

Oxidation and Reduction 1

To succeed with this topic, you need to:

- understand what is meant by ionic and covalent bonding (covered in Factsheet 5 - Bonding), and recognise which compounds have which type of bonding
- be able to do simple arithmetic involving negative numbers
- understand how to produce balanced chemical equations

After working through this Factsheet, you will understand:

- the meaning of the term Oxidation Number
- how oxidation numbers are found and used in compounds
- the definitions of oxidation and reduction in terms of electrons and changes in oxidation numbers
- the use of ionic equations to form balanced chemical equations
- the idea of "redox"

Exam Hint: Oxidation numbers and oxidation/reduction are used throughout all the AS topics. The examination papers assume that you have this knowledge, so when the terms occur in questions you are expected to be able to use them. It is common to see questions on this topic appear several times (as smaller parts) throughout the paper.

Success in this topic relies on you learning the rules thoroughly, and practising plenty of questions so that you are comfortable with the concepts and calculations involved.

Introduction

At GCSE level, the term **oxidation** meant adding oxygen to, or removing hydrogen from, an element or compound. **Reduction** was the reverse – i.e. add hydrogen or remove oxygen.

At AS level oxidation/reduction does **not** have to involve oxygen or hydrogen – it is much more general. Some candidates are initially confused by this – avoid this confusion by **learning** the new definitions in this Factsheet.

Oxidation Number

Definition: Oxidation Number (ON) is the number assigned to an atom or ion to describe its relative state of oxidation or reduction.

To understand what this means and how it works, we will look at the **rules** for assigning oxidation numbers, then work through some examples.

Worked Examples

1. What is the ON of lead in PbCl_4 ?

Chlorine is one of the elements with a fixed ON (see Rules) of -1 .

$$\text{So } \text{Pb} + 4 \times -1 = 0.$$

So ON of Pb is $+4$

2. What is the ON of manganese in Mn_2O_7 ?

Oxygen has a constant ON of -2 (see Rules)

$$2\text{Mn} + 7 \times -2 = 0$$

$$2\text{Mn} = 14$$

So ON of Mn is $+7$

3. What is the ON of iodine in KIO_3 ?

Potassium ($+1$) and oxygen (-2) have fixed ONs (see Rules).

$$\text{So } +1 + \text{I} + 3 \times -2 = 0$$

So ON of I is $+5$

4. What is the ON of chlorine in Cl_2O_3 ?

Cl is **not** -1 when combined with oxygen (see rules). Oxygen is -2 .

$$2\text{Cl} + 3 \times -2 = 0$$

So ON of Cl is $+3$

5. What is the ON of carbon in CO_3^{2-} ?

O has a constant ON of -2 , and this is a polyatomic ion with charge -2 .

$$\text{So } \text{C} + 3 \times -2 = -2$$

So ON of C is $+4$

Assigning Oxidation Numbers - The Rules

- Elements are zero – eg the ON of H in H_2 is 0.
- In ionic compounds, the oxidation number is the same as the charge on the ion. Eg in NaCl, we have the ions Na^+ and Cl^- . The oxidation number of Na is therefore $+1$, and of Cl is -1
- The sum of all the ONs in a compound is zero. Eg in MgCl_2 , the ON of Mg is $+2$, and the ON of each Cl is -1 .
So overall we have $+2 + (2 \times -1) = 0$
- The sum of all the ONs in a polyatomic ion is the same as the charge on the ion. Eg in CO_3^{2-} , C has ON $+4$ and each O has ON -2 .
So $+4 + (3 \times -2) = -2$
- In covalent compounds, one element must be positive ON and another element must be negative, so overall they equal zero.
- The ONs of the following elements in compounds must be learnt:
 - Group 1 (Li, Na, K etc) are $+1$
 - Group 2 (Mg, Ca, Sr etc) are $+2$
 - Al is $+3$
 - O is -2 , except in peroxides (such as H_2O_2) where it is -1
 - H is $+1$ except in metal hydrides (such as NaH) where it is -1
 - F is -1
 - Cl is -1 , except in compounds with O and F

6. What is the ON of iron in $\text{Fe}(\text{CN})_6^{4-}$?

Since the cyanide ion is CN^- , it has an ON of -1

So we have $\text{Fe} + 6 \times -1 = -4$

So ON of Fe is $+2$

NB: The use of roman numerals in the names of ionic compounds now also gives the oxidation number as well as the charge on the metal ion (cation). Eg copper (I) oxide is Cu^+ with O^{2-} - so giving Cu_2O . The copper has a $+1$ oxidation number.

