

Cambridge (CIE) A Level Chemistry



Your notes

Redox Processes: Electron Transfer & Changes in Oxidation Number (Oxidation State)

Contents

- * Oxidation Number Rules
- * Redox Reactions
- * Oxidising & Reducing Agents



Oxidation Numbers

- The oxidation number (also known as oxidation state) is a number given to each atom or ion in a compound to keep track of how many electrons they have
- In a single ion or molecular ion, the oxidation number tells us how many electrons have been lost or gained
- Positive oxidation number = loss of electrons
- Negative oxidation number = gain of electrons

Oxidation number rules

- The oxidation number refers to a *single* atom in a compound

	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 ₂ O 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	F ₂ O Two F atoms = 2 x (-1) = -2



Worked Example

Deducing oxidation numbers

State the oxidation number of the bold atoms in these compounds or ions.

1. **P**₂O₅
2. **S**O₄²⁻
3. H₂**S**
4. **Al**₂Cl₆
5. **N**H₃
6. **Cl**O₂⁻
7. Ca**C**O₃

Answer

1. **P**₂O₅
 - 5 O atoms = $5 \times (-2) = -10$
 - The overall charge of the compound = 0
 - 2 P atoms = +10
 - Oxidation number of 1 P atom = $(+10) / 2 = +5$
2. **S**O₄²⁻
 - 4 O atoms = $4 \times (-2) = -8$
 - The overall charge of the compound = -2
 - The oxidation number of 1 S atom = +6
3. H₂**S**
 - 2 H atoms = $2 \times (+1) = +2$
 - The overall charge of the compound = 0
 - The oxidation number of 1 S atom = -2
4. **Al**₂Cl₆
 - 6 Cl atoms = $6 \times (-1) = -6$
 - The overall charge of the compound = 0
 - 2 Al atoms = +6
 - The oxidation number of 1 Al atom = $(+6) / 2 = +3$
5. **N**H₃
 - 3 H atoms = $3 \times (+1) = +3$
 - The overall charge of the compound = 0
 - The oxidation number of 1 N atom = -3
6. **Cl**O₂⁻
 - 2 O atoms = $2 \times (-2) = -4$
 - The overall charge of the compound = -1
 - The oxidation number of 1 Cl atom = +3
7. Ca**C**O₃

- 3 O atoms = $3 \times (-2) = -6$
- 1 Ca atom = $+2$
- The overall charge of the compound = 0
- The oxidation number of 1 C atom = $+4$



Your notes



Balancing Redox Reactions

- Oxidation numbers can be used to balance chemical equations
- Go through these steps to balance a redox equation:
 1. Write the unbalanced equation and identify the atoms which change in ox. no.
 2. Deduce the oxidation number changes
 3. Balance the oxidation number changes
 4. Balance the charges
 5. Balance the atoms



Worked Example

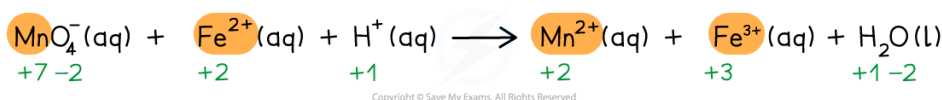
Writing overall redox reactions

Manganate(VII) ions, MnO_4^- , react with Fe^{2+} ions in the presence of acid, H^+ , to form Mn^{2+} ions, Fe^{3+} ions and water.

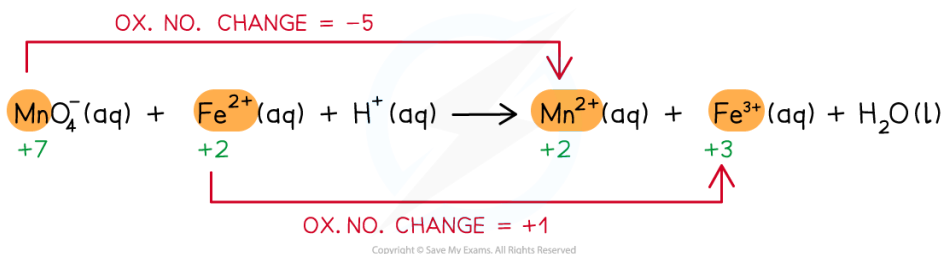
Write the overall redox equation for this reaction.

