Cambridge (CIE) IGCSE Chemistry

Your notes

Redox

Contents

Your notes

Oxidation & reduction

Roman numerals and oxidation numbers

- Transition elements can bond in different ways by forming ions with different charges
- When naming, the charge on the ion is shown by using a Roman numeral after the element's name
- For example, iron can form different ions
  - Iron ions with a 2+ charge, Fe^{2+}, are called **iron(II)** ions
  - Iron ions with a 3+ charge, Fe^{3+}, are called **iron(III)** ions
- The Roman numeral is the **oxidation number** of the element
- For example, iron reacts with oxygen to form iron oxide
  - But the name and formula of the product depend on the oxidation state of the iron ions
  - Iron(II) ions
    - The iron oxide product contains iron(II) ions, Fe^{2+}
    - The name of this product is iron(II) oxide
    - The formula of this product is FeO
  - Iron(III) ions
    - The iron oxide product contains iron(III) ions, Fe^{3+}
    - The name of this product is iron(III) oxide
    - The formula of this product is Fe_2O_3

Worked Example

1. State the oxidation number of the transition metal ion in silver(I) chloride, AgCl.
2. Name, including Roman numeral, the compound with the formula $CoCl_2$.

Answers:

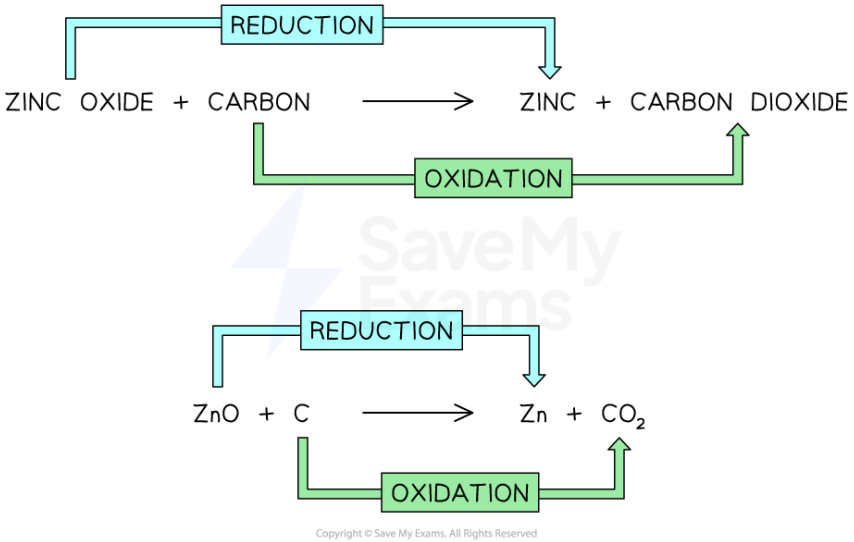
1. The oxidation number of the transition metal ion in silver(I) chloride, AgCl, is 1
   - **Remember:** The Roman numeral gives the oxidation number of the element before
2. The name, including Roman numeral, of the chemical with the formula $CoCl_2$ is cobalt(II) chloride

- $CoCl_2$ contains two chloride ions, each with a 1– charge
- So, the cobalt ion has a 2+ charge
- This means the compound contains cobalt(II) ions, Co^{2+}
- Therefore, it is called cobalt(II) chloride

Your notes

What is a redox reaction?

- A **redox** reaction is where oxidation and reduction take place together at the same time in the same reaction
- In terms of oxygen:
  - **Oxidation** is where **oxygen** is **added** to an element or a compound
  - **Reduction** is where **oxygen** is **removed** from an element or compound
- The reaction between zinc oxide and carbon is an example of a redox reaction
  - Zinc oxide is reduced because it has **lost** oxygen
  - Carbon is oxidised because it has **gained** oxygen



Both reduction and oxidation has occurred in this reaction so it is classed as a redox reaction