Being able to calculate oxidation numbers is the basis for all this topic. Before going on to the rest of this Factsheet, try working through Question 1 at the back to make sure you are happy with finding oxidation numbers.

Oxidation and Reduction

Definition:

Oxidation takes place when an element in a reaction

- Increases its oxidation number
- Loses electrons

Reduction takes place when an element in a reaction

- Decreases its oxidation number
- Gains electrons

To help you remember this, you can use "REG" – **R**eduction **E**lectron **G**ain. If you remember oxidation is the reverse of reduction, then you'll be able to work out that oxidation involves electron loss.

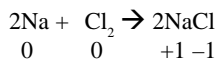
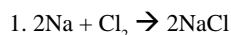
In fact, oxidation never takes place on its own - nor does reduction. When one substance is oxidised in a reaction, another one is reduced. A **Redox** reaction is one in which both reduction and oxidation take place.

To work out which element is oxidised and which is reduced in a reaction, we go through these steps:

- 1 Write down the chemical equation for the reaction.
- 2 Go through and work out the oxidation numbers for every element in the equation.
- 3 Look for an element that has **increased** its oxidation number from one side of the equation to the other - it has been **oxidised**.
- 4 Look for an element that has **decreased** its oxidation number from one side of the equation to the other - it has been **reduced**.

Worked examples

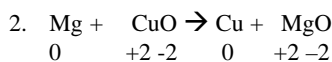
For each reaction, work out which element is being oxidised and which element is being reduced:



Notice that we ignore the numbers in front of formulae/symbols in this exercise.

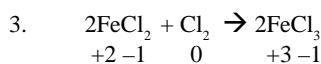
Na: $0 \rightarrow +1$ is oxidation.

Cl: $0 \rightarrow -1$ is reduction



Mg: $0 \rightarrow +2$ is oxidation

Cu: $+2 \rightarrow 0$ is reduction



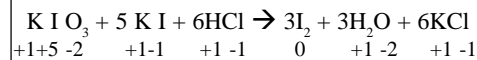
Fe: $+2 \rightarrow +3$ is oxidation

Cl: $0 \rightarrow -1$ is reduction

To practise using oxidation numbers to identify oxidation and reduction, go to Question 2 at the end.

Disproportionation

Disproportionation occurs when an element is both oxidised and reduced in the same reaction. You will meet examples of this when you study Group VII - the halogens. The reaction below shows one example:



Note that the iodine in KIO_3 goes from ON $+5$ to 0 - so it is reduced. The iodine in KI goes from ON -1 to 0 - so it is oxidised.

Oxidising and Reducing Agents

Definition:

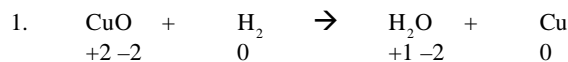
An **oxidising agent** (or **oxidant**) is the species that causes another element to increase in oxidation number (oxidise). It is itself reduced in the reaction.

A **reducing agent** (or **reductant**) is the species that causes another element to decrease in oxidation number (reduce). It is itself oxidised in the reaction.

Note that the oxidising or reducing agent is the **compound** involved in the reaction - not one of the elements in it.

Worked Examples

For each of the following, identify which species are being oxidised and reduced as well as the oxidising and reducing agents.

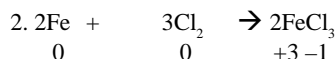


Cu: $+2 \rightarrow 0$ is reduced

H: $0 \rightarrow +1$ is oxidised

CuO is the **oxidising agent**

H_2 is the **reducing agent**



Fe: $0 \rightarrow +3$ is oxidised

Cl: $0 \rightarrow -1$ is reduced

Fe is the **reducing agent**

Cl_2 is the **oxidising agent**

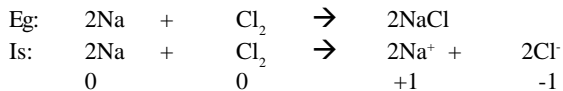
For further practise on identifying oxidising and reducing agents, go to Question 3 at the end.

Oxidising agents are species that "like" to accept electrons - common examples are halogen or oxygen atoms (so oxygen is an oxidising agent...). Reducing agents are species that "like" to give away electrons - common examples are metal atoms and hydrogen.

