Mastering Chemistry

- Book 2B
- Topic 5 Redox Reactions, Chemical Cells and Electrolysis

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20.1 Oxidation and reduction in terms of gain and loss of oxygen (p.45)

- Magnesium burns in air with a white flame. Magnesium oxide is formed.

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

The magnesium gains oxygen. It is oxidised.

Magnesium burns in air to give magnesium oxide

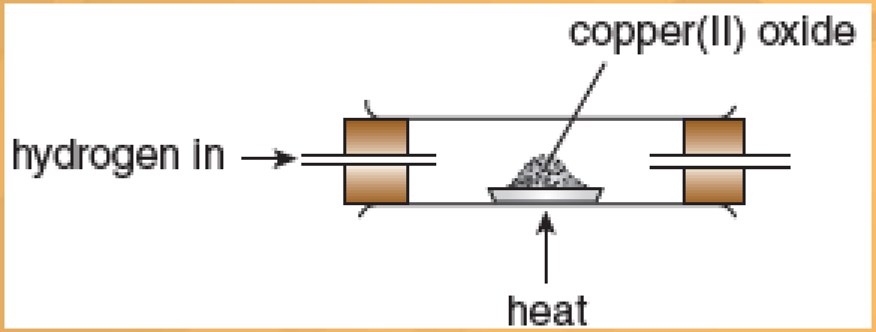
Oxidation (氧化作用) is the gain of oxygen by a chemical species. The chemical species is oxidised.

20.1 Oxidation and reduction in terms of gain and loss of oxygen (p.45)

- Now look at what happens when hydrogen is passed over heated copper(II) oxide. Copper is formed in the process.

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$$

This time copper(II) oxide loses oxygen. It is reduced.



When hydrogen is passed over heated copper(II) oxide, the oxide is reduced to copper

Reduction (還原作用) is the loss of oxygen from a chemical species. The chemical species is reduced.

20.1 Oxidation and reduction in terms of gain and loss of oxygen (p.45)

- The reaction between copper(II) oxide and hydrogen is a redox reaction.

loss of oxygen: reduction

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$$

gain of oxygen: oxidation

Reduction and oxidation always take place together. So the reaction is called a redox reaction (氧化還原反應).

20.1 Oxidation and reduction in terms of gain and loss of oxygen (p.45)

Oxidising agent and reducing agent

- When magnesium and copper(II) oxide are heated together, the following reaction occurs:

$$Mg(s) + CuO(s) \longrightarrow MgO(s) + Cu(s)$$

- Magnesium is oxidised because it gains oxygen — copper(II) oxide causes this to happen and so it is the **oxidising agent (氧化劑)**.
- Copper(II) oxide is reduced because it loses oxygen — magnesium causes this to happen and so it is the **reducing agent (還原劑)**.

20.1 Oxidation and reduction in terms of gain and loss of oxygen (p.45)

An oxidising agent is a chemical species that causes oxidation of another chemical species. The oxidising agent itself becomes reduced in the process.

A reducing agent is a chemical species that causes reduction of another chemical species. The reducing agent itself becomes oxidised in the process.

20.1 Oxidation and reduction in terms of gain and loss of oxygen (p.45)

oxidising agent;
becomes reduced itself; loses oxygen

$$Mg(s) + CuO(s) \longrightarrow MgO(s) + Cu(s)$$

reducing agent;
becomes oxidised itself; gains oxygen

Characteristics of reducing and oxidising agents

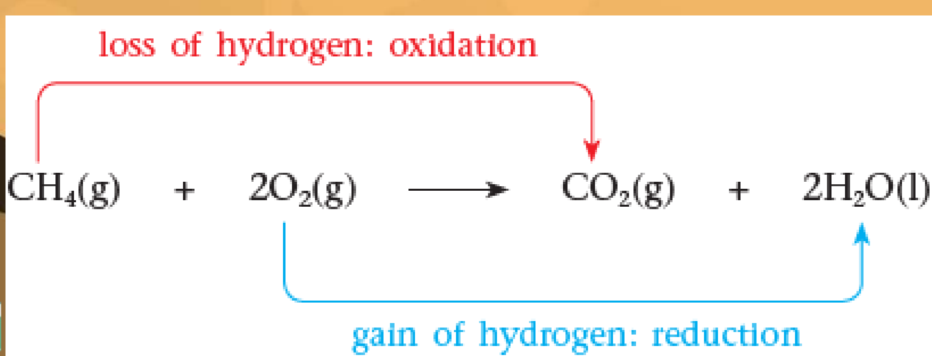
20.2 Oxidation and reduction in terms of loss and gain of hydrogen (p.47)

- Oxidation and reduction can also be defined in terms of loss and gain of hydrogen.