Worked Example

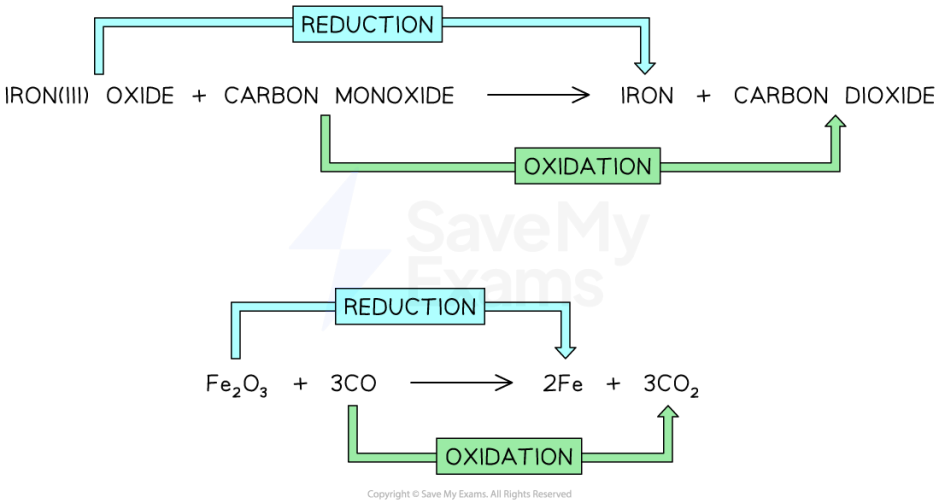
Explain which chemicals that are reduced and oxidised in the reaction between iron(III) oxide and carbon.

iron(III) oxide + carbon monoxide → iron + carbon dioxide

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Answer:

- Iron(III) oxide loses oxygen, so it is reduced
- Carbon monoxide gains oxygen, so it is oxidised



Examiner Tips and Tricks

You may see the term oxidation state used instead of oxidation number.

Although there is a subtle difference between the two terms (this is beyond the scope of this course), they are often used interchangeably.

Usually oxidation number is used to refer to the Roman numerals found within the name.

Your notes

Your notes

Redox & electron transfer

Extended tier only

- Redox reactions can also be defined in terms of electron transfer
- **Oxidation** is a reaction in which an element, ion or compound **loses electrons**
  - The **oxidation number** of the element is **increased**
  - This can be shown in a half-equation, e.g. when silver reacts with chlorine, silver is **oxidised** to silver ions:

$$Ag \rightarrow Ag^+ + e^-$$

- **Reduction** is a reaction in which an element, ion or compound **gains electrons**
  - The **oxidation number** of the element is **decreased**
  - This can be shown in a half-equation, e.g. when oxygen reacts with magnesium, oxygen is **reduced** to oxide ions:

$$O_2 + 4e^- \rightarrow 2O^{2-}$$

- For example, when iron reacts with a compound of copper such as copper sulfate, a displacement reaction occurs

iron + copper sulfate → iron(II) sulfate + copper

$$\mathbf{Fe + CuSO_4 \rightarrow FeSO_4 + Cu}$$

- We can write this as an ionic equation

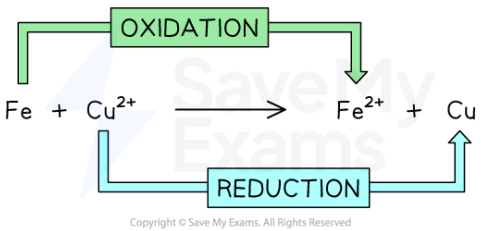
$$\mathbf{Fe + Cu^{2+} + SO_4^{2-} \rightarrow Fe^{2+} + SO_4^{2-} + Cu}$$

- We can then remove the spectator ions to see the overall change

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

- The iron atom has lost electrons to become a positive ion, so has been oxidised
- The positive copper ion has gained electrons to become an atom, so have been reduced

The redox reaction between Fe and Cu^{2+}



The Fe atom is oxidised (loses electrons) and the Cu^{2+} ion is reduced (gains electrons)

Your notes

Worked Example

Which change in the following equation is oxidation?

$$V^{3+} + Fe^{3+} \rightarrow V^{4+} + Fe^{2+}$$

Answer:

- **Step 1** - Identify the changes for each species:
  - V^{3+} to V^{4+}
    - V^{3+} has lost 1 electron
  - Fe^{3+} to Fe^{2+}
    - Fe^{3+} has gained 1 electron
- **Step 2** - Identify each change as either oxidation and reduction
  - V^{3+} to V^{4+} is oxidation
  - Fe^{3+} to Fe^{2+} is reduction
- Therefore, V^{3+} has been oxidised

Examiner Tips and Tricks

Use the mnemonic **OIL-RIG** to remember oxidation and reduction in terms of the movement of electrons:

- **O**xidation **I**s **L**oss
- **R**eduction **I**s **G**ain.

