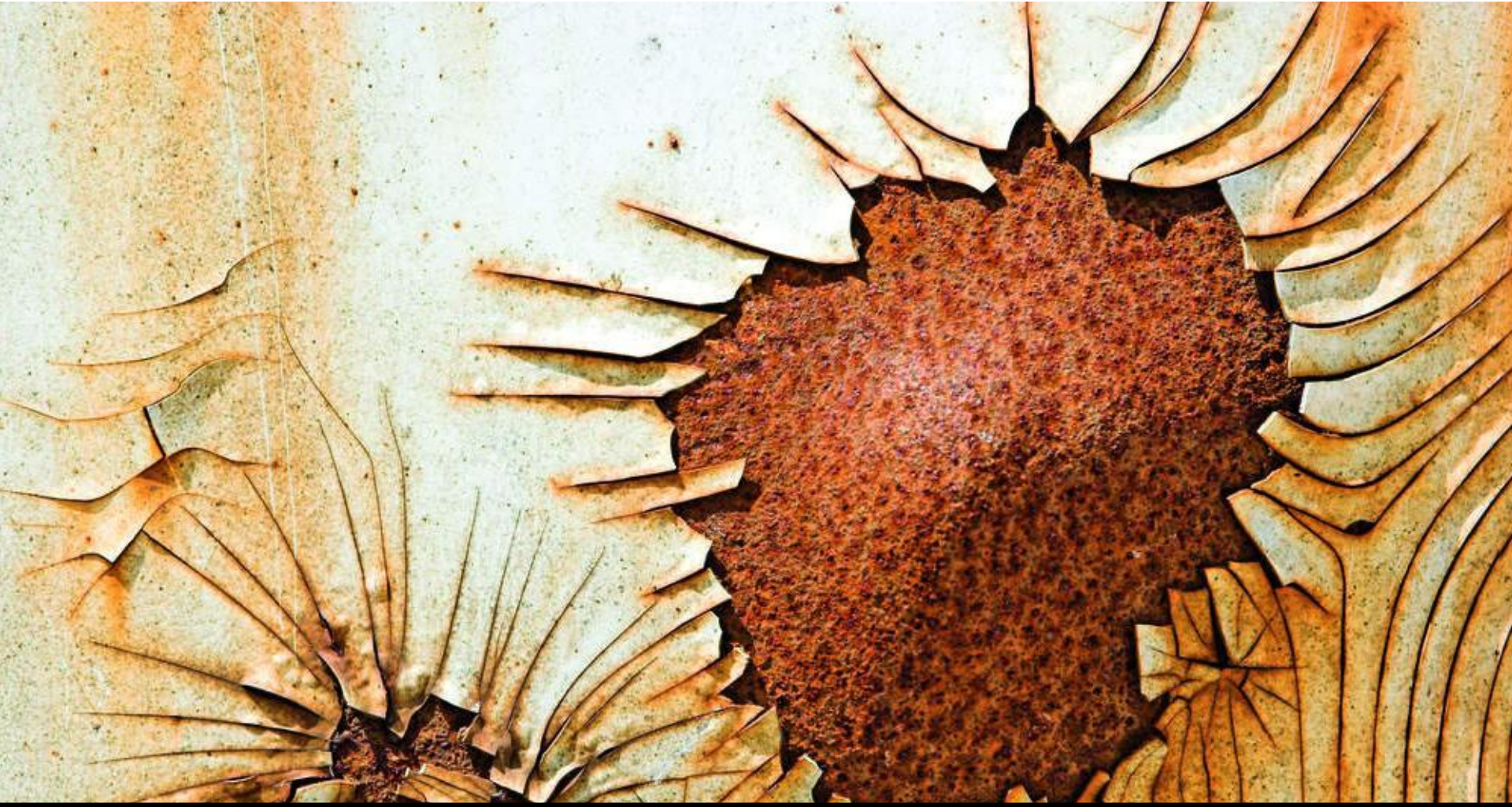

NCEA Chemistry 2.7
Redox AS 91167

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What is this NCEA Achievement Standard?

