydrogen	Oxygen	Fluorine	Group 7
+1	**+2**	**+3**	**+1**	**-2**	**-1**	**-1**
			Except in metal hydrides, MH H = -1	Except in peroxides, H_2O_2 O = -1		Except when combined with oxygen as 'ates' Variable

Oxidation numbers in chemical names:

- Some elements form compounds where they could have a different charge / oxidation states / number.

Transition metals:

Compound	Name	Element with different oxidation state	Oxidation state of that element
$FeCl_2$	Iron (II) chloride	Fe	+2
$FeCl_3$	Iron (III) chloride	Fe	+3

- Roman numerals indicate the oxidation number of the element before it.
- This also occurs with oxyanions:

Oxo – anions: 'ate' molecules

Oxo - anion	Name	Element with different oxidation state	Oxidation state of that element
NO_2^-	Nitrate (III)	N	+3
NO_3^-	Nitrate (V)	N	+5

- These are negative molecules that contain an oxygen atom.

Oxo – acids:

Oxo - acid	Name	Element with different oxidation state	Oxidation state of that element
HNO_2	Nitric (III) acid	N	+3
HNO_3	Nitric (V) acid	N	+5

- These are acids containing oxygen.

Working out formula from oxidation states:

Example:
Give the formula for potassium chlorate (III)

- Potassium is in Gp 1 so its oxidation state = +1
- Chlor**ate** means that chlorine is directly bonded with **oxygen.**
- Oxygen always has an oxidation state = -2
- Chlorate **(III)** tells you the oxidation state of chlorine = +3

K	**Cl**	**O**
+1	+3	-2

- As the compound is neutral, the sum of the oxidation states must add up to = 0
- So 2 x 2- required, therefore 2 x O

	K	**Cl**	**O** $_{\times 2}$	
	+1	+3	-2	
			-2	
Totals:	**+4**		**-4**	**which = 0**

Have a go at the following:

Manganese (IV) oxide

Sodium sulphate (VI)

Sodium sulphate (IV)

Iodate (V) with a 1- charge

Working out oxidation states from formula:
Example:
Work out the oxidation state of manganese in $KMnO_4$

- **K** is in Gp 1 so has an oxidation state = +1
- **O** has a fixed oxidation state = -2 and we have 4 of those
- The compound is neutral overall:

	K	Mn	O_4	
	+1	?	-2	
			-2	
			-2	
			-2	
Totals:	**+1**	**?**	**-8**	**which = 0**

- **Mn** must therefore have an oxidation state of **+7** in order for the sum of the oxidation states to = 0

Have a go at the following: Find the oxidation state of the element in **bold**

VO_3 Mg**S**O_3 Na**Cl**O_3 Na_2**Cr**$_2O_7$

Oxidation states and REDOX reactions

- As oxidation states show the movement of electrons in a reaction it is possible to use these to identify what has been oxidised and what has been reduced:

$Mg_{(s)}$	+	$2HCl_{(aq)}$	→	$MgCl_{2(aq)}$	+	$H_{2(g)}$
0		+1 -1		+2 -1 -1		0

REDOX:

- In this reaction Mg has lost 2e **Oxidation**
- In this reaction each H^+ has gained 1e **Reduction**

Oxidation states:

- Mg oxidation state: **0 → +2** **Oxidation state has increased → Oxidation**
- H^+ oxidation state: **+1 → 0** **Oxidation state has decreased → Reduction**

Summary:

***<u>Oxidation</u>* is an increase in oxidation state**

***<u>Reduction</u>* is a decrease in oxidation state**

(Its oxidation state REDUCES)

Oxidation	Oxidation state	Reduction
	+7	
	+6	
	+5	
	+4	
	+3	
	+2	
	+1	
	0	
	-1	
	-2	
	-3	
	-4	

Other examples:

- Adding the oxidation state makes it even easier:

$2Fe_{(s)}$	+	$3Cl_{2(g)}$	→	$2FeCl_{3(s)}$
0		**0**		**+3 -1**

- Fe oxidation state has been increased from **0 → +3** **Oxidised** (electrons lost)
- Cl oxidation state has been reduced from **0 → -1** **Reduced** (electrons gained)

Have a go at these:

1)

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

__ __ __ __

- ___ oxidation state has gone from ___ → ___ therefore has been _______________
- ___ oxidation state has gone from ___ → ___ therefore has been _______________

2)

$$Mg_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + H_{2(g)}$$

__ __ __ __ __ __ __ __

- ___ oxidation state has gone from ___ → ___ therefore has been _______________
- ___ oxidation state has gone from ___ → ___ therefore has been _______________

3)

$$Li_{(aq)} + HNO_{3(aq)} \rightarrow LiNO_{3(aq)} + H_{2(g)}$$

__ __ __ __ __ __ __ __

- ___ oxidation state has gone from ___ → ___ therefore has been _______________
- ___ oxidation state has gone from ___ → ___ therefore has been _______________

Oxidising and reducing agents:

$$\underset{0}{2Fe_{(s)}} + \underset{0}{3Cl_{2(g)}} \rightarrow \underset{+3\ \ -1}{2FeCl_{3(s)}}$$

- Fe oxidation state has been increased from **0 → +3** **Oxidised** (electrons lost)
- Cl oxidation state has been reduced from **0 → -1** **Reduced** (electrons gained)

So:

- **Fe** can only lose its electrons if there is a species to accept these electrons
- As **Cl** accepted the electrons from iron for it to be oxidised we say that **chlorine** is the **oxidising agent**
- **Cl** can only gain electrons if there is a species to lose these electrons
- As **Fe** gave the electrons to chlorine for it to be reduced we say that **iron** is the **reducing agent:**

Oxidation – Reducing agents
Is
Loss of electrons

Reduction – Oxidising agents
Is
Gain of electrons

Basically:

If it is oxidised it is a reducing agent

If it is reduced it is an oxidising agent

Have a go at these:

1)

$2Ca_{(s)}$ + $O_{2(g)}$ → $2CaO_{(s)}$

__   __   __ __

- ___ oxidation state has gone from ___ → ___ therefore has been _______________

  This makes it an _________________ agent.

