Chem Factsheet

January 2001 Number 11

# 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.
So Pb + 4 × -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)
2Mn + 7 × -2 = 0
2Mn = 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).
So +1 + I + 3 × -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.
2Cl + 3 × -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.
So C + 3 × -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 × -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× -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 $Fe(CN)_6^{4-}$?
Since the cyanide ion is $CN^-$, it has an ON of –1
So we have Fe + 6 × -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:

1. $2Na + Cl_2 \rightarrow 2NaCl$

| $2Na$ + | $Cl_2$ → | $2NaCl$ |
|---|---|---|
| 0 | 0 | +1 –1 |

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

Na: 0 → +1 is oxidation.
Cl: 0 → -1 is reduction

2. 

| Mg + | CuO → | Cu + | MgO |
|---|---|---|---|
| 0 | +2 -2 | 0 | +2 –2 |

Mg: 0 → +2 is oxidation
Cu: +2 → 0 is reduction

3. 

| $2FeCl_2$ + | $Cl_2$ → | $2FeCl_3$ |
|---|---|---|
| +2 –1 | 0 | +3 –1 |

Fe: +2 → +3 is oxidation
Cl: 0 → -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:

| $KIO_3$ + | $5KI$ + | $6HCl$ → | $3I_2$ + | $3H_2O$ + | $6KCl$ |
|---|---|---|---|---|---|
| +1+5 -2 | +1-1 | +1 -1 | 0 | +1 -2 | +1 -1 |

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.

1. 

| CuO + | $H_2$ → | $H_2O$ + | Cu |
|---|---|---|---|
| +2 –2 | 0 | +1 –2 | 0 |

Cu: +2 → 0 is reduced
H: 0 → +1 is oxidised
CuO is the **oxidising agent**
$H_2$ is the **reducing agent**

2. 

| 2Fe + | $3Cl_2$ → | $2FeCl_3$ |
|---|---|---|
| 0 | 0 | +3 –1 |

Fe: 0 → +3 is oxidised
Cl: 0 → -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**

| | | | | | | | |
|---|---|---|---|---|---|---|---|
| Eg: | 2Na | + | $Cl_2$ | → | 2NaCl | | |
| Is: | 2Na | + | $Cl_2$ | → | $2Na^+$ | + | $2Cl^-$ |
| | 0 | | 0 | | +1 | | -1 |

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:

$Na \rightarrow Na^+ + e^-$

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

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.

$Na \rightarrow Na^+ + e^-$

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

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:

$2Na \rightarrow 2Na^+ + 2e^-$

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

Add up:

$2Na + Cl_2 \rightarrow 2NaCl$

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

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

$Cl_2 \rightarrow 2Cl^- - 2e^-$ is **reduction** because electrons are being gained by Cl
0    -1

### Worked example

For $Mg + 2AgNO_3 \rightarrow Mg(NO_3)_2 + 2Ag$, write
a) the ionic equation
b) the two half-equations, identifying which is oxidation and reduction

a) $Mg + 2Ag^+ + \cancel{2NO_3^-} \rightarrow Mg^{2+} + \cancel{2NO_3^-} + 2Ag$

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

$Mg + 2Ag^+ \rightarrow Mg^{2+} + 2Ag$

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

Question 4 provides more practice on half-equations.

### Questions

1. a) What is the oxidation number of Cl in the following?
i) HCl ii) $HClO_3$ iii) $ClO^-$ iv) $Cl_2O_7^{2-}$ v) $ClO_4^-$
b) What is the oxidation number of S in the following?
i) $SO_2$ ii) $Na_2S_2O_3$ iii) $SO_4^{2-}$ iv) $SO_3^{2-}$ v) $SO_3$
c) What is the oxidation number of N in the following?
i) $N_2H_4$ ii) $HNO_3$ iii) $NaNO_2$ iv) HCN v) $N_2O$
d) Find the oxidation number of the identified elements in the following:

| | | |
|---|---|---|
| i) Ba in $BaCl_2$ | ii) Li in $Li_2O$ | iii) P in $P_2O_5$ |
| iv) P in $P_2O_3$ | v) C in CO | vi) I in $I_2$ |
| vii) C in $CCl_4$ | viii) I in $I^-$ | ix) Cr in $CrO_4^{2-}$ |
| x) Br in $BrO_3^-$ | xi) H in LiH | xii) N in $N_2O_4$ |
| xiii) Fe in $Fe(CN)_6^{3-}$ | xiv) Cr in $Cr_2O_7^{2-}$ | xv) P in $PO_4^{3-}$ |
| xvi) O in $H_2O_2$ | xvii) Mn in $MnO_2$ | xviii) Xe in $XeF_4$ |
| xix) N in $NH_3$ | xx) N in $NaNO_3$ | |

