CAIE IGCSE Chemistry

6.4 Redox

Notes

Use a Roman numeral to indicate the oxidation number of an element in a compound

- The oxidation number of an element is the charge of an element/ion in a compound which relates to the electrons gained or lost by the element/ion during the formation of the compound.
- The oxidation number of an element is represented by a Roman numeral placed after the element name inside brackets
- E.g. Iron (II) oxide
  The (II) shows that the iron ion in the compound has an oxidation number of +2.
  E.g. Copper (I) chloride
  The (I) shows that the copper ion in the compound has an oxidation number of +1

Define redox reaction

Redox reactions are reactions where both oxidation and reduction occur simultaneously

Define oxidation and reduction

- Oxidation is the gain of oxygen
- Reduction is the loss of oxygen
- E.g. H_2 + CuO -> Cu + H_2O
  The hydrogen has gained oxygen so has been oxidised
  The copper has lost oxygen so has been reduced

Identify redox reactions as reactions involving gain and loss of oxygen

- Redox reactions involve both oxidation and reduction so there will be the gain and loss of oxygen in the reaction
- E.g. H_2 + ZnO -> Zn + H_2O
  The hydrogen has gained oxygen so has been oxidised
  The zinc has lost oxygen so has been reduced

Identify oxidation and reduction in redox reactions

- To identify what is being oxidised and reduced in redox reactions, look out for whether any elements have gained or lost oxygen in forming their products:

  E.g. $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
  Iron (III) oxide + Carbon monoxide -> Iron + Carbon dioxide
  - Iron oxide -> Iron means iron oxide has lost oxygen so iron oxide has been reduced
  - Carbon monoxide -> Carbon dioxide means more oxygen has been gained to form carbon dioxide so the carbon monoxide has been oxidised

(Extended only) Define oxidation in terms of…

Oxidation is the loss of electrons in a substance, increasing its oxidation number

(Extended only) Define reduction in terms of:

Reduction is the gain of electrons in a substance, decreasing its oxidation number

(Extended only) Identify redox reactions as reactions involving gain and loss of electrons

- To identify redox reactions, it is helpful to split the overall symbol equation into its ionic half equations (remove the spectator ions)
- E.g. Overall equation: $Cl_2 + 2KI \rightarrow 2KCl + I_2$
  Half equations:
  $Cl_2 + 2e^- \rightarrow 2Cl^-$ Chlorine has gained electrons (reduced)
  $2I^- \rightarrow I_2 + 2e^-$ Iodide ions have lost electrons (oxidised)

(Extended only) Identify redox reactions by changes in oxidation number using:

To identify the oxidation number, there are a few rules to help:

(a) The oxidation number of elements in their uncombined state is zero
E.g. The oxidation number of Zn is 0
The oxidation number of Cl_2 is 0
The oxidation number H_2 is 0

(b) The oxidation number of a monatomic ion is the same as the charge on the ion
E.g. The oxidation number of Zn^{+2} is +2
The oxidation number of Cl^- is –1
The oxidation number H^+ is +1

(c) The sum of the oxidation numbers in a compound is zero
E.g. The sum of the oxidation numbers in $ZnCl_2$ is 0
The oxidation number of Zn^{+2} is +2 and the oxidation number of Cl^- is –1 so:

$ZnCl_2$

$(+2) + 2(-1) = 0$

E.g. The sum of the oxidation numbers in HBr is 0
The oxidation number of H^{+1} is +1 and the oxidation number of Br^- is –1 so:

HBr

$(+1) + (-1) = 0$

(d) The sum of the oxidation numbers in an ion is equal to the charge on the ion

- E.g. The oxidation number of SO_4^{-2} is -2
  The oxidation number of S is +6 and the oxidation number of O is –2 so:

SO_4^{-2}

$(+6) + 4(-2) = -2$

Identifying redox reactions using oxidation numbers

- To write balanced redox reactions and confirm that both reduction and oxidation has occurred, oxidation numbers can be used to keep track of the electron changes

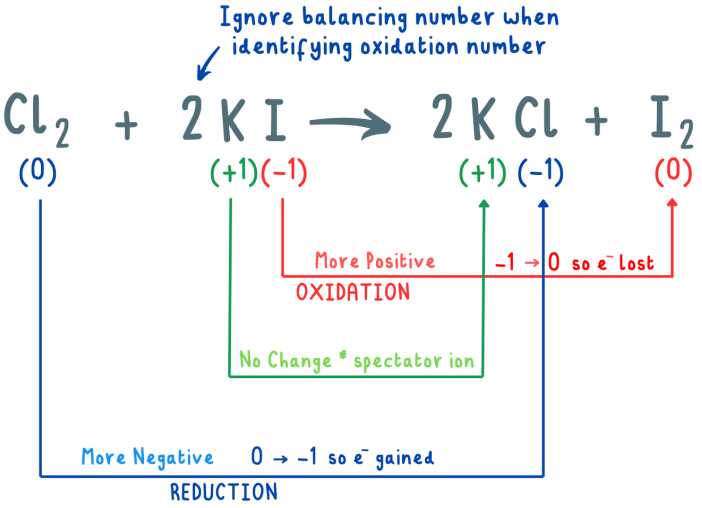
- If the oxidation number of an element/ion has become more negative/decreased = electrons have been gained = REDUCTION
- If the oxidation number of an element/ion has become more positive/increased = electrons have been lost = OXIDATION

- E.g. Identify which species has been oxidised and reduced in the following redox equation using their oxidation numbers:

$Cl_2 + 2KI \rightarrow 2KCl + I_2$

1. Write the oxidation numbers of each element beneath the equation

2. Identify which species have different oxidation numbers from the reactants to the products (ignoring spectator ions)
3. If the oxidation number of an element/ion has become more negative/decreased= electrons have been gained= REDUCTION
   If the oxidation number of an element/ion has become more positive/increased= electrons have been lost = OXIDATION
4. State which species has been oxidised and reduced, stating the change in oxidation number
   Chlorine has been reduced from 0 in Cl_2 to -1 in KCl (Cl^- ions)
   Iodide ions have been oxidised from -1 in KI to 0 in I_2



(Extended only) Identify redox reactions by the colour changes involved when using acidified aqueous potassium manganate(VII) or aqueous potassium iodide

- Redox reactions involve both reduction and oxidation occurring at the same time, so there is oxidising agent and reducing agent present.
- To identify the oxidising agent and reducing agent in a redox reaction, potassium iodide and potassium manganate can be used:

- Acidified aqueous potassium manganate(VII)
  - Tests for reducing agents
  - Potassium manganate(VII) has Mn^{7+} ions which are purple
  - When the Mn^{7+} ions are reduced to Mn^{2+} ions by a reducing agent a colour change occurs from purple to colourless.

- Potassium iodide
  - Tests for oxidising agents
  - Potassium iodide has I^- ions
  - When the I^- ions are oxidised to I_2 by an oxidising agent a colour change in the solution occurs from colourless to brown.
  - Starch solution can also be added, turning the iodine solution blue-black

(Extended only) Define an oxidising agent

- An oxidising agent is a species which brings about oxidation by gaining electrons from other elements/ions. The oxidising agent is itself reduced.
    - Usually a non-metal or a positive ion
    - E.g. $Br_2 + 2e^- \rightarrow 2Br^-$
        - Bromine has gained electrons so has been reduced

(Extended only) Define a reducing agent

- A reduction agent is a species which brings about reduction by losing electrons to another element/ion. The reducing agent is itself oxidised.
    - Usually a metal or a negative ion
    - E.g. $K \rightarrow K^+ + e^-$
        - Potassium has lost an electron so has been oxidised

(Extended only) Identify oxidising agents and reducing agents in redox reactions

- Oxidising agents and reducing agents can be identified in redox reactions by the ionic half equation method or by using oxidation numbers:

E.g. Overall equation: $Cl_2 + 2KI \rightarrow 2KCl + I_2$

Half equations:

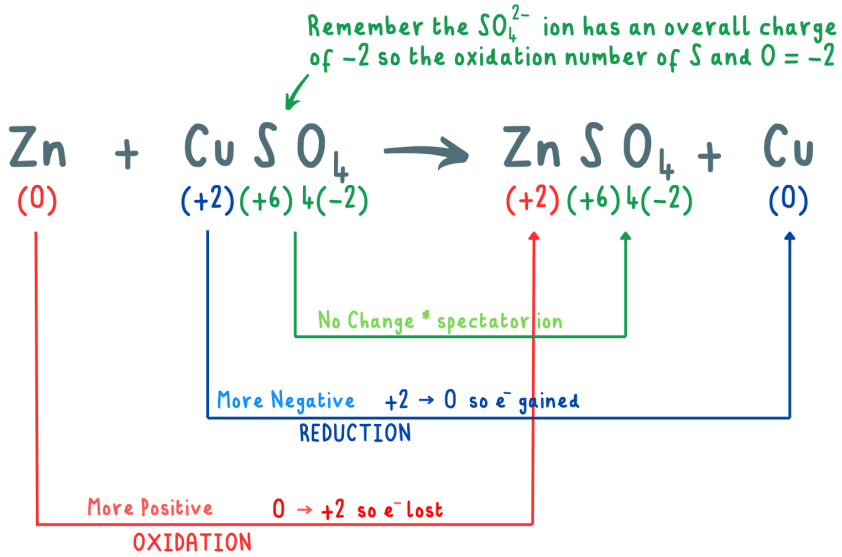
$2I^- \rightarrow I_2 + 2e^-$

Iodide ions have lost (donated) electrons to the chlorine, bringing about reduction, whilst being oxidised itself so the iodide ions are the reducing agent

$Cl_2 + 2e^- \rightarrow 2Cl^-$

Chlorine has gained electrons from the iodide ions, bringing about oxidation, whilst being reduced itself so the chlorine is the oxidising agent.

E.g. Identify the oxidising and reducing agents in the following redox reaction using oxidation numbers : $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$



Zinc has become oxidised from 0 in Zn to +2 in $ZnSO_4$, losing electrons and bringing about reduction whilst being oxidised itself = Zinc is the reducing agent.
Copper ions have become reduced from +2 in $CuSO_4$ to 0 in Cu , gaining electrons and bringing about oxidation whilst being reduced itself = Copper ions are the oxidising agent