Oxidation is the loss of hydrogen from a chemical species.
Reduction is the gain of hydrogen by a chemical species.

- Natural gas consists mainly of methane, $CH_4(g)$. It can be used for cooking. The chemical equation for the complete combustion of methane is shown below.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$



The redox reaction that occurs when methane burns completely in oxygen

20.2 Oxidation and reduction in terms of loss and gain of hydrogen (p.47)

- Oxygen causes the oxidation of methane and so it is the oxidising agent. Methane causes the reduction of oxygen and so it is the reducing agent.

oxidising agent;
gains hydrogen

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

reducing agent;
loses hydrogen

Methane is the reducing agent and oxygen is the oxidising agent in the redox reaction that occurs when methane burns in oxygen

20.2 Oxidation and reduction in terms of loss and gain of hydrogen (p.47)

Practice 20.1

1 Zinc reacts with copper(II) oxide to form zinc oxide and copper when heated.

$$Zn(s) + CuO(s) \longrightarrow ZnO(s) + Cu(s)$$

a) Explain, in terms of the gain and loss of oxygen, why this is a redox reaction.

b) Identify the reducing agent. Explain your answer.

2 Hydrogen sulphide reacts with chlorine to form sulphur and hydrogen chloride.

$$H_2S(g) + Cl_2(g) \longrightarrow S(s) + 2HCl(g)$$

a) Explain, in terms of the gain and loss of hydrogen, why this is a redox reaction.

b) Identify the oxidising agent. Explain your answer.

3 In each of the following reactions, identify which chemical species is oxidised and which is reduced.

a) $CO(g) + Ag_2O(s) \longrightarrow CO_2(g) + 2Ag(s)$

b) $MnO_2(s) + 4HCl(aq) \longrightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$

20.2 Oxidation and reduction in terms of loss and gain of hydrogen (p.47)

Practice 20.1 (continued)

1 a) Zinc gains oxygen. It undergoes oxidation.
Copper(II) oxide loses oxygen. It undergoes reduction.
Thus, this is a redox reaction.
b) Zinc is the reducing agent.
It causes the reduction of copper(II) oxide.

2 a) Hydrogen sulphide loses hydrogen.
It undergoes oxidation.
Chlorine gains hydrogen. It undergoes reduction.
Thus, this is a redox reaction.
b) Chlorine is the oxidising agent.
It causes the oxidation of hydrogen sulphide.

3 a) CO(g) is oxidised.
Ag_2O(s) is reduced.
b) MnO_2(s) is reduced.
HCl(aq) is oxidised.

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

- When magnesium burns in air, magnesium oxide is formed.

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

- Magnesium oxide is an ionic compound. It is made up of magnesium ions (Mg^{2+}) and oxide ions (O^{2-}).
- During the reaction, each magnesium atom loses two electrons to form a magnesium ion (Mg^{2+}).
- Each oxygen atom gains two electrons to form an oxide ion (O^{2-}). There are two oxygen atoms in an oxygen molecule, so four electrons are gained.

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

- These two processes can be described by the following half equations:

Oxidation	$Mg \longrightarrow Mg^{2+} + 2e^-$
Reduction	$O_2 + 4e^- \longrightarrow 2O^{2-}$

Oxidation is the loss of electron(s) from a chemical species.
Reduction is the gain of electron(s) by a chemical species.

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

- Mnemonic for redox reactions in terms of loss and gain of electrons.

Oxidation
Is
Loss of electron(s)

Reduction
Is
Gain of electron(s)

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

- Consider the burning of sodium in chlorine gas to form sodium chloride.

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

Burning of sodium in chlorine gas

- By definition, sodium undergoes oxidation while chlorine undergoes reduction.

Oxidation	$Na \longrightarrow Na^+ + e^-$
Reduction	$Cl_2 + 2e^- \longrightarrow 2Cl^-$

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

Oxidising agent and reducing agent

- Consider the displacement reaction between zinc and copper(II) sulphate solution.

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

- This reaction can be shown as two half equations:

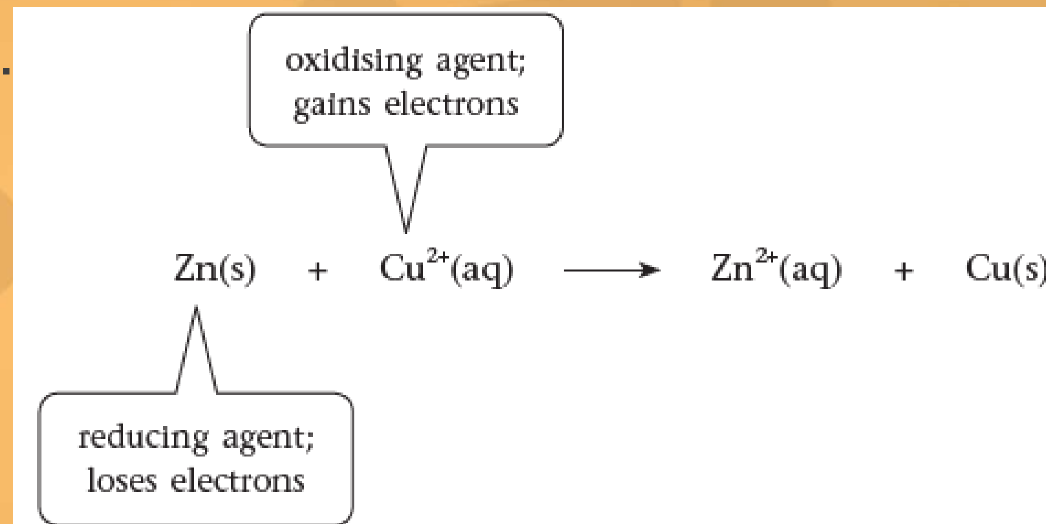
electron loss (oxidation) $Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

electron gain (reduction) $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

The displacement reaction between zinc and copper(II) sulphate solution

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

- Zinc is the reducing agent and copper(II) ion is the oxidising agent in the displacement reaction between zinc and copper(II) sulphate solution.



When an oxidising agent reacts, it gains electrons and is therefore, reduced.
When a reducing agent reacts, it loses electrons and is therefore, oxidised.

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

General trend of the reducing power of metals and the oxidising power of metal ions

- Metals towards the top of the electrochemical series lose electrons more readily to form ions. In other words, they are strong reducing agents.
- Ions of metals towards the top of the series are less likely to gain electrons to form metals. Hence ions of these metals are weak oxidising agents.

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

- The electrochemical series of metals is shown below:

Metal ion				Metal
$K^{+}(aq)$	+	e^{-}	$\rightleftharpoons$	$K(s)$
$Ca^{2+}(aq)$	+	$2e^{-}$	$\rightleftharpoons$	$Ca(s)$
$Na^{+}(aq)$	+	e^{-}	$\rightleftharpoons$	$Na(s)$
$Mg^{2+}(aq)$	+	$2e^{-}$	$\rightleftharpoons$	$Mg(s)$
$Al^{3+}(aq)$	+	$3e^{-}$	$\rightleftharpoons$	$Al(s)$
$Zn^{2+}(aq)$	+	$2e^{-}$	$\rightleftharpoons$	$Zn(s)$
$Fe^{2+}(aq)$	+	$2e^{-}$	$\rightleftharpoons$	$Fe(s)$
$Pb^{2+}(aq)$	+	$2e^{-}$	$\rightleftharpoons$	$Pb(s)$
$2H^{+}(aq)$	+	$2e^{-}$	$\rightleftharpoons$	$H_2(g)$
$Cu^{2+}(aq)$	+	$2e^{-}$	$\rightleftharpoons$	$Cu(s)$
$Ag^{+}(aq)$	+	e^{-}	$\rightleftharpoons$	$Ag(s)$
$Au^{+}(aq)$	+	e^{-}	$\rightleftharpoons$	$Au(s)$

oxidising power of metal ions increasing

reducing power of metals increasing

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

A metal high in the electrochemical series is a strong reducing agent while its ion is a weak oxidising agent.

- Towards the bottom of the electrochemical series, the metals become progressively weaker reducing agents, but the oxidising power of their ions increases.

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

- Metals above hydrogen in the series are stronger reducing agents than hydrogen and should reduce hydrogen ions to hydrogen gas. For example, magnesium reduces the hydrogen ions in dilute hydrochloric acid to hydrogen gas.

$$Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_2(g)$$

- Metals below hydrogen in the series will not react with acids. For example, copper and silver do not react with hydrochloric acid.

