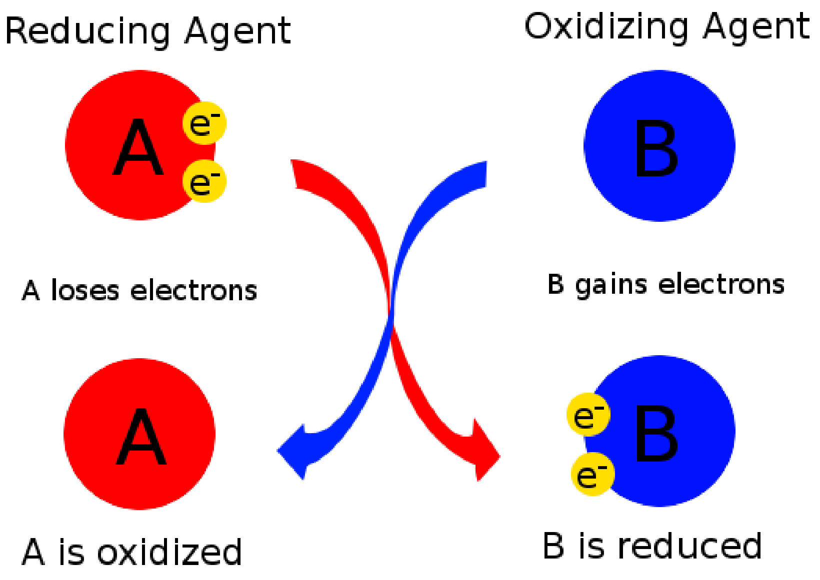
UNIT 6

REDOX REACTIONS

PART 1 – OXIDATION AND REDUCTION



Contents

Key words: oxidation, reduction, oxidation number, oxoanion, oxidising agent, reducing agent, half-equation, redox reaction, rusting, galvanisation

Units which must be completed before this unit can be attempted:
Unit 1 – Atomic Structure and the Periodic Table
Unit 2 – Particles, Structure and Bonding
Unit 3 – Amount of Substance
Unit 4 – Introduction to Physical Chemistry

1) Oxidation and Reduction

(a) Definitions of Oxidation and Reduction

(i) In terms of electron transfer

Oxidation and reduction can be defined in terms of electron transfer:

Oxidation is the loss of electrons. When a species loses electrons it is said to be oxidised.
Eg $Na \rightarrow Na^+ + e^-$ (each sodium atom loses one electron)
$2I^- \rightarrow I_2 + 2e^-$ (each iodide ion loses one electron, so two in total)

Reduction is the gain of electrons. When a species gains electrons it is said to be reduced.
Eg $Cl_2 + 2e^- \rightarrow 2Cl^-$ (each chlorine atom gains one electron, so two in total)
$Al^{3+} + 3e^- \rightarrow Al$ (each aluminium ion gains three electrons)

Processes which show the gain or loss of electrons by a species are known as **half-equations**. They show simple oxidation or reduction processes.

(ii) In terms of the transfer of hydrogen or oxygen

In organic chemistry, it is sometimes more convenient to consider oxidation and reduction in terms of the gain and loss of oxygen and hydrogen:

Oxidation is the gain of oxygen or the loss of hydrogen:

When a species gains oxygen or loses hydrogen, it is said to be oxidised:
Eg $C_2H_6O + O \rightarrow C_2H_4O + H_2O$ (ethanol loses two hydrogen atoms when it is oxidised to ethanal)
ethanol ethanal
Eg $C_2H_4O + O \rightarrow C_2H_4O_2$ (ethanal gains one oxygen atom when it is oxidised to ethanoic acid)
ethanol ethanal

Reduction is the gain of hydrogen or the loss of oxygen:
Eg $C_2H_3N + 2H \rightarrow C_2H_5N$ (ethanenitrile gains two hydrogen atoms when it is reduced to ethylamine)
ethanol ethanal
Eg $C_6H_5NO_2 + 6H \rightarrow C_6H_7N + 2H_2O$ (nitrobenzene loses two oxygen atoms and gains two hydrogen
nitrobenzene phenylamine atoms one oxygen atom when it is reduced to phenylamine)

In fact, these reactions can also be described as oxidation or reduction in terms of the carbon atom gaining or losing electrons, but it is simpler to consider the movement of oxygen and hydrogen in these cases.

(b) Oxidation numbers

The oxidation number of an atom is the charge that would exist on an individual atom if the bonding were completely ionic.

In simple ions, the oxidation number of the atom is the charge on the ion:

- Na^+, K^+, H^+ all have an oxidation number of +1.
- Mg^{2+}, Ca^{2+}, Pb^{2+} all have an oxidation number of +2.
- Cl^-, Br^-, I^- all have an oxidation number of -1.
- O^{2-}, S^{2-} all have an oxidation number of -2.

In molecules or compounds, the sum of the oxidation numbers on the atoms is zero.

- SO_3; oxidation number of S = +6, oxidation number of each O = -2 $\quad +6 + 3(-2) = 0$
- H_2O_2; oxidation number of H = +1, oxidation number of O = -1 $\quad 2(+1) + 2(-1) = 0$

SCl_2; oxidation number of S = +2, oxidation number of Cl = -1 $\quad +2 + 2(-1) = 0$

In polyatomic ions, the sum of the oxidation numbers on the atoms is equal to the overall charge on the ion.

- SO_4^{2-}; oxidation number of S = +6, oxidation number of O = -2 $\quad +6 + 4(-2) = -2$
- PO_4^{3-}; oxidation number of P = +5, oxidation number of O = -2 $\quad +5 + 4(-2) = -3$
- ClO^-; oxidation number of Cl = +1, oxidation number of O = -2 $\quad +1 +(-2) = -1$

In elements, the oxidation number of each atom is zero.

- In Cl_2, S, Na and O_2 all atoms have an oxidation number of zero.

Many atoms, such as S, N and Cl, can exist in a variety of oxidation states. The oxidation number of these atoms can be calculated by assuming that the oxidation number of the other atom is fixed (usually O at -2).

All group I atoms always adopt the +1 oxidation state in their compounds.
All group II atoms adopt the +2 oxidation state in their compounds.
Aluminium always adopts the +3 oxidation state in its compounds.
Fluorine always adopts the -1 oxidation state in its compounds.
Hydrogen adopts the +1 oxidation state in its compounds unless it is bonded to a metal, Silicon or boron in which case it adopts the -1 oxidation state.
Oxygen adopts the -2 oxidation state in its compounds unless it is bonded to a group I or group II metal or hydrogen (with which it sometimes adopts the -1 oxidation state), or with fluorine (with which it adopts the +2 oxidation state).

The oxidation numbers of all other atoms in their compounds can vary.
By following the above guidelines, the oxidation number of any atom in a compound or ion can be deduced.

Test Your Progress: Topic 6 Exercise 1
Deduce the oxidation numbers of the following atoms:

(a) Si in SiF_4	(h) S in $S_2O_3^{2-}$	(o) I in IO_3^-
(b) S in H_2S	(i) Cl in ClO^-	(p) Cl in Cl_2O_7
(c) Pb in PbO_2	(j) Cl in ClO_3^-	(q) O in OF_2
(d) S in H_2SO_4	(k) Tl in TlCl	(r) Fe in Fe_3O_4
(e) N in NO_3^-	(l) C in CaC_2	(s) S in $S_4O_6^{2-}$
(f) N in NO_2^-	(m) H in AlH_3	(t) C in HCN
(g) I in I_2	(n) C in $C_2O_4^{2-}$	

During oxidation and reduction, the oxidation numbers of atoms change.

If an atom is oxidized, its oxidation number increases
(ie it becomes more +ve or less –ve)
If an atom is reduced, its oxidation number decreases
(ie it becomes less +ve or more –ve)

These ideas can be summarized in the following table:

Oxidation	Loss of electrons	Increase in oxidation number
Reduction	Gain of electrons	Decrease in oxidation number

(c) Naming inorganic ions and compounds

Oxidation numbers are used when naming compounds according to the internationally agreed IUPAC rules:

(i) Binary compounds

- Ionic compounds are named by stating the cation followed by the anion
- Binary covalent compounds are named by stating the atom with a positive oxidation number followed by the atom with a negative oxidation number

Simple cations (and atoms in a positive oxidation state in binary covalent compounds) are named using the name of the atom followed by its oxidation number in brackets and Roman numerals:

K^+ is potassium (I) and Cu^+ is copper (I)	+1 = I
Mg^{2+} is magnesium (II) and Cu^{2+} is copper (II)	+2 = II
Al^{3+} is aluminium (III) and Fe^{3+} is iron (III)	+3 = III
C in CO_2 is carbon (IV)	+4 = IV
P in P_4O_{10} is phosphorus (V)	+5 = V
W in WO_3 is tungsten (VI)	+6 = VI
Mn in Mn_2O_7 is manganese (VII)	+7 = VII

Simple anions (and atoms in a negative oxidation state in binary covalent compounds) are named by changing the final one or two syllables of the atom to -ide
Cl^- is chloride, H^- is hydride
O^{2-} is oxide, S^{2-} is sulphide
N^{3-} is nitride, P^{3-} is phosphide

Examples: CO_2 is carbon (IV) oxide (although it is usually known as carbon dioxide)
H_2O is hydrogen (I) oxide (although it is usually known as water)
CuCl is copper (I) chloride
$CuCl_2$ is copper (II) chloride

Test Your Progress: Topic 6 Exercise 2
Deduce the IUPAC names of the following binary compounds:

(a) CuO	(h) PbS	(o) Fe_2O_3
(b) Cu_2O	(i) NH_3	(p) MnO_2
(c) PbO_2	(j) OF_2	(q) CCl_4
(d) Ca_3N_2	(k) H_2Te	(r) Cl_2O_7
(e) CO	(l) UF_6	(s) $SnCl_4$
(f) NO_2	(m) AlH_3	(t) $SnCl_2$
(g) NCl_3	(n) FeO	(u) SO_2

Note: many of these compounds have more commonly names. In particular, it is common to leave out the Roman numeral if the atom has only one known oxidation number (such as sodium or magnesium).
Sodium (I) chloride is the IUPAC name for NaCl but sodium chloride is correct!
Magnesium (II) chloride is the IUPAC name for $MgCl_2$ but magnesium chloride and magnesium dichloride are correct!

(ii) Polyatomic anions and non-binary compounds

In anions containing more than one atom, such as CO_3^{2-}, NO_3^- and SO_4^{2-}, one of the atoms has a positive oxidation number and the other, usually oxygen, has a negative oxidation number.

These anions are named by changing the last one or two syllables of the atom with a positive oxidation number to -ate, and then adding the oxidation state of that atom in brackets and Roman numerals:

Sometimes the presence of oxygen is indicated by using the prefix oxo-, preceded by the number of oxygen atoms if more than one (two = di, three = tri, four = tetra, five = penta, six = hexa)

CO_3^{2-} is called carbonate (IV) or trioxocarbonate (IV)
NO_3^- is called nitrate (V) or trioxonitrate (V)
SO_4^{2-} is called sulphate (VI) or tetraoxosulphate (VI)

Note: the oxo prefix is usually omitted in common usage.

If the atom with a negative oxidation number is not oxygen, is must be mentioned in the prefix, with its final one or two syllables replaced with -o:
$CoCl_4^{2-}$ is called tetrachlorocobaltate (II)

Test Your Progress: Topic 6 Exercise 3
Deduce the IUPAC names of the following compounds:

(a) Na_2SO_3	(f) $NaNO_2$
(b) $KClO_3$	(g) $(NH_4)_2SO_4$
(c) $NaClO$	(h) $KMnO_4$
(d) $CuCO_3$	(i) K_2CrO_4
(e) $Mg(NO_3)_2$	(j) K_2PtF_6

(d) Half-equations

Many oxidation and reduction processes involve polyatomic ions or molecules and the half-equations for these processes are more complex.

There are two ways to create half-equations:

Method 1: (this shows you straight away whether oxidation or reduction is taking place)

- Identify the atom being oxidised or reduced, and make sure there are the same number of that atom on both sides (by balancing)
- insert the number of electrons being gained or lost:
  (on the left if reduction, on the right if oxidation)
  No of electrons gained/lost =
  change in oxidation number x number of atoms changing oxidation number
- balance O atoms by adding H_2O
- balance H atoms by adding H^+

Example: Write a balanced half-equation for the process $SO_3^{2-} \rightarrow SO_4^{2-}$
The oxidation number of S is increasing from +4 to +6, so S is being oxidised
There is one S on each side, so the S is already balanced
The S is losing two electrons, so two electrons are required on the right:

$$SO_3^{2-} \rightarrow SO_4^{2-} + 2e^-$$

There are three O atoms on the left and four on the right, so one water is needed on the left:

$$SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2e^-$$

There are two H atoms on the left and none on the right, so two H ions are needed on the right:

$$\mathbf{SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2e^-}$$

Method 2: (this does not use oxidation numbers and is easier in more complex processes)

- Identify the atom being oxidised or reduced, and make sure there are the same number of that atom on both sides (by balancing)
- balance O atoms by adding water
- balance H atoms by adding H^+
- add the necessary number of electrons to ensure the charge on both sides is the same

Example: Write a balanced half-equation for the process $H_2SO_4 \rightarrow H_2S$
S is being reduced; there is one sulphur on each side, so the S is already balanced
There are four O atoms on the left and none on the right, so four waters are needed on the right:

$$H_2SO_4 \rightarrow H_2S + 4H_2O$$

There are two H atoms on the left and ten on the right, so eight H ions are needed on the left:

$$H_2SO_4 + 8H^+ \rightarrow H_2S + 4H_2O$$

The total charge on the left is +8 and on the right is 0. So eight electrons must be added to the left to balance the charge:

$$H_2SO_4 + 8H^+ + 8e^- \rightarrow H_2S + 4H_2O$$

Test Your Progress: Topic 6 Exercise 4 Deduce the half-equations for the following processes, and state whether they are oxidation or reduction:	
(a) $PbO_2 \rightarrow Pb^{2+}$ (b) $Cl^- \rightarrow Cl_2$ (c) $S_2O_3^{2-} \rightarrow S_4O_6^{2-}$ (d) $IO_3^- \rightarrow I_2$ (e) $I^- \rightarrow I_2$ (f) $ClO^- \rightarrow ClO_3^-$	(g) $ClO^- \rightarrow Cl^-$ (h) $H_2SO_4 \rightarrow SO_2$ (i) $Br^- \rightarrow Br_2$ (j) $H_2SO_4 \rightarrow S$ (k) $H_2SO_4 \rightarrow H_2S$ (l) $SO_3^{2-} \rightarrow SO_4^{2-}$

(e) Redox reactions

Half-equations consider gain and loss of electrons, but in fact electrons cannot be created or destroyed; they can only be transferred from species to species. Gain of electrons by one species necessarily involves loss of electrons by another. Oxidation and reduction thus always occur simultaneously; an oxidation is always accompanied by a reduction and vice versa. Any reaction consisting of the oxidation of one species and the reduction of another is known as a **redox** reaction.

A redox reaction can be derived by combining an oxidation half-equation with a reduction half-equation in such a way that the total number of electrons gained is equal to the total number of electrons lost.

Eg $H_2SO_4 + 8H^+ + 8e^- \rightarrow H_2S + 4H_2O$ - reduction
$2I^- \rightarrow I_2 + 2e$ – oxidation
(the oxidation half-equation must be multiplied by 4 to equate the electrons)
$8I^- \rightarrow 4I_2 + 8e^-$
overall: $H_2SO_4 + 8H^+ + 8I^- \rightarrow H_2S + 4H_2O + 4I_2$ - redox

Eg $Al^{3+} + 3e^- \rightarrow Al$ - reduction
$2O^{2-} \rightarrow O_2 + 4e^-$ - oxidation
(the reduction half-equation must be multiplied by 4 and the oxidation half-equation by 3 to equate the electrons)
$4Al^{3+} + 12e^- \rightarrow 4Al$
$6O^{2-} \rightarrow 3O_2 + 12e^-$
overall: $4Al^{3+} + 6O^{2-} \rightarrow 4Al + 3O_2$ – redox

Test Your Progress: Topic 6 Exercise 5 Use the half-equations from Exercise 4 to write balanced equations for the following redox reactions:	
(a) PbO_2 with Cl^- to make Pb^{2+} and Cl_2 (b) $S_2O_3^{2-}$ with I_2 to make $S_4O_6^{2-}$ and I^- (c) IO_3^- with I^- to make I_2 (d) ClO^- to make ClO_3^- and Cl^- (e) H_2SO_4 with Br^- to make SO_2 and Br_2	(f) H_2SO_4 with I^- to make I_2 and H_2S (g) ClO^- with I^- to make Cl^- and I_2 (h) PbO_2 with SO_3^{2-} to make SO_4^{2-} and Pb^{2+} (i) H_2SO_4 with I^- to make S and I_2

(f) Oxidising agents and reducing agents

The species which is reduced is accepting electrons from the other species and thus causing it to be oxidised. It is thus an **oxidising agent**.
H_2SO_4, Al^{3+} and Cl_2 are all oxidising agents.

The species which is oxidised is donating electrons to another species and thus causing it to be reduced. It is thus a **reducing agent**.
Na, O^{2-}, I^- and $S_2O_3^{2-}$ are all reducing agents.

A redox reaction can thus be described as a transfer of electrons from a reducing agent to an oxidising agent.

Eg $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$
Half-equations: $I_2 + 2e \rightarrow 2I^-$ (reduction)
$2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e$ (oxidation)
I_2 is the oxidising agent; $S_2O_3^{2-}$ is the reducing agent.

Eg $H_2SO_4 + 8H^+ + 8I^- \rightarrow H_2S + 4H_2O + 4I_2$
Half-equations: $H_2SO_4 + 8H^+ + 8e \rightarrow H_2S + 4H_2O$ (reduction)
$2I^- \rightarrow I_2 + 2e$ (oxidation)
H_2SO_4 is the oxidising agent, I^- is the reducing agent

Test Your Progress: Topic 6 Exercise 6
Identify the reducing agents and the oxidising agents in Topic 6 Exercise 5

2) Common Redox Reactions

(a) Reaction of Metals with Acids

Most metals react with acids to make a salt and hydrogen.

Eg $Mg + 2HCl \rightarrow MgCl_2 + H_2$ (salt = magnesium chloride)
$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$ (salt = zinc sulphate)

This reaction is an example of a **redox** reaction.
The metal is oxidised from 0 to +2 (in the above reactions):
$Mg \rightarrow Mg^{2+} + 2e^-$ or $Zn \rightarrow Zn^{2+} + 2e^-$

The H^+ ion in the acid is reduced from +1 to 0 (in hydrogen)
$2H^+ + 2e^- \rightarrow H_2$

The overall equation can therefore be written:
$Mg + 2H^+ \rightarrow Mg^{2+} + H_2$ or $Zn + 2H^+ \rightarrow Zn^{2+} + H_2$
The chloride and sulphate ions are not part of the reaction and can be omitted from the equation.

This reaction can be a useful way of preparing salts:

Practical: prepare a sample of zinc sulphate from zinc and sulphuric acid

1) Measure out 20 cm^3 of 0.5 $moldm^{-3}$ sulphuric acid into a 100 cm^3 beaker.
2) Warm the beaker gently on a tripod until the temperature reaches 50 ^oC.
3) Add 1 g of zinc granules gradually to the beaker over a period of 2 minutes, stirring gently.
4) Heat gently for a few minutes.
5) Allow the mixture to cool.
6) Place a folded piece of filter paper inside a filter funnel, and then place the funnel into the neck of a 250 cm^3 conical flask.
7) Pour the warm mixture into the filter funnel and allow the solution to filter through. A colourless solution should collect in the conical flask.
8) Rinse the beaker and then pour the colourless solution back into it.
9) Heat the mixture gently until 80% of the water has evaporated.
10) Label the beaker with your name and leave it for a week.
11) Use a spatula to remove any crystals in the beaker and dry them with a piece of filter paper.

How many moles of sulphuric acid (H_2SO_4) are used in the reaction?
How many moles of zinc (Zn) are used in the reaction?
Why is it important that the Zn is in excess?

https://www.youtube.com/watch?v=uwbJcDO3vCM

(b) Reaction of Metals with Bases

Most metals do not react with bases as the H^+ concentration is too low, although the most reactive metals react with water at any pH.
A few metals which form amphoteric oxides (such as Al) can react with bases to form an oxoanion:

$Al + 3NaOH \rightarrow Na_3AlO_3 + 1.5H_2$

Just as with acids, the metal is oxidised from 0 to (in this case) +3. The H in NaOH is reduced from +1 to 0. This is therefore another example of a redox reaction.

(c) Metal Displacement Reactions

Metals will displace any less reactive metal from a solution of its salt.

Eg $Zn + CuSO_4 \rightarrow ZnSO + CuSO_4$

The more reactive metal (eg Zn) is oxidised from 0 to (in this case) +2 ($Zn \rightarrow Zn^{2+} + 2e^-$)
The less reactive metal (eg Cu) is reduced from (in this case) +2 to 0 ($Cu^{2+} + 2e^- \rightarrow Cu$)
The sulphate ion is not directly involved in the reaction.

The simplified chemical equation can therefore be written: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$.
This is therefore another example of a redox reaction.

(d) Formation and Reduction of Metallic Oxides

(i) Formation of Metallic Oxides

Most metals react with directly with oxygen to produce oxides:

Eg $2Mg + O_2 \rightarrow 2MgO$
$4Al + 3O_2 \rightarrow 2Al_2O_3$

During these reactions, the metal is oxidised from 0 to its typical oxidation state in oxides.
(eg $Mg \rightarrow Mg^{2+} + 2e^-$, or $Al \rightarrow Al^{3+} + 3e^-$)
The oxygen is reduced from 0 to -2.

The reaction of metals with oxygen to form oxides is an example of a redox reaction.

(ii) Reduction of Metallic Oxides

Many metals are found in nature (ie in their ores) as metallic oxides. They have to be converted from the metal oxide to the metal, usually be reacting them with a more reactive element.

Eg $Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$
$ZnO + C \rightarrow Zn + 2CO$
$CuO + H_2 \rightarrow Cu + H_2O$

In all of these reactions, the metal in the oxide is reduced from +2 or +3 to 0.
The more reactive element is oxidised. (H from 0 to +1, C from 0 to +2 or +4, Al from 0 to +3).

The reaction of metal oxides is therefore an example of a redox reaction.

Test Your Progress: Topic 6 Exercise 6
Write balanced symbol equations (with or without spectator ions) for the following redox reactions. In each case, state which atom is oxidised and which atom is reduced by writing r (for reduced) or o (for oxidised) under the atom in the equation.

(a) Zinc with hydrochloric acid to form zinc (II) chloride
(b) Aluminium with nitric acid
(c) Zinc (II) sulphate with magnesium
(d) Copper with silver (I) nitrate to form copper (II) nitrate
(e) Silver with oxygen to form silver (I) oxide
(f) Iron (III) oxide with carbon monoxide to form iron and carbon dioxide
(g) Titanium (IV) chloride with magnesium to form titanium and magnesium chloride

(e) Tests for oxidising and reducing agents

(i) Test for oxidising agents

To determine whether something is an oxidising agent, you need to add a reducing agent and observe whether a reaction takes place.

The easiest reducing agent to use is **potassium iodide (KI).** The iodide ions are readily oxidised to iodine. The oxidation number of the I increases from -1 to 0 ($2I^- \rightarrow I_2 + 2e^-$). Iodine ($I_2$) is a brown colour in solution. If a brown colour forms when KI is added to a solution, then that solution must contain a reducing agent (something in that solution is oxidised). The presence of iodine can be confirmed by adding a small quantity of starch and observing an intense blue/black solution.

Eg Fe^{3+} is an oxidising agent. It is readily reduced to Fe^{2+} by iodide ions:

$$2Fe^{2+} + 2I^- \rightarrow 2Fe^{2+} + I_2$$

On addition of KI to a solution containing Fe^{2+} ions, a brown colour of iodine will appear. This brown colour will turn blue/black if starch is added.

(ii) Test for reducing agents

To determine whether something is a reducing agent, you need to add an oxidising agent and observe whether a reaction takes place.

The easiest oxidising agents to use are **acidified potassium dichromate (VI) ($K_2Cr_2O_7$) or acidified potassium manganate (VII)**.

$K_2Cr_2O_7$ is an orange colour, but when it reacts with reducing agents in acidic conditions the Cr is reduced from +6 to +3 and the resulting solution is green. Reducing agents will therefore turn an acidified solution of potassium dichromate from orange to green.

$KMnO_4$ is a dark purple colour, but when it reacts with reducing agents in acidic conditions the Mn is reduced from +7 to +2 and the resulting solution is colourless. Reducing agents will therefore decolorise an acidified solution of potassium dichromate.

Practical: testing for oxidising and reducing agents

1) Measure out approximately 1 cm^3 of iron (II) sulphate into three different test tubes.
2) Measure out approximately 1 cm^3 of iron (III) sulphate into three different test tubes.
3) Add approximately 1 cm^3 of potassium iodide to the first test tube containing Fe^{2+} and the first test tube containing Fe^{3+}. Which solution turns brown?
4) Then add a few drops of starch indicator to both mixtures. Which solution turns blue/black?
5) Add approximately 1 cm^3 of acidified potassium dichromate (VI) to the second test tube containing Fe^{2+} and the second test tube containing Fe^{3+}. Which solution turns green?
6) Add approximately 1 cm^3 of acidified potassium manganate (VII) to the second test tube containing Fe^{2+} and the second test tube containing Fe^{3+}. In which solution does the purple colour disappear?

www.youtube.com/watch?v=fpG6XRg2gSU
www.youtube.com/watch?v=nRlabYLKvsQ
www.youtube.com/watch?v=r9iexLFfedY

(f) The Rusting of Iron

Rusting is a redox reaction between iron, oxygen and water. It involves a number of different chemical reactions but can be summarised as:

$$4Fe + 6H_2O + 3O_2 \rightarrow 4Fe(OH)_3$$

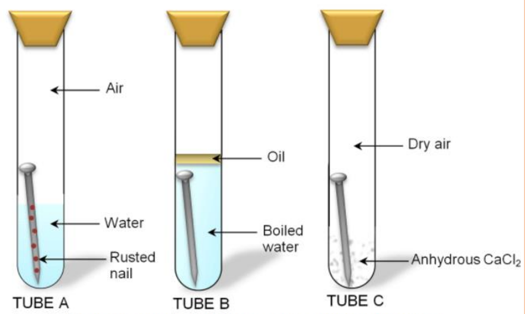
The iron is oxidised from 0 to +3. The oxygen is reduced from 0 to -2.

Most metals react with oxygen and water in the environment, but usually the oxides or hydroxides formed stick to the surface of the metal and protect it from further corrosion. Most metals are therefore protected naturally by their oxide layer. The oxides and hydroxides of iron, however, flake off the surface and so provide the iron metal with no protection. Hence the oxidation process continues indefinitely.

It can be shown experimentally that both air and water are necessary for rusting to happen:

Practical: demonstrating what oxygen and water are both required for rusting

1) Place three iron nails in three different test tubes (see below)
2) Add water to the first test tube until it is covering half of the nail
3) Add recently boiled water (no oxygen) to the second test tube until it completely covers the nail, and then add a thin layer of vegetable oil
4) Add no water to the third test tube. Instead add a drying agent such as $CaCl_2$
5) Leave for a week and then observe the condition of all three nails



Investigation the conditions under which iron nail rusts

Practical: demonstrating that acids, bases, salts and heating increase the rate of rusting

1) Place five iron nails in five different test tubes
2) Add water to the first test tube until it is covering half of the nail
3) Add 1 $moldm^{-3}$ HCl to the second test tube until it is covering half of the nail
4) Add 1 $moldm^{-3}$ NaOH to the third test tube until it is covering half of the nail
5) Add 1 $moldm^{-3}$ HCl to the fourth test tube until it is covering half of the nail
6) Add 1 $moldm^{-3}$ NaCl to the fifth test tube until it is covering half of the nail
7) Leave the first four test tubes in a warm place for a week; leave the fifth test tube in a fridge for a week
8) Observe the degree of rusting on each nail.

Any solutions containing electrolytes such as acids, bases or salts will increase the rate of rusting.
Acids will actually also dissolve the rust (by neutralising it) and it may appear not to be causing rusting, but the metal is still corroding.

Iron is a very widely used metal and the rusting of iron is a major economic problem. There are many items in the home which are prone to rusting, including fridges, freezers, food containers - any metallic object made of iron. It is important to prevent iron from rusting as much as possible and there are a number of methods available for doing this.

The simplest methods involve oiling, greasing and painting. This provides limited protection to the iron by preventing water and oxygen from coming into contact with the surface of the iron. Titanium nitride (TiN) is another inert protective layer which can be applied to an iron surface to prevent rusting. The protection will only last until it is scratched or damaged.

If the iron is not under water, maintaining a dry environment through the presence of drying agents in the room will also prevent rusting.

Iron can also be protecting by adding a thin layer of another metal to the surface of the iron. Tin and zinc are the metals most commonly used for this purpose. The coating of iron with a thin layer of zinc is known as **galvanisation**. Zinc and tin both get oxidised by air, but the oxide formed sticks to the surface of the metal, providing protection.

The thin layer of tin or zinc can be applied either by dipping the iron into the molten metal (known as hot-dipping) or by electroplating. The principles of electroplating will be discussed in the next section.

There are some other methods for preventing rusting which involve electrochemistry. These will also be discussed in the next section.