Identifying redox reactions

Extended tier only

Identifying redox reactions using oxidation numbers

- The **oxidation number** is a number assigned to an atom or ion in a compound
- It shows the number of electrons that an atom has lost, gained or shared in forming a compound
  - So, the oxidation number helps you to keep track of the movement of electrons in a redox process
- It is written as a +/- sign followed by a number
  - Positive oxidation number = loss of electrons

- Negative oxidation number = gain of electrons
- For example, aluminium in a compound usually has the oxidation number of +3 indicating it has lost 3 electrons
- **Careful:** It is easy to confuse oxidation number with charge which is written by a number followed by a +/- sign)

- A few simple rules help guide you through the process of determining the oxidation number of any element

Your notes

Rules for assigning oxidation numbers

	Rule	Example
1	The oxidation number of any uncombined element is zero	H_2 Zn O_2
2	Many atoms or ions have fixed oxidation numbers in compounds	Group 1 elements are always +1 Group 2 elements are always +2 Fluorine is always –1 Hydrogen is +1, except in hydrides like NaH where it is –1 Oxygen is –2, except in peroxides where it is in –1 and in F_2O where it is +2
3	The oxidation number of an element in a monoatomic ion is always the same as the charge	Zn^{2+} = +2 Fe^{3+} = +3 Cl^- = –1
4	The sum of the oxidation numbers in a compound is zero	NaCl Na = +1 Cl = –1 Sum of oxidation numbers = 1 – 1 = 0
5	The sum of the oxidation numbers in an ion is equal to the charge on the ion	SO_4^{2-} S = +6 Four O atoms = 4 x (–2) = –8 Sum of oxidation numbers = 6 – 8 = –2
6	In either a compound or an ion, the more electronegative element is given the negative oxidation number	F_2O Two F atoms = 2 x (–1) = –2 O = +2

- Redox reactions can be identified by the changes in the **oxidation number** when a reactant goes to a product

Your notes

Worked Example

The equation for the reaction between chlorine and potassium iodide is shown below.

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

Identify which species has been:

1. Oxidised
2. Reduced

Answer:

1. The species that has been oxidised is iodine
    - $2I^- \rightarrow I_2 + 2e^-$
    - The oxidation number of I^- is -1
    - The oxidation number of iodine in I_2 is 0
    - The oxidation number has increased so the iodide ions have been oxidised / lost electrons
2. The species that has been reduced is chloride ions
    - $Cl_2 + 2e^- \rightarrow 2Cl^-$
    - The oxidation number of chlorine as Cl_2 is 0
    - The oxidation number of Cl^- is -1
    - The oxidation number has decreased so the Cl_2 has been reduced / gained electrons

Identifying redox reactions by colour changes

- The tests for redox reactions involve the observation of a colour change in the solution being analysed
- Two common examples are acidified potassium manganate(VII), and potassium iodide
- Potassium manganate(VII), $KMnO_4$, is an **oxidising agent** which is often used to test for the presence of **reducing** agents
- When acidified potassium manganate(VII) is added to a reducing agent its colour changes from **purple** to **colourless**

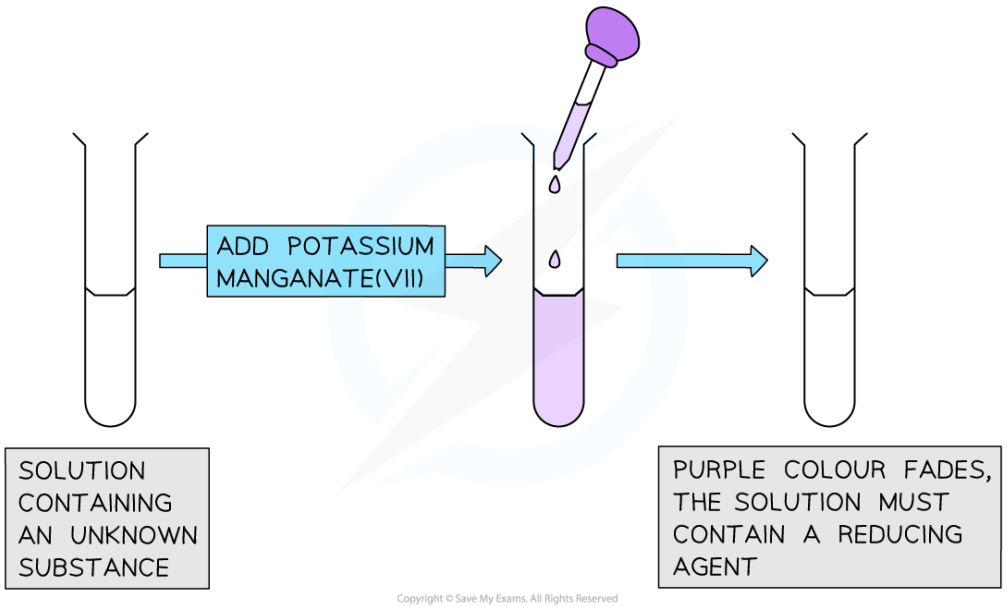


Diagram to show the colour change when potassium manganate(VII) is added to a reducing agent