When a student achieves a standard, they gain a number of credits. Students must achieve a certain number of credits to gain an NCEA certificate (80 for Level 2)

The standard you will be assessed on is called **Chemistry 2.7 Demonstrate understanding of oxidation-reduction**

It will be internally (in Class) assessed as part of a **Examination with a practical component** and will count towards **3 credits** for your Level 2 NCEA in Chemistry

What are the main steps required in this Internal Assessment?

AS91167 Demonstrate understanding of oxidation-reduction

Interpretation of evidence for Achieved

Demonstrate understanding involves:
describing, identifying, naming, giving an account of oxidation-reduction and describing oxidation-reduction reactions. This requires the use of chemistry vocabulary, symbols and conventions.

- ❑ Key reactant and product species correctly identified for reactions

- ❑ Correctly identifies half reactions as reduction or oxidation (oxidation number or electrons)

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Interpretation of evidence for Merit

Demonstrate in-depth understanding involves making and explaining links between oxidation-reduction reactions, observations and equations. This requires explanations that use chemistry vocabulary, symbols and conventions.

- ❑ A balanced half equation is written for one reactions
- ❑ Oxidation and reduction are identified in terms of oxidation number or electron transfer for one reaction
- ❑ Observations are linked to species for one reaction

Aiming for Excellence

Interpretation of evidence for Excellence

Demonstrate comprehensive understanding involves justifying, evaluating, comparing and contrasting, or analysing links between oxidation-reduction reactions, observations and equations. This requires the consistent use of chemistry vocabulary, symbols and conventions.

- ❑ A balanced overall equation is written with no errors
- ❑ Oxidation and reduction are identified in terms of oxidation number or electron transfer for reactions
- ❑ Observations are linked to species for reactions

In this Achievement Standard Oxidation-reduction is limited to:

- oxidation numbers
- electron transfer in reactions
- oxidants and/or reductants
- observations for reactions
- balanced oxidation-reduction half equations
- overall balanced oxidation-reduction equations.

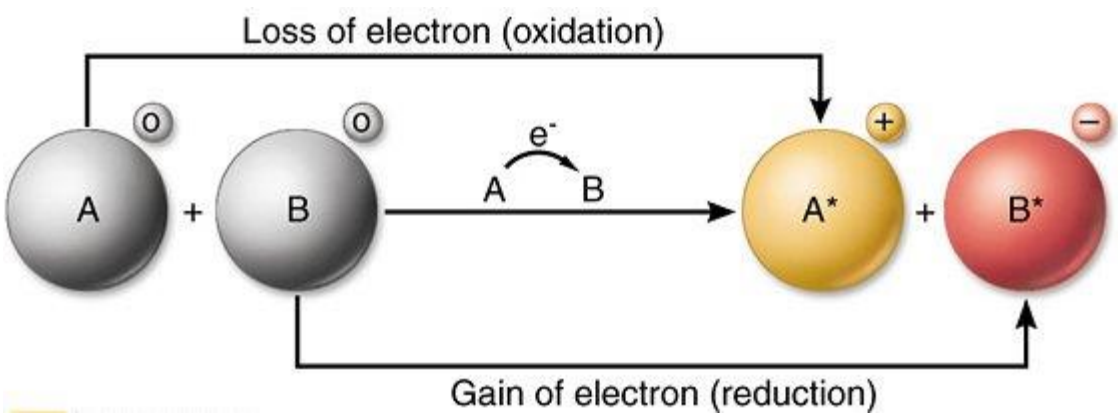
- oxidants include a selection from, but not limited to: O_2, I_2, Br_2, Cl_2, OCl^- H^+,Fe^{3+}, Cu^{2+}, H_2O_2, MnO_4^-/H^+, $Cr_2O_7^{2-}/H^+$, concentrated HNO_3, IO_3^-

- reductants include a selection from, but not limited to, metals, C, H_2, Fe^{2+}, Br^-, I^-, H_2S, SO_2 ,SO_3^{2-}, HSO_3^-, H_2O_2

Chemical Reactions - reactants & products

A **chemical reaction** is a process that produces a chemical change to one or more substances.
A chemical reaction will produce one or more **new substances.** Other observations may include a temperature change, a colour change or production of gas.
Chemicals that are used in a chemical reaction are known as **reactants**. Those that are formed are known as **products**.
Oxidation – Reduction reactions are a specific type of reaction where electrons are transferred.

Reactants → Products



A reactant and what product it changes into after the redox reaction is known as a **species** i.e. Cu changing to Cu^{2+} so Cu/Cu^{2+} is the species

Redox terms

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Reduction and oxidation occur in pairs of reactants

* A redox reaction is where one reactant is oxidised and the other reactant is reduced.

Oxidation of one reactant	Reduction of the other reactant
❑ loss of electrons and a	❑ gain of electrons
❑ loss of hydrogen and a	❑ gain of hydrogen
❑ gain of oxygen and a	❑ loss of oxygen

Oxidation numbers are used to determine what is oxidised and what is reduced in a reaction. These will be explained later

Electron transfer

Electron transfer is most common type of Redox reaction – and the one we will be looking at in this Internal

An Iron nail left in copper sulfate

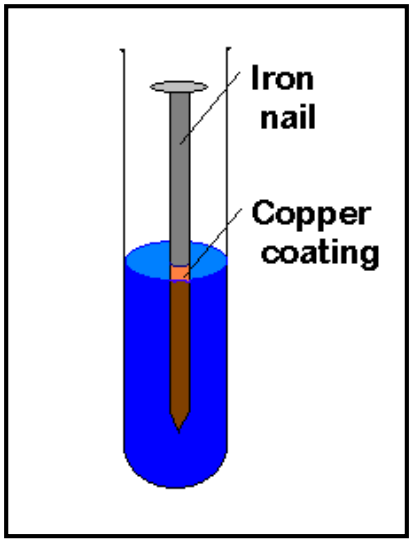
Copper is reduced – gained electrons

Oxidising agent (oxidant)

$$Fe_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$$

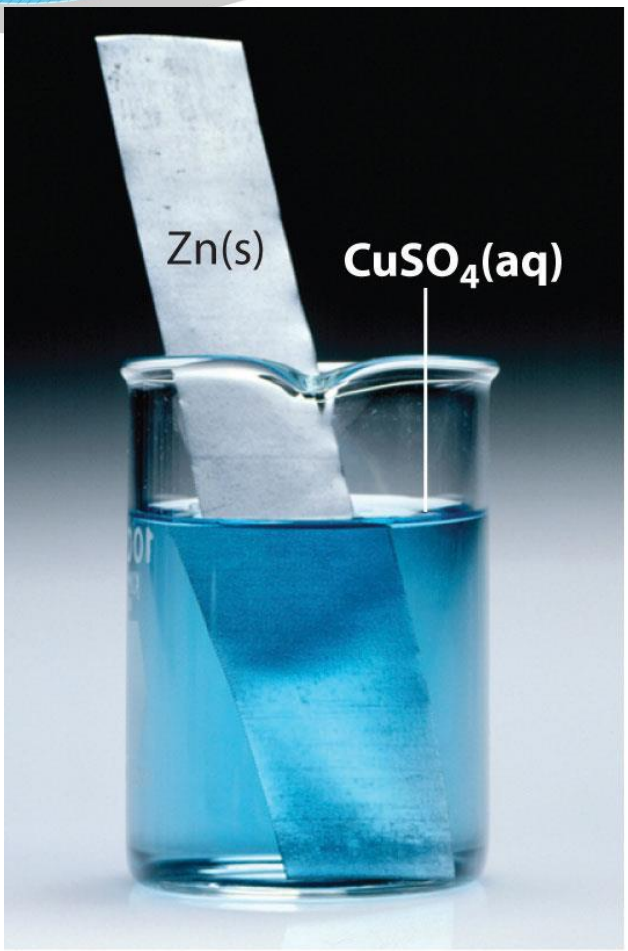
Iron is oxidised – lost electrons

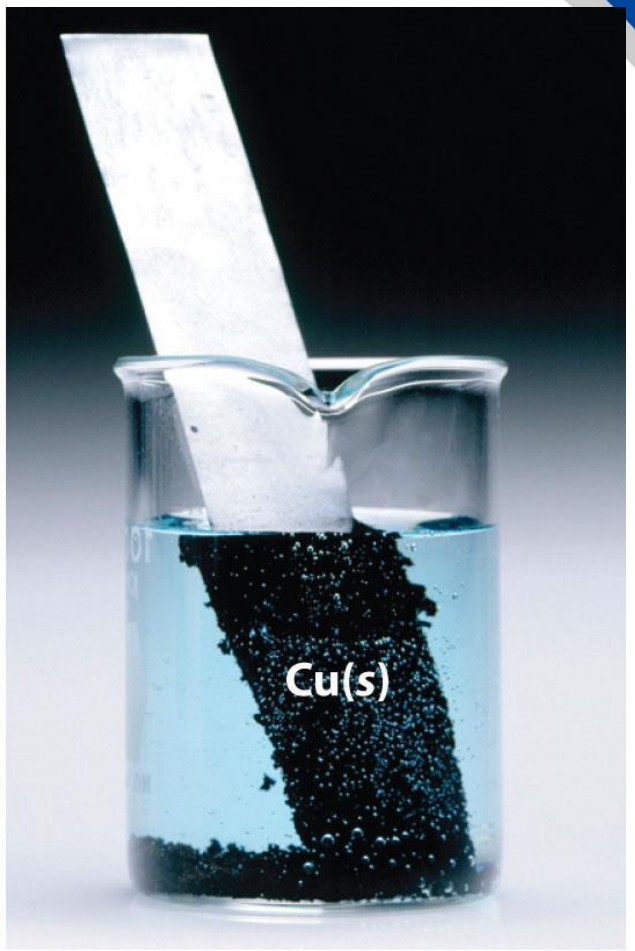
Reducing Agent (reductant)



Electron transfer

During electron transfer Redox reactions we often just write **ionic equations**.
For example the Cu^{2+} ions come from the $CuSO_4$ but only the Cu^{2+} is written into the equation. The SO_4^{2-} ions are **spectators** as they **play no part in the reaction**. They are also in solution and detached from the Cu^{2+} ions





$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

Iron Ore smelting

Iron oxide is reduced – lost oxygen

Oxidising agent (oxidant)

$$2Fe_2O_{3(s)} + 3C_{(s)} \longrightarrow 4Fe_{(s)} + 3CO_{2(g)}$$

carbon is oxidised – gained oxygen

Reducing Agent (reductant)

Iron Ore

Sulfur production

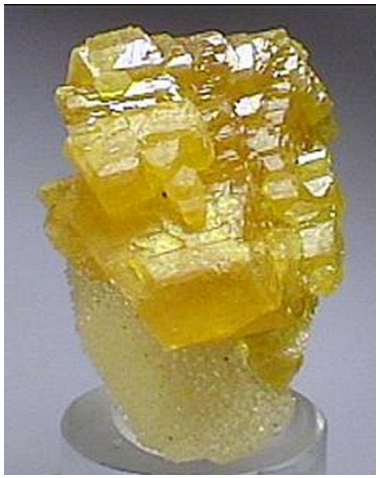
Hydrogen sulphide is oxidised – lost hydrogen

Reducing Agent (reductant)

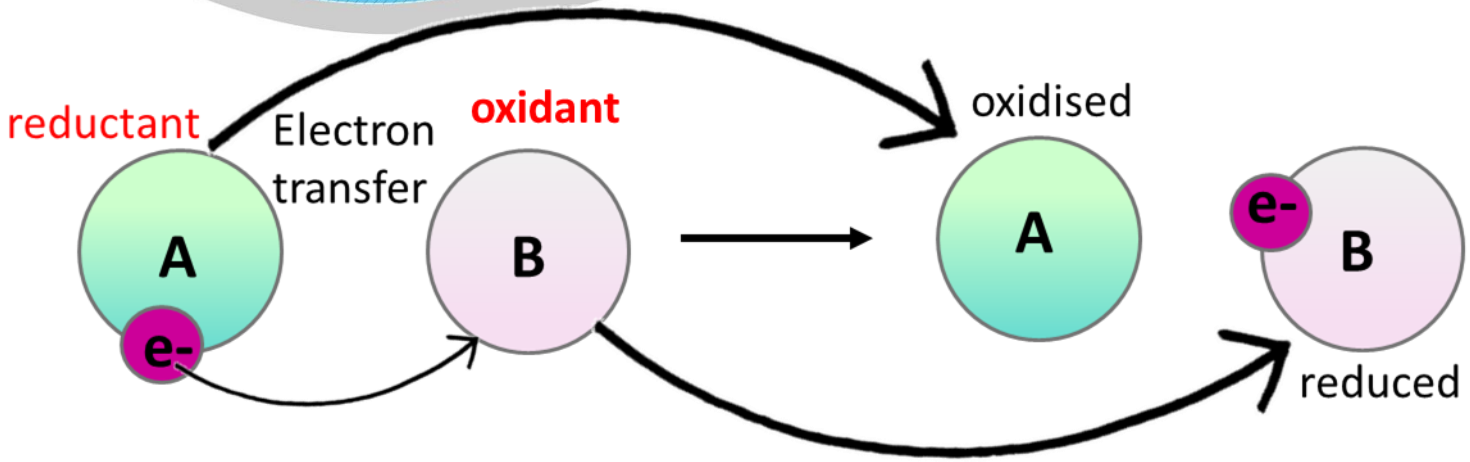
$$2H_2S_{(g)} + O_{2(g)} \longrightarrow 2S_{(s)} + 2H_2O_{(l)}$$

Oxygen gas is reduced – gained hydrogen

Oxidising agent (oxidant)



Summary of Terms



LEO (loss electrons oxidation) A	GER (gain electrons reduction) B
Reductant Acts as a reducing agent to B ❑ is oxidised ❑ loses electrons	Oxidant Acts as an oxidising agent to A ❑ is reduced ❑ gains electrons

Oxidants (reduced)			Reductants (oxidised)		
$O_{2(g)}$ Oxygen	→	$O^{2-}_{(aq)}$ Oxide	$C_{(s)}$ Carbon	→	$CO_{2(g)}$ Carbon Dioxide
$IO_3^-{}_{(aq)}$ Iodate	→	$I_{2(s)}$ Iodine	$SO_3^{2-}{}_{(aq)}$ Sulfite ion	→	$SO_4^{2-}{}_{(aq)}$ Sulfate ion
$OCl^-{}_{(aq)}$ Hypochlorite	→	$Cl^-{}_{(aq)}$ Chloride	$H_{2(g)}$ Hydrogen gas	→	$H^+{}_{(aq)}$ Hydrogen ion
Conc $HNO_{3\,(aq)}$ Nitric acid	→	$NO_{2(g)}$ nitrous dioxide	$Fe^{2+}{}_{(aq)}$ Iron (ii) ion	→	$Fe^{3+}{}_{(aq)}$ Iron (iii) ion
$H_2O_{2\,(aq)}$ Hydrogen peroxide	→	$H_2O_{(l)}$ Water	$I^-{}_{(aq)}$ Iodide ion	→	$I_{2(s)}$ Iodine
$MnO_4^-/H^+{}_{(aq)}$ Permanganate	→	$Mn^{2+}{}_{(aq)}$ Manganese ion	$H_2O_{2(l)}$ Hydrogen peroxide	→	$O_{2(g)}$ oxygen gas
$Cr_2O_7^{2-}/H^+{}_{(aq)}$ Dichromate	→	$Cr^{3+}{}_{(aq)}$ Chromium ion	$Br^-{}_{(aq)}$ Bromide	→	$Br_{2(l)}$ Bromine
$S_{(s)}$ sulfur	→	$H_2S_{(g)}$ Hydrogen sulfide	$SO_{2(g)}$ Sulfur dioxide	→	$SO_4^{2-}{}_{(aq)}$ Sulfate ion

Redox couples

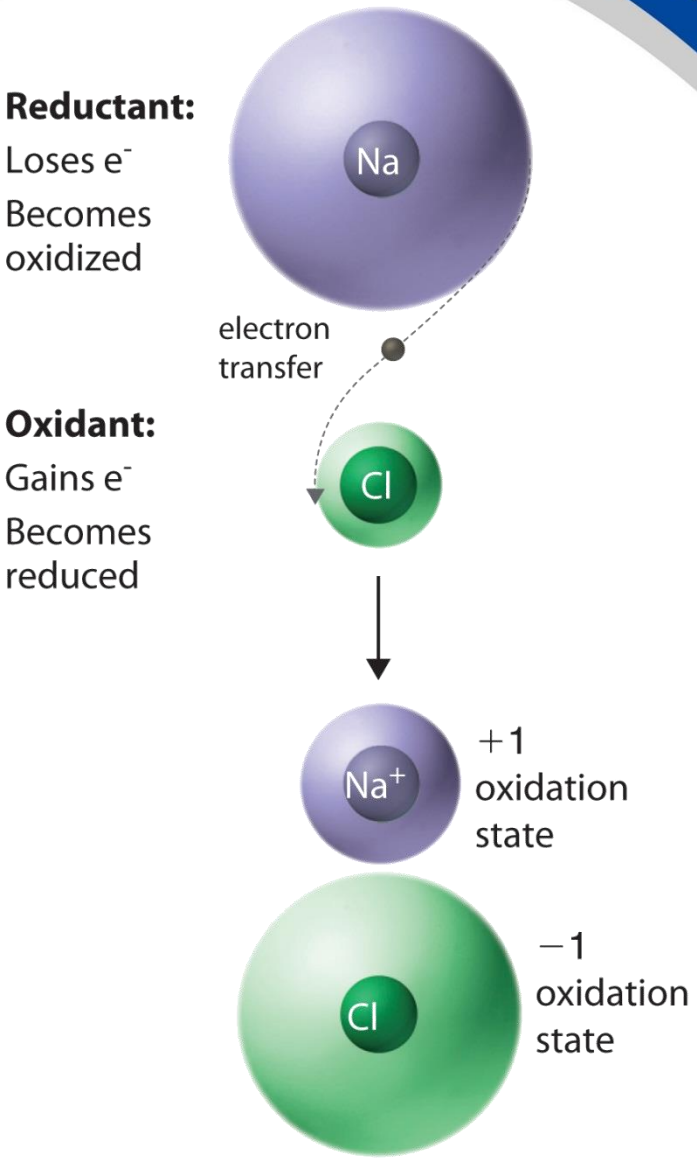
Oxidants (reduced)	Reductants (oxidised)
O_2 / O^{2-}	C / CO_2
IO_3^- / I_2 (brown)	SO_3^{2-} / SO_4^{2-}
OCl^- / Cl-	H_2 / H^+
Conc HNO_3 / NO_2 (brown)	Fe^{2+} (green) / Fe^{3+} (orange rust)
H_2O_2 / H_2O	I^- / I_2(brown)
MnO_4^-/H^+ (purple) / Mn^{2+} (colourless)	H_2O_2 / O_2
$Cr_2O_7^{2-}/H^+$ (orange)/ Cr^{3+}(green)	Br^- / Br_2 (brown liquid)
Fe^{3+} (orange rust)/ Fe^{2+} (green)	SO_2 / SO_4^{-2}
Cu^{2+} (blue)/ Cu(orange solid)	H_2S / S(yellow solid)
I_2 (brown) / I^-	HSO_3^- / SO_4^{2-}
H^+ / H_2	
Cl_2 (pale green gas)/ Cl-	
Br_2 (brown liquid)/ Br^-	

Observations of Oxidant and reductant colours

To judge the colours of oxidants and reductants requires observation before and after a redox reaction has taken place. It is a good idea to make precise notes on the colour rather than just state orange or green - as in the case of dichromate and the chromium ions

Oxidation Numbers

Oxidation numbers can be used to predict whether a species – the reactant and its product – are undergoing oxidation or reduction. The oxidation number is assigned to a **single atom only** and the corresponding atom in the product using a **set of rules**.
If the oxidation number **increases** from reactant to product then **oxidation** has taken place. If the oxidation number **decreases** from reactant to product then **reduction** has taken place.



Oxidation Numbers and Rules

The Oxidation Number (ON) gives the 'degree' of oxidation or reduction of an element.
They are assigned to a **INDIVIDUAL** atom using the following rules.

Elements

Oxidation number = 0

For example

Fe H_2

0 0

Hydrogen atom

(not as element)

Oxidation number = +1

For example

$\underline{H}Cl$ $\underline{H}_2SO_4$

+1 +1

Except Hydrides

Oxidation number = -1

For example $Li\underline{H}$

-1

Oxygen atom

(not as element)

Oxidation number = -2

For example

$Mn\underline{O}_4^-$ $C\underline{O}_2$

-2 -2

Except peroxides

Oxidation number = -1

for example $H_2\underline{O}_2$

-1

Oxidation Numbers and Rules

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Monatomic ions

Oxidation number = charge

For example

Fe^{2+} Cl^{-}

+2 -1

Polyatomic ions

Sum of Oxidation number = charge

For example

MnO_4^{-}

+7 -2

Because
Total charge = -1
And
Oxygen = -2

$+7 + (4x -2) = -1$

Mn O

Molecules

Sum of Oxidation number = 0

For example

CO_2

+4 -2

Because
Total charge = 0
And
Oxygen = -2

$+4 + (2x-2) = 0$

C O

Oxidation Number Summary

Oxidation is a loss of electrons and causes an **increase** in ON

Reduction is a gain of electrons and causes an **decrease** in ON

Oxidation of Fe^{2+}

$Fe^{2+} \longrightarrow Fe^{3+} + e^-$

+2 +3

Fe has increased ON

(+2 to +3) caused by

a loss of electrons e-

Reduction of MnO_4^-

$MnO_4^- + 5e^- \longrightarrow Mn^{2+}$

+7 +2

Mn has decreased ON

(+7 to +2) caused by

a gain in electrons e-

OXIDATION and **RED**UCTION always occur together. The electrons lost by one atom are gained by another atom.

This is called a **REDOX** reaction.

Using Oxidation numbers to identify oxidants and reductants

$$Cr_2O_7^{2-} + I^- \rightarrow Cr^{3+} + I_2$$

<u>What has been oxidised and what has been reduced?</u>

STEP ONE – write the ON for each atom using rules (not oxygen or hydrogen)

$$\underline{Cr}_2O_7^{2-} + I^- \rightarrow Cr^{3+} + I_2$$

+6 -1 +3 0

Decrease - reduction Increase - oxidation

STEP TWO – Identify the atom that has had its ON increased. It is **Oxidised**

I- has increased **ON** (-1 to 0) so *I-* is *Oxidised. (the reductant)*

STEP THREE – Identify the atom that has decreased ON. It is **reduced**.

Cr has decreased **ON** (+6 to +3) so $Cr_2O_7^{2-}$ is *Reduced.(the oxidant)*

Balancing Redox equations

A balanced redox equation is broken into two half-equations, to show how electrons are transferred.

$$Fe_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$$

Reduction half equation - oxidant is reduced

$$Fe \longrightarrow Fe^{2+} + 2e^{-}$$

Oxidation half equation – reductant is oxidised

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

Steps to Balancing Redox equations

Rules e.g. $Cr_2O_7^{2-} \rightarrow Cr^{3+}$

1. Assign oxidation numbers and identify element oxidised or reduced.

$$\overset{(+6)(-2)}{Cr_2O_7^{2-}} \rightarrow \overset{(+3)}{Cr^{3+}}$$

2. Balance atom no. for element oxidised or reduced (other than oxygen and hydrogen)

$$Cr_2O_7^{2-} \rightarrow \mathbf{2}Cr^{3+}$$

3. Balance the Oxygen using H_2O

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + \mathbf{7H_2O}$$

4. Use H^+ (acidic conditions) to balance the hydrogen

$$\mathbf{14H^+} + Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

5. Balance charge by adding electrons (LHS on oxidants RHS on reductants)

$$14H^+ + Cr_2O_7^{2-} + \mathbf{6e^-} \rightarrow 2Cr^{3+} + 7H_2O$$

6. Check balance of elements and charges.

Steps to Balancing Redox equations – Example ONE

1. Write half equation by identifying reactant and product	2. Balance atoms that are not O or H	3. Balance O by adding H_2O and H by adding H^+	4. Balance charge by adding electrons
$MnO_4^- \longrightarrow Mn^{2+}$	Atoms already balanced	$MnO_4^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O$ Balance O by adding $4H_2O$ and H by adding 8 H+	$MnO_4^- + 8H^+$ Total charge +7 $\longrightarrow$ $Mn^{2+} + 4H_2O$ Total charge +2 Add 5 electrons (e-)

Balance the half equation for the reduction of MnO_4^- to Mn^{2+}

$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$$

Steps to Balancing Redox equations – Example TWO

1. Write half equation by identifying reactant and product	2. Balance atoms that are not O or H	3. Balance O by adding H_2O and H by adding H^+	4. Balance charge by adding electrons

Balance the half equation for the oxidation of Fe^{2+} to Fe^{3+}

$Fe^{2+} \longrightarrow Fe^{3+}$	Atoms already balanced	There are no O or H atoms to balance	$Fe^{2+} \longrightarrow Fe^{3+} + e^-$ 2+ = 3+ (-1) 1 electron to balance charge

$$Fe^{2+} \longrightarrow Fe^{3+} + e^-$$

Joining half equations together

Rules e.g. $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

And $Fe^{2+} \rightarrow Fe^{3+} + e^-$

1. The two half equations must have electrons on <u>opposite</u> sides of the equation
2. Place the two equations one under the other
3. The electron numbers must equal each other – if not multiply one or both equations to the lowest common denominator (multiply every reactant/product)

$$5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$$

4. Cancel out the electrons

$$MnO_4^- + 8H^+ + \cancel{5e^-} \rightarrow Mn^{2+} + 4H_2O$$
$$5Fe^{2+} \rightarrow 5Fe^{3+} + \cancel{5e^-}$$

5. Cancel out the same number of H^+ and H_2O if present on both sides
6. Join the remainder together

$$\boxed{MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}}$$

Observations

Observations must **link** the species (the reactant and the product it changes into) to the colour changes and/or appearance of gas.
Question: Mix potassium dichromate solution and iron(ii) sulfate solution

When potassium dichromate solution is mixed with iron (ii) sulfate solution the orange solution changes to a green colour.

1. Identify the reactants from the question and write two half equations

   $Cr_2O_7^{2-} \rightarrow Cr^{3+}$ The reactants are given but you will have to remember

   $Fe^{3+} \rightarrow Fe^{2+}$ the products

2. Use oxidation numbers to identify which species have been oxidised and which species has been reduced

3. Write colours beside/underneath each reactant/product of each species – these will be collected from observation of the reaction (or memory/question)

$Cr_2O_7^{2-}$ (bright orange) Cr^{3+} (green) Fe^{3+} (rust orange) Fe^{2+} (pale green)

4. Write a comprehensive summary of this information

Orange dichromate ion, $Cr_2O_7^{2-}$ is reduced to green chromium ion, Cr^{3+} and the rust orange Fe^{2+} ion is oxidised to pale green Fe^{3+} ion, so over all, the colour is from an **orange solution to a green solution**

Observations

Question: Mix potassium dichromate solution and iron(ii) sulfate solution

1. Oxidation numbers

$\overset{+6}{Cr_2O_7^{2-}} \rightarrow \overset{+3}{Cr^{3+}}$ the oxidation number of dichromate reduces from +6 to +3 of the chromium ion, therefore this reaction is a **REDUCTION** reaction (and $Cr_2O_7^{2-}$ is the oxidant)

$\overset{+2}{Fe^{2+}} \rightarrow \overset{+3}{Fe^{3+}}$ the oxidation number iron (ii) ion increases from +2 to the +3 of iron (iii) ion, therefore this reaction is an **OXIDATION** reaction (and Fe^{2+} is the reductant)

2. Balancing

$14H^+ + Cr_2O_7^{2-} + \mathbf{6e^-} \rightarrow 2Cr^{3+} + 7H_2O$

$(Fe^{2+} \rightarrow Fe^{3+} + e^-) \times 6$

$14H^+ + Cr_2O_7^{2-} + \mathbf{6Fe^{2+}} \rightarrow 2Cr^{3+} + 7H_2O + \mathbf{6Fe^{3+}}$

$Cr_2O_7^{2-}$ gains 6 electrons therefore this is a reduction reaction and $Cr_2O_7^{2-}$ is the oxidant
Fe^{2+} loses 1 electron therefore this is an oxidation reaction and Fe^{2+} is the reductant

3. Observations

Orange dichromate ion, $Cr_2O_7^{2-}$ is reduced to green chromium (iii) ion, Cr^{3+} and the rust orange Fe^{2+} ion is oxidised to pale green Fe^{3+} ion.

Observations

Sometimes two products will both have a distinct colour and mixed together produce a colour that favours one more than the other. Here the pale green of the Cr^{3+} ion is masked by the strong colour of the brown I_2. This is where **actual observation** of the reaction is **important** and **notes are made at the time**.