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

| | |
|---|---|
| a) $Zn + Ag_2O \rightarrow ZnO + 2Ag$ | b) $2K + Br_2 \rightarrow 2KBr$ |
| c) $PbO_2 + 4H^+ + Sn^{2+} \rightarrow Pb^{2+} + Sn^{4+} + 2H_2O$ | d) $N_2 + 3H_2 \rightarrow 2NH_3$ |
| e) $2K + H_2 \rightarrow 2KH$ | f) $Mg + 2HCl \rightarrow MgCl_2 + H_2$ |
| g) $Sn^{2+} + 2Fe^{3+} \rightarrow Sn^{4+} + 2Fe^{2+}$ | h) $2Cu + O_2 \rightarrow 2CuO$ |
| i) $KIO_3 + 2Na_2S_2O_3 \rightarrow KIO + 2Na_2SO_4$ | j) $Cu^{2+} + Cu \rightarrow 2Cu^+$ |
| k) $2S_2O_3^{2-} + I_2 \rightarrow 2I^- + S_4O_6^{2-}$ | l) $2FeI_2 + I_2 \rightarrow 2FeI_3$ |
| m) $IO_3^- + 5I^- + 6H^+ \rightarrow 3I_2 + 3H_2O$ | n) $I_2 + I^- \rightarrow I_3^-$ |
| o) $Zn + 2HNO_3 \rightarrow Zn(NO_3)_2 + H_2$ | |

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

4. For each of the following equations
i) write the two half equations
ii) identify the half-equations as oxidation or reduction

a) $Zn + 2AgCl \rightarrow ZnCl_2 + 2Ag$
b) $PbO_2 + 4HCl \rightarrow PbCl_2 + Cl_2 + 2H_2O$
c) $2K + H_2 \rightarrow 2KH$
d) $Mg + 2HCl \rightarrow MgCl_2 + H_2$
e) $PH_3 + 2O_2 \rightarrow H_3PO_4$

### Answers

| | | | | | | |
|---|---|---|---|---|---|---|
| 1 | a) | i) –1 | ii) +5 | iii) +1 | iv) +6 | v) +7 |
| | b) | i) +4 | ii) +2 | iii) +6 | iv) +4 | v) +6 |
| | c) | i) –2 | ii) +5 | iii) +3 | iv) –5 | v) +1 |
| | d) | i) +2 | ii) +1 | iii) +5 | iv) +3 | v) +2 |
| | | vi) 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

| | | | | | |
|---|---|---|---|---|---|
| a) Zn | $Ag^+$ | b) K | $Br_2$ | c) $Sn^{2+}$ | $Pb^{4+}$ |
| d) $H_2$ | $N_2$ | e) K | $H_2$ | f) Mg | $H^+$ |
| g) $Sn^{2+}$ | $Fe^{3+}$ | h) Cu | $O_2$ | i) I | S |
| j) Cu | $Cu^{2+}$ | k) S | $I_2$ | l) $Fe^{2+}$ | $I_2$ |
| m) $I^-$ | I | n) $I^-$ | I | o) Zn | $H^+$ |

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

| | | | | | |
|---|---|---|---|---|---|
| a) $Ag_2O$ | Zn | b) $Br_2$ | K | c) $PbO_2$ | $Sn^{2+}$ |
| d) $N_2$ | $H_2$ | e) $H_2$ | K | f) HCl | Mg |
| g) $Fe^{3+}$ | $Sn^{2+}$ | h) $O_2$ | Cu | i) $KIO_3$ | $Na_2SO_3$ |
| j) $IO_3^-$ | $I^-$ | k) $I_2$ | $S_2O_3^{2-}$ | l) $I_2$ | $FeI_2$ |
| m) $Cu^{2+}$ | Cu | n) $I_2$ | $I^-$ | o) $HNO_3$ | Zn |

4.

| | | | | |
|---|---|---|---|---|
| a) | $Zn \rightarrow Zn^{2+} + 2e^-$ | oxidation | $Ag^+ + e^- \rightarrow Ag$ | reduction |
| b) | $2Cl^- \rightarrow Cl_2 + 2e^-$ | oxidation | $Pb^{4+} + 2e^- \rightarrow Pb^{2+}$ | reduction |
| c) | $K \rightarrow K^+ + e^-$ | oxidation | $H_2 \rightarrow 2H^- - 2e^-$ | reduction |
| d) | $Mg \rightarrow Mg^{2+} + 2e^-$ | oxidation | $2H^+ + 2e^- \rightarrow H_2$ | reduction |
| e) | $P^{3-} \rightarrow P (+5) + 2e^-$ | oxidation | $O_2 + 2e^- \rightarrow 2O^{2-}$ | reduction |

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