Exam Hint: To help you avoid getting confused about oxidising and reducing agents, you should remember that an oxidising agent oxidises **something else** - so it is reduced itself.

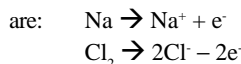
Ionic and Half Equations

As you have worked through this Factsheet and answered questions 1, 2 and 3, you have been using the concept of **ionic equations**



Writing ionic equations was covered in Factsheet No 3, but we will revisit the concept as we write half equations which involve **electrons**.

If we look at the sodium and chlorine reaction, then the two half-equations are:

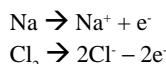


The **rules** for including the electrons are:

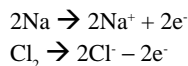


- If you have a + ion, you need to + electron(s)
- If you have a - ion, you need to - electron(s)
- the amount of charge tells you how many electrons – eg 2+ gives 2e⁻, 3- gives 3e⁻

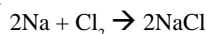
We can use two half-equations to produce the normal balanced chemical equation by balancing the electrons. Multiply each equation by the number of electrons in the other equation (ignoring signs), then add together. The electrons must cancel out.



Multiply the Na equation by the number of electrons in the Cl equation - that's 2. Multiply the Cl equation by the number of electrons in the Na equation - that's 1:



Add up:



Half-equations are important in redox work because of the definition of oxidation and reduction in terms of oxidation numbers and electrons (remember REG!)

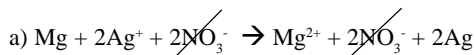
$\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$ is **oxidation** because electrons are being lost by Na
0 +1

$\text{Cl}_2 \rightarrow 2\text{Cl}^- - 2\text{e}^-$ is **reduction** because electrons are being gained by Cl
0 -1

Worked example

For $\text{Mg} + 2\text{AgNO}_3 \rightarrow \text{Mg}(\text{NO}_3)_2 + 2\text{Ag}$, write

- the ionic equation
- the two half-equations, identifying which is oxidation and reduction



Removing the NO_3^- which is a **spectator ion** (does not take part in the reaction) and so "cancels out" leaves:



- $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$ oxidation, as Mg loses electrons
 $2\text{Ag}^+ + 2\text{e}^- \rightarrow 2\text{Ag}$ reduction, because Ag⁺ gains electrons

Question 4 provides more practice on half-equations.

Questions

- What is the oxidation number of Cl in the following?
i) HCl ii) HClO₃ iii) ClO[•] iv) Cl₂O₇²⁻ v) ClO₄⁻
 - What is the oxidation number of S in the following?
i) SO₂ ii) Na₂S₂O₃ iii) SO₄²⁻ iv) SO₃²⁻ v) SO₃
 - What is the oxidation number of N in the following?
i) N₂H₄ ii) HNO₃ iii) NaNO₂ iv) HCN v) N₂O
 - Find the oxidation number of the identified elements in the following:
i) Ba in BaCl₂ ii) Li in Li₂O iii) P in P₂O₅
iv) P in P₂O₃ v) C in CO vi) I in I₂
vii) C in CCl₄ viii) I in I⁻ ix) Cr in CrO₄²⁻
x) Br in BrO₃⁻ xi) H in LiH xii) N in N₂O₄
xiii) Fe in Fe(CN)₆³⁻ xiv) Cr in Cr₂O₇²⁻ xv) P in PO₄³⁻
xvi) O in H₂O₂ xvii) Mn in MnO₂ xviii) Xe in XeF₄
xix) N in NH₃ xx) N in NaNO₃