- ___ oxidation state has gone from ___ → ___ therefore has been _______________

  This makes it a _________________ agent.

2)

$Sr_{(s)}$ + $H_2SO_{4(aq)}$ → $SrSO_{4(aq)}$ + $H_{2(g)}$

__   __ __ __   __ __ __   __

- ___ oxidation state has gone from ___ → ___ therefore has been _______________

  This makes it an _________________ agent.

- ___ oxidation state has gone from ___ → ___ therefore has been _______________

  This makes it a _________________ agent.

3)

$Na_{(s)}$ + $HNO_{3(aq)}$ → $NaNO_{3(aq)}$ + $H_{2(l)}$

__   __ __ __   __ __ __   __

- ___ oxidation state has gone from ___ → ___ therefore has been _______________

  This makes it an _________________ agent.

- ___ oxidation state has gone from ___ → ___ therefore has been _______________

  This makes it a _________________ agent.

Constructing simple half equations:

- **Write species for the equation and make sure the atoms / ions and the charges balance by adding electrons on the appropriate side:**

Example:
The oxidation of Fe to Fe^{3+}:

$$\mathbf{Fe \rightarrow Fe^{3+} + 3e^{-}}$$

The reduction of chlorine to chloride:

$$\mathbf{\tfrac{1}{2}Cl_2 + e^{-} \rightarrow Cl^{-}}$$

Constructing complex half equations:
Example:
The reduction of MnO_4^{-} to Mn^{2+}

Step 1: Write the reactants and products making sure the REDOX element balances:

$$MnO_4^{-} \rightarrow Mn^{2+}$$

Step 2: Balance the oxygen's by adding water to the opposite side:

$$MnO_4^{-} \rightarrow Mn^{2+} + 4H_2O$$

Step 3: Balance the hydrogen's by adding hydrogen ions to the opposite side:

$$MnO_4^{-} + 8H^{+} \rightarrow Mn^{2+} + 4H_2O$$

Step 4: Balance the charges using electrons:

$$MnO_4^{-} + 8H^{+} + 5e^{-} \rightarrow Mn^{2+} + 4H_2O$$

Have a go at these:

1) **Bromide to bromine**

2) **Nitric acid to nitrogen dioxide**

3) **Sulphuric acid to Sulphur dioxide**

4) **Hydrogen peroxide to oxygen**

Constructing redox equations from half equations

- This is done by balancing the numbers of electrons **lost** by one half equation with those being **gained** by another half equation.
- Any species the same on both sides can then be cancelled.

Example:

$$Fe \rightarrow Fe^{3+} + 3e^-$$

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

Step 1: Balance the half equations using the electrons:

$$2Fe \rightarrow 2Fe^{3+} + 6e^- \quad x2$$

$$3Cu^{2+} + 6e^- \rightarrow 3Cu \quad x3$$

Step 2: Add the half equations together and cancel out the electrons:

$$2Fe \rightarrow 2Fe^{3+} + 6e^-$$

$$3Cu^{2+} + 6e^- \rightarrow 3Cu$$

$$2Fe + 3Cu^{2+} + \cancel{6e^-} \rightarrow 3Cu + 2Fe^{3+} + \cancel{6e^-}$$

This gives:

$$2Fe + 3Cu^{2+} \rightarrow 3Cu + 2Fe^{3+}$$

Have a go at these:

1) $$Al \rightarrow Al^{3+} + 3e^-$$

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

2) $$Cu \rightarrow Cu^+ + e^-$$

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

Constructing REDOX equations from text:

- Identify the REDOX species and write balanced half equations.
- Balance the 2 half equations by balancing the numbers of electrons **lost** by one half equation with those being **gained** by another half equation.
- Any species the same on both sides can then be cancelled.

Example:

Sulphuric acid reacts with **hydrogen peroxide** forming **sulphur dioxide** and **oxygen**

Step 1: Identify and construct the 2 redox half equations balancing O with H_2O and H with H^+:

$$H_2SO_4 + 2H^+ + 2e^- \rightarrow SO_2 + 2H_2O$$

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

Step 2: Balance the electrons:

$$H_2SO_4 + 2H^+ + 2e^- \rightarrow SO_2 + 2H_2O$$

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

Step 3: Add the half equations together and cancel out the electrons / any other species on both sides:

$$H_2SO_4 + 2H^+ + 2e^- \rightarrow SO_2 + 2H_2O$$

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

$$H_2SO_4 + \cancel{2H^+} + \cancel{2e^-} + H_2O_2 \rightarrow SO_2 + 2H_2O + O_2 + \cancel{2H^+} + \cancel{2e^-}$$

To give:

$$\mathbf{H_2SO_4 + H_2O_2 \rightarrow SO_2 + 2H_2O + O_2}$$