
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):



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

Magnesium:



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

Oxygen:



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

New definitions:

Oxidation
Is
Loss of electrons

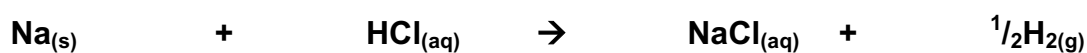
Reduction
Is
Gain of electrons

- 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:



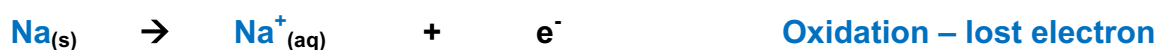
2) Convert to ionic equation and identify spectator ions:



3) Remove spectator ions and identify what will be in each $\frac{1}{2}$ equation:

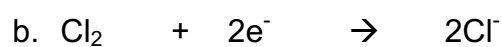


4) Write the half equation and determine REDOX using OILRIG:



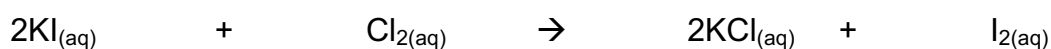
Example questions:

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



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

Overall reaction:



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 O₂ H₂ All have Ox state: 0

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

Na¹⁺ = +1 Cl¹⁻ = -1 Al³⁺ = 3+

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

NaCl	Li ₂ O	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 (SO₄²⁻) = 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 ₂ O 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
- Chlorate 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_{x2}
+1	+3	-2

Totals:	+4	-2	-2	which = 0
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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	O₄
+1	?	-2
		-2
		-2
		-2

Totals: +1 ? -8 which = 0

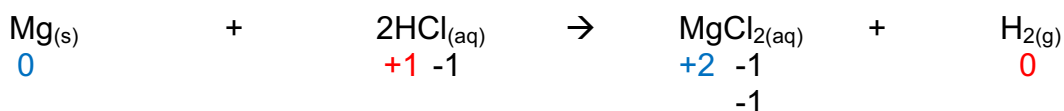
- **Mn** must therefore have an oxidation state of **+7** in order for the sum of the oxidation states to = 0

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



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:



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:

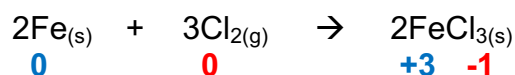
Oxidation is an increase in oxidation state

Reduction 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:

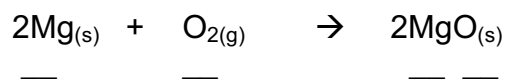
- Adding the oxidation state makes it even easier:



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

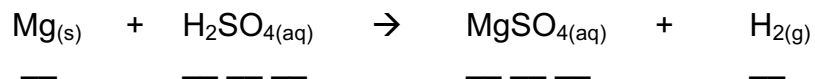
Have a go at these:

1)



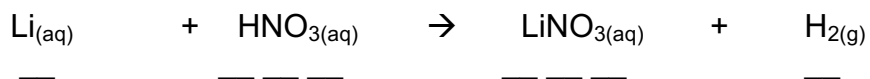
- ___ oxidation state has gone from ___ \rightarrow ___ therefore has been _____
- ___ oxidation state has gone from ___ \rightarrow ___ therefore has been _____

2)



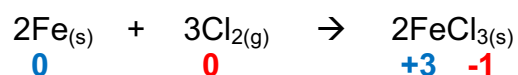
- ___ oxidation state has gone from ___ \rightarrow ___ therefore has been _____
- ___ oxidation state has gone from ___ \rightarrow ___ therefore has been _____

3)



- ___ oxidation state has gone from ___ \rightarrow ___ therefore has been _____
- ___ oxidation state has gone from ___ \rightarrow ___ therefore has been _____

Oxidising and reducing agents:



- 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

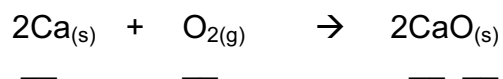
Basically:

If it is oxidised it is a reducing agent

If it is reduced it is an oxidising agent

Have a go at these:

1)



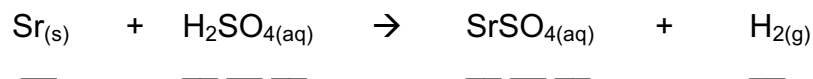
- ___ oxidation state has gone from ___ \rightarrow ___ therefore has been _____

This makes it an _____ agent.