- Potassium iodide, KI, is a **reducing** agent which is often used to test for the presence of **oxidising** agents
- When added to an acidified solution of an oxidising agent such as aqueous chlorine or hydrogen peroxide (H_2O_2), the solution turns a **red-brown** colour due to the formation of iodine, I_2

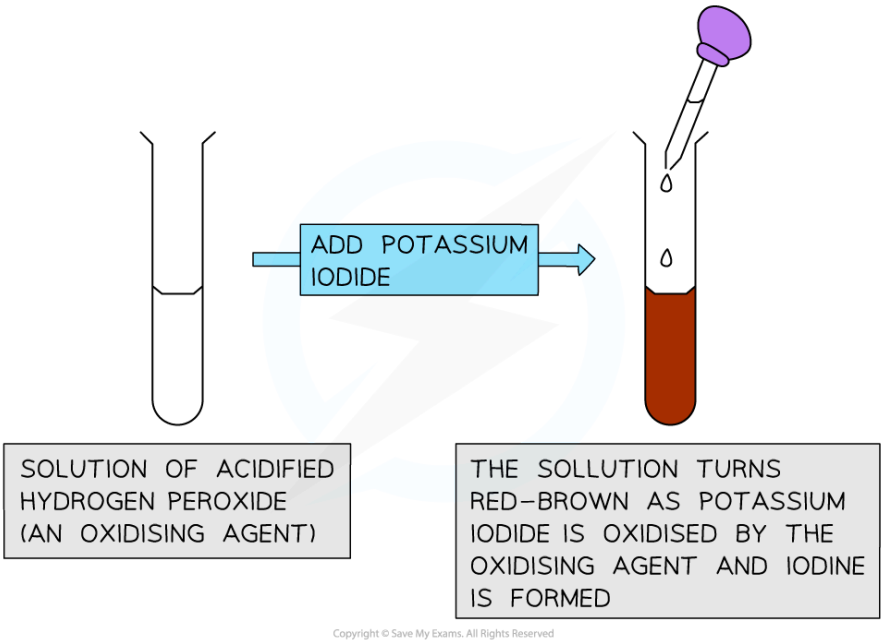


Diagram to show the colour change when potassium iodide is added to an oxidising agent

- The potassium iodide is oxidised as it loses electrons
- The hydrogen peroxide is reduced
- Therefore, potassium iodide is acting as a **reducing agent**


Your notes

Oxidising & reducing agents

Extended tier only

What is an oxidising agent?

- An oxidising agent is a substance that **oxidises** another substance, and becomes **reduced** in the process
- An oxidising agent **gains** electrons as another substance loses electrons
- Common examples include hydrogen peroxide, fluorine and chlorine

What is a reducing agent?

- A reducing agent is a substance that **reduces** another substance, and becomes **oxidised** in the process
- A reducing agent **loses** electrons as another substance gains electrons
- Common examples include carbon and hydrogen
- The process of reduction is very important in the chemical industry as a means of extracting metals from their ores

Identifying oxidising and reducing agents

$$CuO + H_2 \rightarrow Cu + H_2O$$

- Hydrogen is reducing the CuO
- Hydrogen is itself oxidised as it has gained oxygen / lost electrons
- So, the **reducing agent** is **hydrogen**:

$$H_2 \rightarrow 2H^+ + 2e^-$$

- CuO is reduced by hydrogen
  - This means that the hydrogen is oxidised by CuO
- CuO is reduced as it has lost oxygen / gained electrons
- So, the **oxidising agent** is **copper oxide**

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

Worked Example

When iron reacts with bromine to form iron(II) bromide, a redox reaction reaction occurs:

$$Fe + Br_2 \rightarrow FeBr_2$$

Which species is acting as the reducing agent in this reaction?

Answer

1. **Step 1** - Write half equations to work out what has gained/lost electrons
   - $Fe \rightarrow Fe^{2+} + 2e^-$
   - $Br_2 + 2e^- \rightarrow 2Br^-$
   - Fe loses electrons; Br_2 gains electrons
2. **Step 2** - Deduce what has been oxidised/reduced (remember OIL RIG)
   - Fe has been oxidised as it has lost electrons
   - Br_2 has been reduced as it has gained electrons
3. **Step 3** - Identify the reducing agent
   - Fe is the reducing agent as it has been oxidised by losing electrons and caused Br_2 to be reduced as it gained electrons

Your notes