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

Practice 20.2

1 Iron reacts with chlorine according to the equation below.

$$2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$$

Explain, in terms of gain and loss of electrons, why this is a redox reaction.

Iron loses electrons. It undergoes oxidation.
Chlorine gains electrons. It undergoes reduction.
Thus, this is a redox reaction.

2 In each of the following reactions, identify whether the underlined chemical species has been oxidised, reduced or neither.

a) $\underline{Mg(s)} + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$

Mg(s) loses electrons. It has been oxidised.

b) $\underline{Ag^+(aq)} + Cl^-(aq) \longrightarrow AgCl(s)$

$Ag^+(aq)$ does not lose / gain electrons. It is neither oxidised nor reduced.

20.3 Oxidation and reduction in terms of loss and gain of electrons (p.49)

Practice 20.2 (continued)

3 The oxidising power of three ions increases in the order as shown below:

$$Rh^{2+}(aq) < Pd^{2+}(aq) < Pt^{2+}(aq)$$

Complete the table below to show whether displacement reactions occur between metals Rh, Pd and Pt with their ions.

Ion \ Metal	Rh	Pd	Pt
$Rh^{2+}(aq)$	—	✘	✘
$Pd^{2+}(aq)$	✔	—	✘
$Pt^{2+}(aq)$	✔	✔	—

20.4 Oxidation numbers (p.54)

- The oxidation number of an element is an imaginary charge assigned to it according to a set of rules. In assigning an oxidation number to each element in a compound, it is assumed that the compound is ionic.

Rules for working out oxidation numbers

There are some general rules for working out oxidation numbers.

1 The oxidation number of an element is zero.
For example, the oxidation number of oxygen in O_2 is zero.

2 The oxidation number of the element in a monoatomic ion is the charge on that ion.
For example, the oxidation number of iron in Fe^{2+} ion is +2.

20.4 Oxidation numbers (p.54)

Rules for working out oxidation numbers

3 The sum of the oxidation numbers of the elements in a neutral compound is zero.
For example, in sodium chloride (NaCl), the oxidation number of sodium (+1) plus the oxidation number of chlorine (–1) equals zero.

4 The sum of the oxidation numbers of the elements in a polyatomic ion is the charge on that ion.
For example, in sulphate ion (SO_4^{2-}), the oxidation number of sulphur (+6) plus four times the oxidation number of oxygen (–2) equals –2.

20.4 Oxidation numbers (p.54)

Rules for working out oxidation numbers

5 Some elements have fixed oxidation numbers in all of their compounds.

Element	Oxidation number
Group I metals	+1
Group II metals	+2
Fluorine	–1
Hydrogen	+1 (except when it is combined with a reactive metal, for example in sodium hydride, NaH, where it is –1)
Oxygen	–2 (except in a few compounds such as hydrogen peroxide, H_2O_2, where it is –1 and oxygen difluoride, OF_2, where it is +2)

20.4 Oxidation numbers (p.54)

Q (Example 20.1)

Work out the oxidation number of nitrogen in

a) NH_3.

b) NO_2^-.

c) NO_3^-.

A

a) The oxidation number of H in NH_3 is +1.
Suppose the oxidation number of N in NH_3 is x.
$x + 3 \times (+1) = 0$
$\therefore x = +3$

b) The oxidation number of O in NO_2^- is –2.
Suppose the oxidation number of N in NO_2^- is x.
$x + 2 \times (-2) = -1$
$\therefore x = +3$

20.4 Oxidation numbers (p.54)

Q (Example 20.1) (continued)

A

c) The oxidation number of O in NO_3^- is –2.
Suppose the oxidation number of N in NO_3^- is x.
$x + 3 \times (-2) = -1$
$\therefore x = +5$

20.4 Oxidation numbers (p.54)

Q (Example 20.2)

Work out the oxidation number of chlorine in
a) NaOCl.
b) ClO_2^-.
c) $HClO_3$.

A

a) The oxidation numbers of Na and O in NaOCl are +1 and –2 respectively.
Suppose the oxidation number of Cl in NaOCl is *x*.
$+1 + (-2) + x = 0$
$\therefore x = +1$

b) The oxidation number of O in ClO_2^- is –2.
Suppose the oxidation number of Cl in ClO_2^- is *x*.
$x + 2 \times (-2) = -1$
$\therefore x = +3$

20.4 Oxidation numbers (p.54)

Q (Example 20.2) (continued)

A

c) The oxidation numbers of H and O in $HClO_3$ are +1 and –2 respectively.
Suppose the oxidation number of Cl in $HClO_3$ is *x*.
$(+1) + x + 3 \times (-2) = 0$
$\therefore x = +5$

20.4 Oxidation numbers (p.54)

Practice 20.3

Work out the oxidation number of the underlined element in each of the following chemical species.

a) $\underline{S}O_2$
b) $KH\underline{S}O_4$
c) $\underline{O}_2F_2$
d) $Na_3\underline{P}O_4$
e) $\underline{Mn}O_4^-$
f) $[\underline{Cu}Cl_4]^{2-}$

20.4 Oxidation numbers (p.54)

Practice 20.3 (continued)

a) The oxidation number of O in SO_2 is –2.
Suppose the oxidation number of S in SO_2 is x.
$x + 2 \times (-2) = 0$
$\therefore x = +4$

b) The oxidation numbers of K, H and O in $KHSO_4$ are +1, +1 and –2 respectively.
Suppose the oxidation number of S in $KHSO_4$ is x.
$(+1) + (+1) + x + 4 \times (-2) = 0$
$\therefore x = +6$

c) The oxidation number of F in O_2F_2 is –1.
Suppose the oxidation number of O in O_2F_2 is x.
$2x + 2 \times (-1) = 0$
$\therefore x = +1$

20.4 Oxidation numbers (p.54)

Practice 20.3 (continued)

d) Na_3PO_4 contains Na+ ions and PO_4^{3-} ions.
The oxidation number of O in PO_4^{3-} is –2.
Suppose the oxidation number of P in PO_4^{3-} is *x*.
$x + 4 \times (-2) = -3$
$\therefore x = +5$

e) The oxidation number of O in MnO_4^- is –2.
Suppose the oxidation number of Mn in MnO_4^- is *x*.
$x + 4 \times (-2) = -1$
$\therefore x = +7$

f) The oxidation number of Cl in $[CuCl_4]^{2-}$ is –1.
Suppose the oxidation number of Cu in $[CuCl_4]^{2-}$ is *x*.
$x + 4 \times (-1) = -2$
$\therefore x = +2$

20.5 Oxidation and reduction in terms of changes in oxidation numbers (p.57)

- Consider the following half equations:

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

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

$$2I^-(aq) \longrightarrow I_2(aq) + 2e^-$$

In all three cases, electrons are lost so the processes are oxidations.

- The oxidation number of the element increases in each case:
  - Zn increases from 0 to +2;
  - Fe increases from +2 to +3;
  - I increases from –1 to 0.

20.5 Oxidation and reduction in terms of changes in oxidation numbers (p.57)

- Consider the following half equations:

  $Br_2(aq) + 2e^- \longrightarrow 2Br^-(aq)$

  $Cr^{3+}(aq) + e^- \longrightarrow Cr^{2+}(aq)$

  $2H^+(aq) + 2e^- \longrightarrow H_2(g)$

  In all three cases, electrons are gained so the processes are reductions.

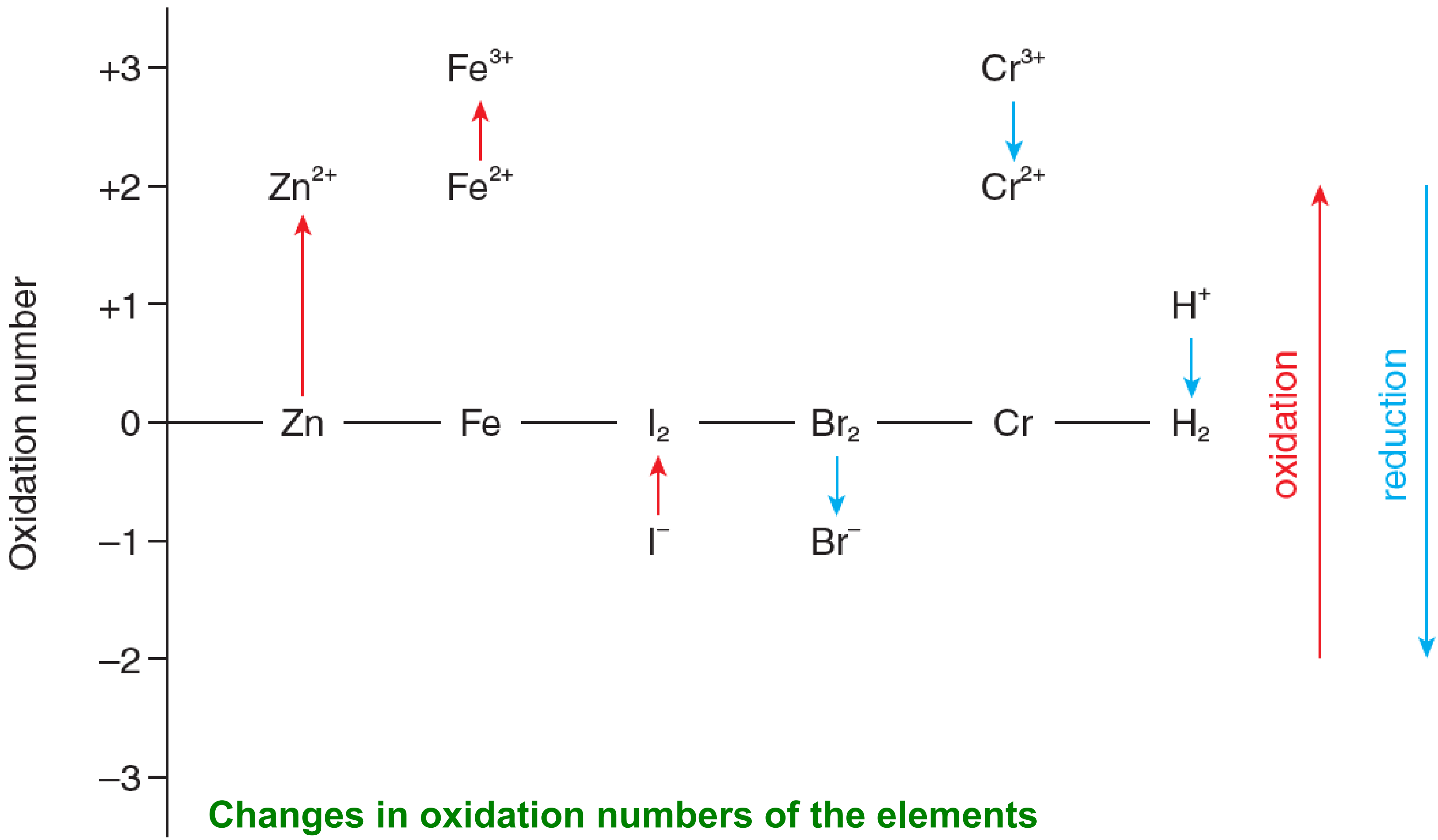
- The oxidation number of the element decreases in each case:
  - Br decreases from 0 to –1;
  - Cr decreases from +3 to +2;
  - H decreases from +1 to 0.

20.5 Oxidation and reduction in terms of changes in oxidation numbers (p.57)

Oxidation occurs when the oxidation number of an element in a chemical species increases.

Reduction occurs when the oxidation number of an element in a chemical species decreases.

20.5 Oxidation and reduction in terms of changes in oxidation numbers (p.57)



Changes in oxidation numbers of the elements

20.5 Oxidation and reduction in terms of changes in oxidation numbers (p.57)

- The definitions of oxidation and reduction are shown below:

Defined in terms of	Oxidation	Reduction
gain and loss of oxygen	gain of oxygen	loss of oxygen
loss and gain of hydrogen	loss of hydrogen	gain of hydrogen
loss and gain of electron(s)	loss of electron(s)	gain of electron(s)
changes in oxidation numbers	increase in oxidation number	decrease in oxidation number

20.5 Oxidation and reduction in terms of changes in oxidation numbers (p.57)

Oxidising agent and reducing agent

- When zinc reacts with concentrated nitric acid, the oxidation numbers of zinc and nitrogen change as shown below.

$$Zn(s) + 4HNO_3(aq) \longrightarrow Zn(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

oxidation number: 0 (Zn), +5 (N in HNO_3), +2 (Zn in $Zn(NO_3)_2$), +4 (N in NO_2)

- The oxidation number of zinc increases from 0 to +2. Hence zinc undergoes oxidation.
- The oxidation number of nitrogen decreases from +5 to +4. Hence concentrated nitric acid undergoes reduction.

20.5 Oxidation and reduction in terms of changes in oxidation numbers (p.57)

- Zinc is the reducing agent and concentrated nitric acid is the oxidising agent in the reaction between zinc and concentrated nitric acid.

oxidising agent;
the oxidation number
of N decreases

$$Zn(s) + 4HNO_3(aq) \longrightarrow Zn(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

reducing agent;
the oxidation number
of Zn increases

20.5 Oxidation and reduction in terms of changes in oxidation numbers (p.57)

An oxidising agent increases the oxidation number of an element in another chemical species.

The oxidation number of an element in the oxidising agent decreases.

A reducing agent decreases the oxidation number of an element in another chemical species.

The oxidation number of an element in the reducing agent increases.

20.5 Oxidation and reduction in terms of changes in oxidation numbers (p.57)

- The characteristics of oxidising agents and reducing agents are shown below:

Oxidising agent	Reducing agent
causes oxidation	causes reduction
becomes reduced itself	becomes oxidised itself
loses oxygen	gains oxygen
gains hydrogen	loses hydrogen
gains electron(s)	loses electron(s)
oxidation number of an element in it decreases	oxidation number of an element in it increases

20.6 Using oxidation numbers to identify redox reactions (p.60)

- Redox reactions can be easily identified by
  - deducing the oxidation numbers of elements in the chemical species present in the equation;
  - examining the numbers to see if the oxidation number of any element has changed. If it has, the reaction is a redox reaction. If there is no change in oxidation numbers during the chemical reaction, then the reaction is not a redox reaction.

20.6 Using oxidation numbers to identify redox reactions (p.60)

Q (Example 20.3)

For each of the following reactions, identify whether it is a redox reaction by using changes in oxidation numbers.

a) $Cu(s) + 2AgNO_3(aq) \longrightarrow 2Ag(s) + Cu(NO_3)_2(aq)$

b) $2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$

c) $Ca(HCO_3)_2(aq) \longrightarrow CaCO_3(s) + H_2O(l) + CO_2(g)$

A

Consider each equation and deduce the oxidation numbers of elements in the chemical species.

a) $Cu(s) + 2AgNO_3(aq) \longrightarrow 2Ag(s) + Cu(NO_3)_2(aq)$

Cu	0			+2	increase in oxidation number
Ag		+1	0		decrease in oxidation number

Hence this is a redox reaction.

20.6 Using oxidation numbers to identify redox reactions (p.60)

Q (Example 20.3) (continued)

A

b) $2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$

Cl	+5	–1	decrease in oxidation number
O	–2	0	increase in oxidation number

Hence this is a redox reaction.

20.6 Using oxidation numbers to identify redox reactions (p.60)

Q (Example 20.3) (continued)

A

c) $Ca(HCO_3)_2(aq) \longrightarrow CaCO_3(s) + H_2O(l) + CO_2(g)$

	$Ca(HCO_3)_2(aq)$	$CaCO_3(s)$	$H_2O(l)$	$CO_2(g)$
Ca	+2	+2		
H	+1		+1	
C	+4	+4		+4
O	–2	–2	–2	–2

The oxidation numbers of all the elements remain unchanged. Hence this is not a redox reaction.

20.6 Using oxidation numbers to identify redox reactions (p.60)

Practice 20.4

1 For each of the following reactions, identify whether it is a redox reaction by using changes in oxidation numbers.

a) $SO_2(g) + 2H_2S(g) \longrightarrow 3S(s) + 2H_2O(l)$

b) $MgO(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2O(l)$

c) $2Al(s) + 3H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

2 Potassium iodide solution turns yellow with time due to the following redox reaction:

$$4KI(aq) + 2CO_2(g) + O_2(g) \longrightarrow 2K_2CO_3(aq) + 2I_2(aq)$$

a) Identify the oxidising agent and the reducing agent.

b) Explain why the potassium iodide solution turns yellow.

20.6 Using oxidation numbers to identify redox reactions (p.60)

Practice 20.4 (continued)

1 a) $SO_2(g) + 2H_2S(g) \longrightarrow 3S(s) + 2H_2O(l)$

S in SO_2	+4		0	decrease in oxidation number
S in H_2S		−2	0	increase in oxidation number

Hence this is a redox reaction.

b) $MgO(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2O(l)$

Mg	+2			+2	
O	−2				−2
H		+1			+1
Cl			−1	−1	

The oxidation numbers of all the elements remain unchanged.