- ___ oxidation state has gone from ___ \rightarrow ___ therefore has been _____

This makes it a _____ agent.

2)



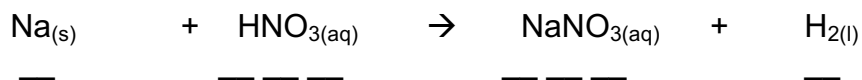
- ___ oxidation state has gone from ___ \rightarrow ___ therefore has been _____

This makes it an _____ agent.

- ___ oxidation state has gone from ___ \rightarrow ___ therefore has been _____

This makes it a _____ agent.

3)



- ___ oxidation state has gone from ___ \rightarrow ___ therefore has been _____

This makes it an _____ agent.

- ___ oxidation state has gone from ___ \rightarrow ___ 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 Fe³⁺:



The reduction of chlorine to chloride:



Constructing complex half equations:

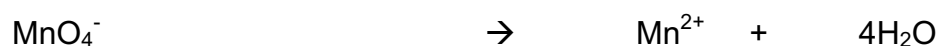
Example:

The reduction of MnO₄⁻ to Mn²⁺

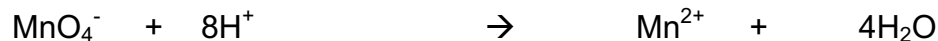
Step 1: Write the reactants and products making sure the REDOX element balances:



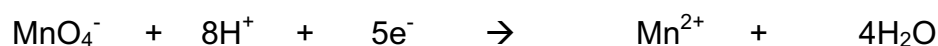
Step 2: Balance the oxygen's by adding water to the opposite side:



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



Step 4: Balance the charges using electrons:



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.

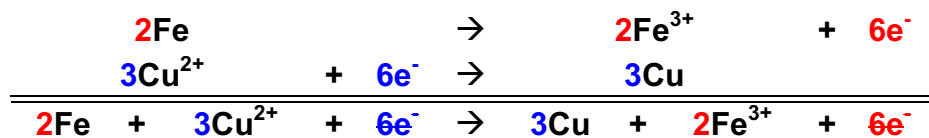
Example:



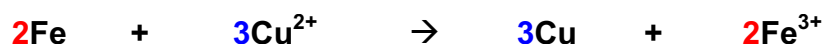
Step 1: Balance the half equations using the electrons:



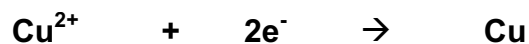
Step 2: Add the half equations together and cancel out the electrons:



This gives:



Have a go at these:



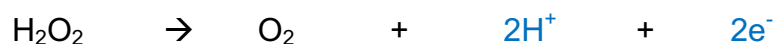
Constructing REDOX equations from text:

- 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.

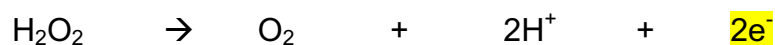
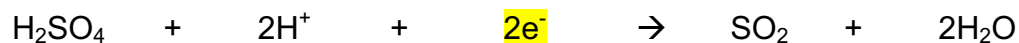
Example:

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

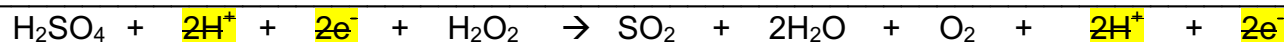
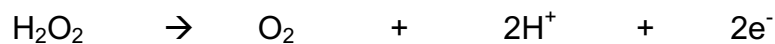
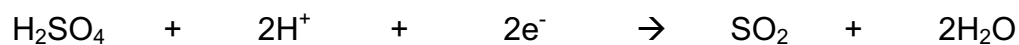
Step 1: Identify and construct the 2 redox half equations balancing O with H₂O and H with H⁺:



Step 2: Balance the electrons:



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



To give:

