Chemistry AS 91167
C2.7 Redox

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)

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

Standard of 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:

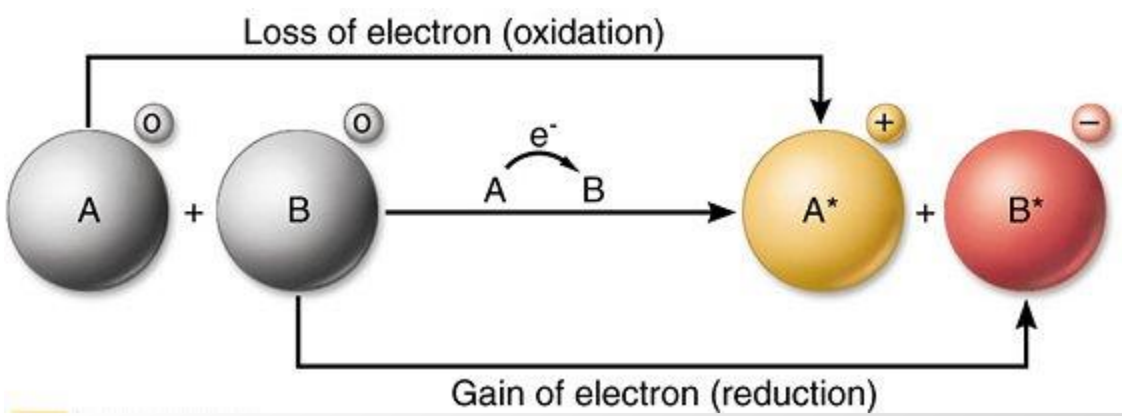
- oxidation numbers
- electron transfer in reactions
- oxidants and/or reductants
- observations for reactions
- balanced oxidation-reduction half equations
- overall balanced oxidation-reduction equations.

- 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

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

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

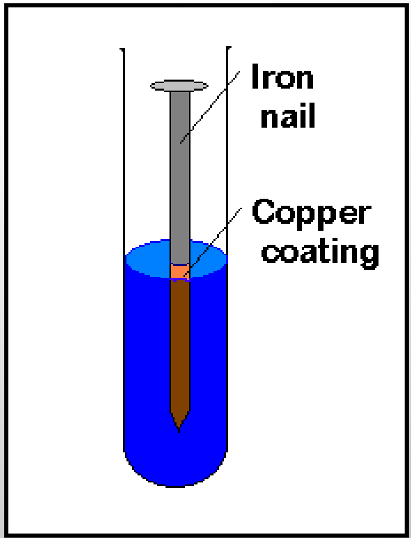
Copper is reduced – gained electrons

Oxidising agent (oxidant)

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

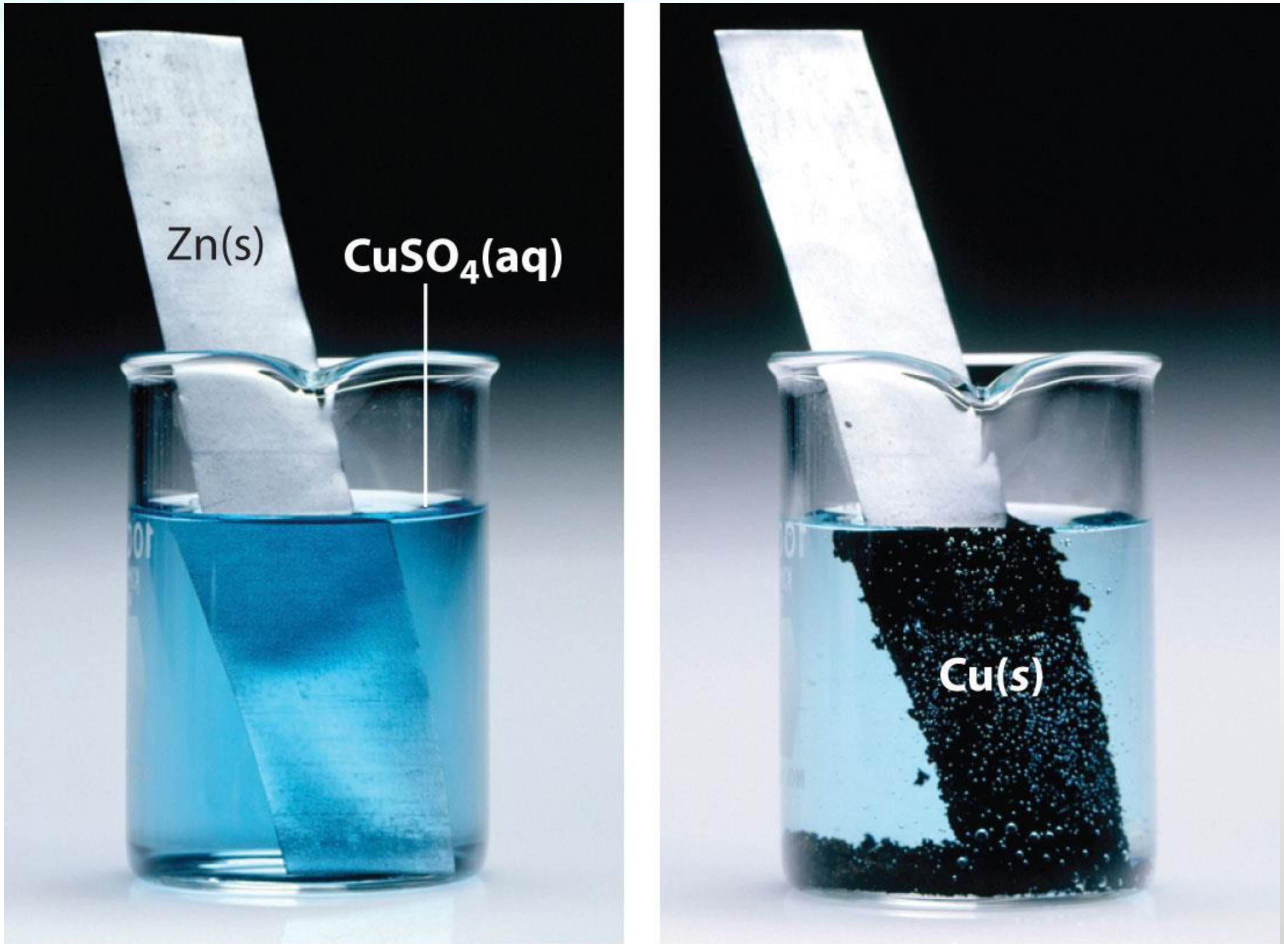
Iron is oxidised – lost electrons

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



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

Oxygen transfer

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

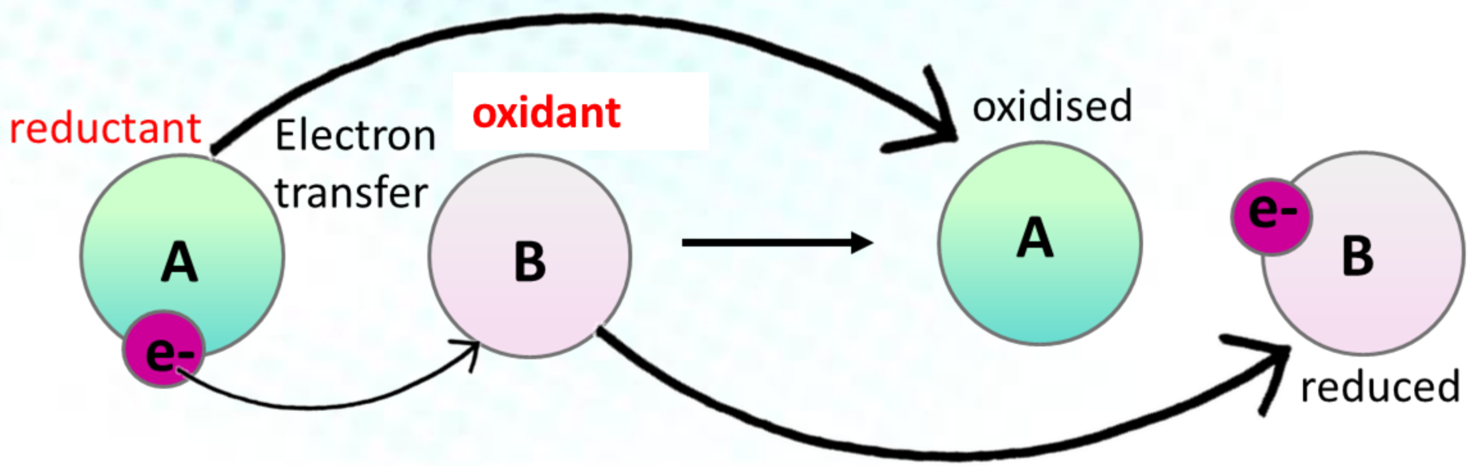
Reducing Agent (reductant)

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

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^-	

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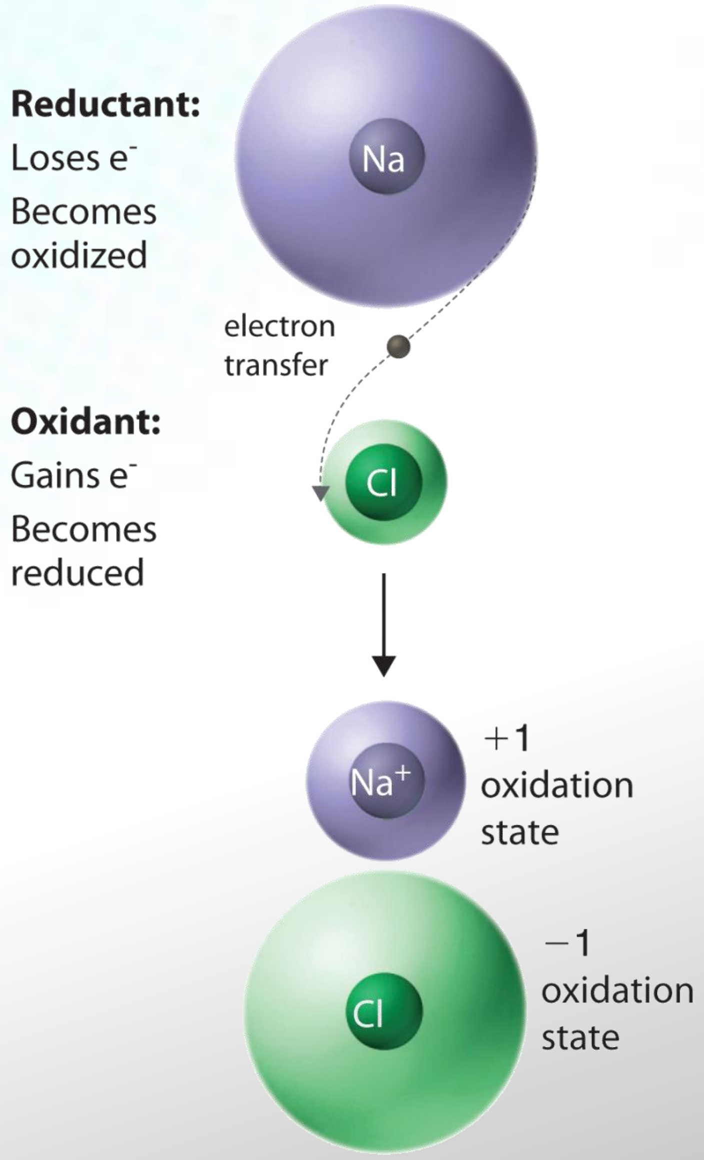
Observations of Oxidant and reductant colours

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	Hydrogen atom (not as element)	Oxygen atom (not as element)
Oxidation number = 0	Oxidation number = +1	Oxidation number = -2
For example	For example	For example
Fe: 0 H_2: 0	$\underline{H}Cl$: +1 $\underline{H}_2SO_4$: +1	$Mn\underline{O}_4^-$: -2 $C\underline{O}_2$: -2
	Except Hydrides Oxidation number = -1 For example $Li\underline{H}$: -1	Except peroxides Oxidation number = -1 for example $H_2\underline{O}_2$: -1

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Oxidation Numbers and Rules

Monatomic ions	Polyatomic ions	Molecules
Oxidation number = charge	Sum of Oxidation number = charge	Sum of Oxidation number = 0
For example	For example	For example
Fe^{2+} Cl^{-}	MnO_4^-	CO_2
+2 -1	+7 -2	+4 -2
	Because Total charge = -1 And Oxygen = -2	Because Total charge = 0 And Oxygen = -2
	$+7 + (4x\ -2) = -1$	$+4 + (2x-2) = 0$
	Mn O	C O

Oxidation Number Summary

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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.

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

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

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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 reduction of MnO_4^- to Mn^{2+}

$MnO_4^- \rightarrow Mn^{2+}$	Atoms already balanced	$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$ Balance O by adding $4H_2O$ and H by adding 8 H+	$MnO_4^- + 8H^+$ Total charge +7 $\rightarrow$ $Mn^{2+} + 4H_2O$ Total charge +2 Add 5 electrons (e-)

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

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

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

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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**.