Answer

- **Step 1:** Write the unbalanced equation and identify the atoms which change in oxidation number:



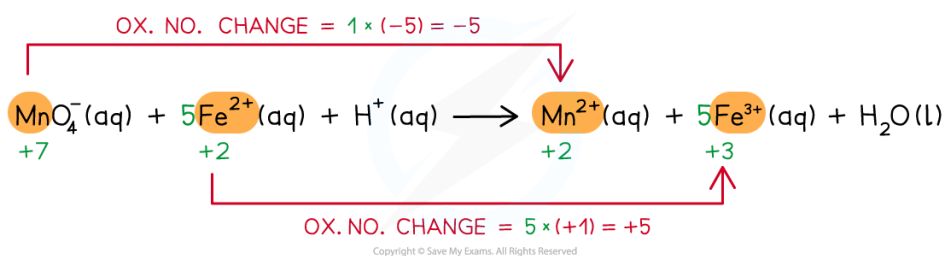
- **Step 2:** Deduce the oxidation number changes:



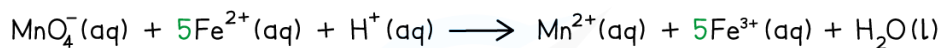
- **Step 3:** Balance the oxidation number changes:



Your notes



- Step 4: Balance the charges:



IGNORING H^+

• TOTAL \oplus CHARGE = $(5 \times 2+) = 10+$

• TOTAL \oplus CHARGE = $5 \times (3+) + (2+) = 17+$

• TOTAL \ominus CHARGE = $1-$

• TOTAL \ominus CHARGE = 0

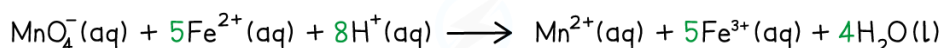
TOTAL CHARGE REACTANTS = $9+$

TOTAL CHARGE PRODUCTS = $17+$

THUS 8H^+ IONS ARE NEEDED TO BALANCE THE CHARGES ON BOTH SIDES

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- Step 5: Balance the atoms:



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Redox & Disproportionation Reactions

Oxidation

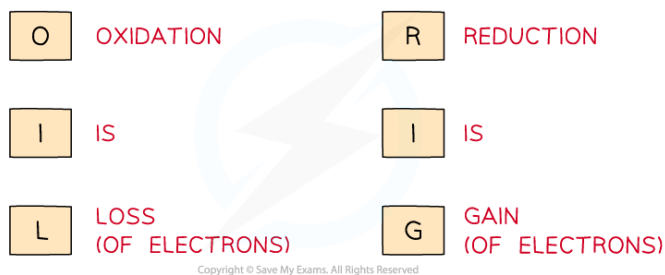
- Oxidation is the **gain of oxygen**, e.g.:
 - $\text{Cu} + \text{H}_2\text{O} \rightarrow \text{CuO} + \text{H}_2$
 - Cu has gained an oxygen and is oxidised
- Oxidation is also the **loss of a hydrogen**, e.g.:
 - $2\text{NH}_3 + 3\text{Br}_2 \rightarrow \text{N}_2 + 6\text{HBr}$
 - N has lost a hydrogen and is oxidised
- Oxidation is also the **loss of electrons**, e.g.:
 - $\text{Cu}^{2+} + \text{Mg} \rightarrow \text{Mg}^{2+} + \text{Cu}$
 - Mg has lost two electrons and is oxidised
- Oxidation causes an **increase in oxidation number**, e.g.:
 - $\text{Cu}^{2+} + \text{Mg} \rightarrow \text{Mg}^{2+} + \text{Cu}$

- The oxidation number of Mg changes from 0 to +2, thus Mg is oxidised

Reduction

- Reduction is the **loss of oxygen**, e.g.:
 - $\text{Cu} + \text{H}_2\text{O} \rightarrow 2\text{CuO} + \text{H}_2$
 - H has been reduced
- Reduction is also the **gain of a hydrogen**, e.g.:
 - $2\text{NH}_3 + 3\text{Br}_2 \rightarrow \text{N}_2 + 6\text{HBr}$
 - Br has been reduced
- Reduction is also the **gain of electrons**, e.g.:
 - $\text{Cu}^{2+} + \text{Mg} \rightarrow \text{Mg}^{2+} + \text{Cu}$
 - Cu has been reduced
- Reduction causes a **decrease in oxidation number**, e.g.:
 - $\text{Cu}^{2+} + \text{Mg} \rightarrow \text{Mg}^{2+} + \text{Cu}$
 - The oxidation number of Cu changes from +2 to 0, thus Cu is reduced

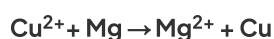
OIL RIG acronym



Use the acronym "Oil Rig" to help you remember the definitions of oxidation and reduction

Redox reactions

- Redox reactions are reactions in which oxidation and reduction take place together
 - While one species is oxidising, another is reducing in the same reaction
- For example:



- Cu has been reduced from +2 to 0
- Mg has been oxidised from 0 to +2



Your notes





Worked Example

Oxidation and reduction

In each of the following equations, state which reactant has been oxidised and which has been reduced.

- $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$
- $\text{Mg} + \text{Fe}^{2+} \rightarrow \text{Mg}^{2+} + \text{Fe}$
- $\text{CO} + \text{Ag}_2\text{O} \rightarrow 2\text{Ag} + \text{CO}_2$

Answer

■ Answer 1:

- Oxidised: Na as the oxidation number has increased from 0 to +1
- Reduced: Cl_2 as the oxidation number has decreased from 0 to -1

■ Answer 2:

- Oxidised: Mg as the oxidation number has increased by 2
- Reduced: Fe^{2+} as the oxidation number has decreased by 2

■ Answer 3:

- Oxidised: C as it has gained oxygen
- Reduced: Ag as it has lost oxygen



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Disproportionation reactions

- A **disproportionation reaction** is a reaction in which the same species is both oxidised and reduced at the same time

Example disproportion reaction



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Worked Example

Balancing disproportionation reactions

Balance the disproportionation reaction which takes place when chlorine is added to hot concentrated aqueous sodium hydroxide.

The products are Cl^- ions, ClO_3^- ions and water.

Answer

- **Step 1:** Write the unbalanced equation and identify the atoms that change in oxidation number:

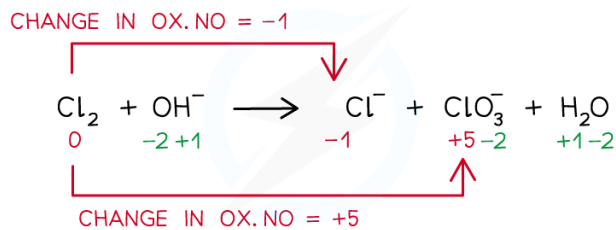


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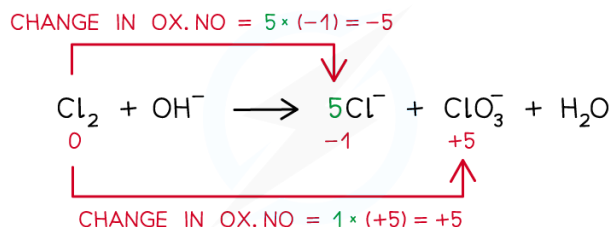
Your notes

- Step 2: Deduce the oxidation number changes:



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- Step 3: Balance the oxidation number changes:



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- Step 4: Balance the charges:



• TOTAL \oplus CHARGE = 0

• TOTAL \oplus CHARGE = 0

• TOTAL \ominus CHARGE = 1-

• TOTAL \ominus CHARGE = $5 \times (1-) + (1-) = 6-$

TOTAL CHARGE REACTANTS = 1-

TOTAL CHARGE PRODUCTS = 6-

THUS 6 OH⁻ IONS ARE NEEDED TO BALANCE THE CHARGES ON BOTH SIDES

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- Step 5: Balance the atoms:



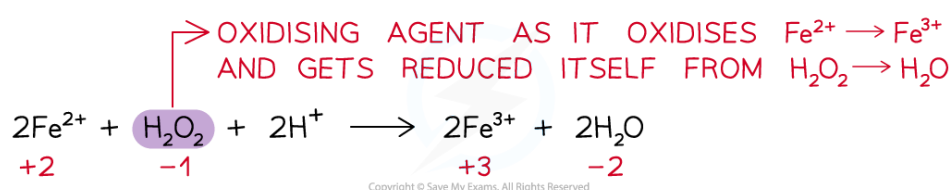
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Oxidising & Reducing Agents

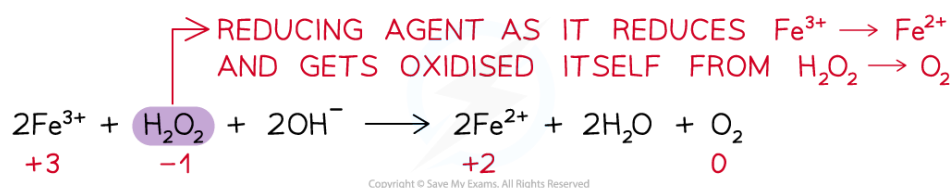
Oxidising agent

- An **oxidising agent** is a substance that **oxidises** another atom or ion by causing it to lose electrons
- An oxidising agent itself gets **reduced**
 - This means that an oxidising agent **gains electrons**
- Therefore, the **oxidation number** of the oxidising agent **decreases**
- For example:



Reducing agent

- A **reducing agent** is a substance that **reduces** another atom or ion by causing it to gain electrons
- A reducing agent itself gets **oxidised**
 - This means that a reducing agent **loses / donates electrons**
- Therefore, the **oxidation number** of the reducing agent **increases**
- For example:



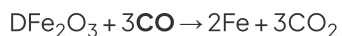
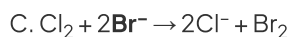
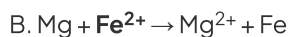
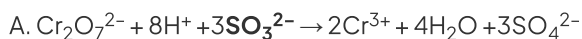
- For a reaction to be recognised as a redox reaction, there must be both an oxidising and reducing agent
- Some substances can act both as oxidising and reducing agents
- Their nature is dependent upon what they are reacting with and the reaction conditions





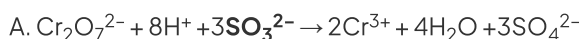
Worked Example

In which of the following reactions is the species in bold acting as an oxidising agent?

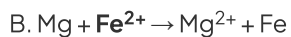


Answer

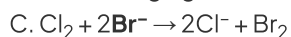
- Oxidising agents are substances that oxidise other species, gain electrons and are themselves reduced.



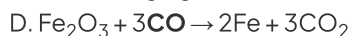
- The SO_3^{2-} has an oxidation number of -2
- This becomes SO_4^{2-} , which still has an oxidation number of -2
- There is no change in oxidation number so this equation cannot be the correct answer



- The Fe^{2+} has an oxidation number of $+2$
- This becomes Fe , which has an oxidation number of 0
- The oxidation number has decreased, which means that the **Fe^{2+} has acted as an oxidising agent and this equation is the correct answer**



- The Br^- has an oxidation number of -1
- This becomes Br_2 , which has an oxidation number of 0
- The oxidation number has increased, which means that Br^- has acted as a reducing agent and this equation cannot be the correct answer



- The CO has an oxidation number of 0
- This becomes CO_2 , which still has an oxidation number of 0
- There is no change in oxidation number so this equation cannot be the correct answer

Roman numerals

- Roman numerals are used to show the **oxidation states of transition metals**, which can have more than one oxidation number
- Iron can be both $+2$ and $+3$, so Roman numerals are used to distinguish between them
 - Fe^{2+} in FeO can be written as **iron(II) oxide**
 - Fe^{3+} in Fe_2O_3 can be written as **iron(III) oxide**



Your notes

Worked Example

Give the full systematic names of the following compounds:

1. FeCl_2
2. HClO_4
3. NO_2
4. $\text{Mg}(\text{NO}_3)_2$
5. K_2SO_4

Answer

■ **Answer 1:** FeCl_2

- The oxidation number of 2 Cl atoms = -2
- FeCl_2 has no overall charge
- So, the oxidation number of Fe is $+2$
- Therefore, the systematic name is iron(II) chloride

■ **Answer 2:** HClO_4

- The oxidation number of the H atom = $+1$
- The oxidation number of 4 O atoms = -8
- HClO_4 has no overall charge
- So, the oxidation number of Cl is $+7$
- Therefore, the systematic name is chloric(VII) acid

■ **Answer 3:** NO_2

- The oxidation number of 2 O atoms = -4
- NO_2 has no overall charge
- So, the oxidation number of N is $+4$
- Therefore, the systematic name is nitrogen(IV) oxide

■ **Answer 4:** $\text{Mg}(\text{NO}_3)_2$

- The systematic name is magnesium nitrate
- This is a salt of the common acid, which means that they are named without including the oxidation number of the non-metal

■ **Answer 5:** K_2SO_4

- The systematic name is potassium sulfate
- Again, this is a salt of the common acid, so it is named without including the oxidation number of the non-metal



Your notes