2. Identify which species is being oxidised and which species is being reduced in each of the following reactions.

- $\text{Zn} + \text{Ag}_2\text{O} \rightarrow \text{ZnO} + 2\text{Ag}$
- $2\text{K} + \text{Br}_2 \rightarrow 2\text{KBr}$
- $\text{PbO}_2 + 4\text{H}^+ + \text{Sn}^{2+} \rightarrow \text{Pb}^{2+} + \text{Sn}^{4+} + 2\text{H}_2\text{O}$
- $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$
- $2\text{K} + \text{H}_2 \rightarrow 2\text{KH}$
- $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$
- $\text{Sn}^{2+} + 2\text{Fe}^{3+} \rightarrow \text{Sn}^{4+} + 2\text{Fe}^{2+}$
- $2\text{Cu} + \text{O}_2 \rightarrow 2\text{CuO}$
- $\text{KIO}_3 + 2\text{Na}_2\text{S}_2\text{O}_3 \rightarrow \text{KIO} + 2\text{Na}_2\text{SO}_4$
- $\text{Cu}^{2+} + \text{Cu} \rightarrow 2\text{Cu}^+$
- $2\text{S}_2\text{O}_3^{2-} + \text{I}_2 \rightarrow 2\text{I}^- + \text{S}_4\text{O}_6^{2-}$
- $2\text{FeI}_2 + \text{I}_2 \rightarrow 2\text{FeI}_3$
- $\text{IO}_3^- + 5\text{I}^- + 6\text{H}^+ \rightarrow 3\text{I}_2 + 3\text{H}_2\text{O}$
- $\text{I}_2 + \text{I}^- \rightarrow \text{I}_3^-$
- $\text{Zn} + 2\text{HNO}_3 \rightarrow \text{Zn}(\text{NO}_3)_2 + \text{H}_2$

3. Go back to question 2 and identify the oxidising agents and reducing agents

4. For each of the following equations

- write the two half equations
- identify the half-equations as oxidation or reduction

- $\text{Zn} + 2\text{AgCl} \rightarrow \text{ZnCl}_2 + 2\text{Ag}$
- $\text{PbO}_2 + 4\text{HCl} \rightarrow \text{PbCl}_2 + \text{Cl}_2 + 2\text{H}_2\text{O}$
- $2\text{K} + \text{H}_2 \rightarrow 2\text{KH}$
- $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$
- $\text{PH}_3 + 2\text{O}_2 \rightarrow \text{H}_3\text{PO}_4$

Answers

- i) -1 ii) +5 iii) +1 iv) +6 v) +7
 - i) +4 ii) +2 iii) +6 iv) +4 v) +6
 - i) -2 ii) +5 iii) +3 iv) -5 v) +1
 - i) +2 ii) +1 iii) +5 iv) +3 v) +2
 - i) 0 vii) +4 viii) -1 ix) +6 x) +5
 - xi) -1 xii) +4 xiii) +3 xiv) +6 xv) +5
 - xvi) -1 xvii) +4 xviii) +4 xix) -3 xx) +5

2. The species first given is oxidised, the second is reduced

- Zn Ag⁺ b) K Br₂ c) Sn²⁺ Pb⁴⁺
- H₂ N₂ e) K H₂ f) Mg H⁺
- Sn²⁺ Fe³⁺ h) Cu O₂ i) I S
- Cu Cu²⁺ k) S I₂ l) Fe²⁺ I₂
- I⁻ I n) I⁻ I o) Zn H⁺

3. The oxidising agent is given first, the reducing agent second.

- Ag₂O Zn b) Br₂ K c) PbO₂ Sn²⁺
- N₂ H₂ e) H₂ K f) HCl Mg
- Fe³⁺ Sn²⁺ h) O₂ Cu i) KIO₃ Na₂SO₃
- IO₃⁻ I⁻ k) I₂ S₂O₃²⁻ l) I₂ FeI₂
- Cu²⁺ Cu n) I₂ I⁻ o) HNO₃ Zn

- $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ oxidation $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$ reduction
 - $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$ oxidation $\text{Pb}^{4+} + 2\text{e}^- \rightarrow \text{Pb}^{2+}$ reduction
 - $\text{K} \rightarrow \text{K}^+ + \text{e}^-$ oxidation $\text{H}_2 \rightarrow 2\text{H} - 2\text{e}^-$ reduction
 - $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$ oxidation $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ reduction
 - $\text{P}^{3-} \rightarrow \text{P} (+5) + 2\text{e}^-$ oxidation $\text{O}_2 + 2\text{e}^- \rightarrow 2\text{O}^{2-}$ reduction

Acknowledgements: This Factsheet was researched and written by Sam Goodman & Kieron Heath. Curriculum Press, Unit 305B, The Big Peg, 120 Vyse Street, Birmingham, B13 6NF ChemistryFactsheets may be copied free of charge by teaching staff or students, provided that their school is a registered subscriber. No part of these Factsheets may be reproduced, stored in a retrieval system, or transmitted, in any other form or by any other means, without the prior permission of the publisher. ISSN 1351-5136