Hence this is NOT a redox reaction.

c) $2Al(s) + 3H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

Al	0		+3		increase in oxidation number
H		+1		0	decrease in oxidation number

Hence this is a redox reaction.

20.6 Using oxidation numbers to identify redox reactions (p.60)

Practice 20.4 (continued)

2 a) Oxidising agent: O_2(g)
Reducing agent: KI(aq)
b) Due to the formation of I_2(aq).

20.7 Naming inorganic compounds (p.62)

- The concept of oxidation number is used in the modern chemical naming of inorganic compounds. This is called the **Stock system (司托克系統)**.
- Some compounds and ions contain an element that can have more than one oxidation number. Their names should include the oxidation number written as a Roman numeral in brackets.

20.7 Naming inorganic compounds (p.62)

- For example, there are two types of chloride of iron. They are named iron(II) chloride and iron(III) chloride.

Iron(II) chloride and iron(III) chloride

20.7 Naming inorganic compounds (p.62)

- In iron(II) chloride, the oxidation number of iron is +2. The compound contains Fe^{2+} ions and its chemical formula is $FeCl_2$.
- In iron(III) chloride, the oxidation number of iron is +3. The compound contains Fe^{3+} ions and its chemical formula is $FeCl_3$.
- The system is used to name chemical species containing transition metals, as well as tin and lead from Group IV of the Periodic Table where variable oxidation numbers are shown.
- For metals from Groups I, II and III, it is not necessary to indicate the oxidation number of the metal. For example, the chloride of calcium is named calcium chloride rather than calcium(II) chloride.

20.7 Naming inorganic compounds (p.62)

Naming oxoanions

- An oxoanion is a negative ion containing oxygen. Under the Stock system, the name always ends in –ate to show that oxygen is present.
- Names of oxoanions need to have some indication of the oxidation number of the element other than oxygen.
- For example, the oxidation number of nitrogen in the NO_2^- ion is +3 and hence the ion is named nitrate(III) ion. The oxidation number of nitrogen in the NO_3^- ion is +5 and hence the ion is named nitrate(V) ion.

20.7 Naming inorganic compounds (p.62)

- Some examples of naming oxoanions.

Chemical formula of oxoanion	Oxidation number of central element	Name by Stock system	Common name
CrO_4^{2-}	+6	chromate(VI) ion	chromate ion
$Cr_2O_7^{2-}$	+6	dichromate(VI) ion	dichromate ion
MnO_4^{2-}	+6	manganate(VI) ion	manganate ion
MnO_4^{-}	+7	manganate(VII) ion	permanganate ion
OCl^{-}	+1	chlorate(I) ion	hypochlorite ion
ClO_3^{-}	+5	chlorate(V) ion	chlorate ion
SO_3^{2-}	+4	sulphate(IV) ion	sulphite ion
SO_4^{2-}	+6	sulphate(VI) ion	sulphate ion

20.7 Naming inorganic compounds (p.62)

Practice 20.5

Write the chemical formula for each of the ions below.

Ion	Ionic charge	Chemical formula
Nitrate(V)	–1	NO_3^-
Sulphate(IV)	–2	SO_3^{2-}
Chlorate(III)	–1	ClO_2^-

20.8 Half equations of oxidation and reduction (p.64)

- The processes of oxidation and reduction happen at the same time in a redox reaction — one chemical species is oxidised while the other is reduced.
- A redox reaction can, therefore, be broken down into two half reactions, which can be represented by half equations to show the electron loss or gain.
- For example, when zinc is added to dilute hydrochloric acid, zinc is oxidised to zinc ions. The half equation is:

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

20.8 Half equations of oxidation and reduction (p.64)

- There are two points to notice when writing a half equation such as this:
  - This is an oxidation, so the electrons are on the right-hand side of the equation.
  - The equation must be balanced for charge as well as the numbers of atoms. In this example, both sides add up to a charge of zero.

- In the reaction between zinc and dilute hydrochloric acid, hydrogen ions are reduced to hydrogen. The half equation is:

  $$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

  This is a reduction, so the electrons are on the left-hand side of the equation.

20.8 Half equations of oxidation and reduction (p.64)

More complex half equations

- Some oxidising agents work in the presence of acids.
- For example, in the presence of dilute sulphuric acid, potassium permanganate solution is a strong oxidising agent. The permanganate ion (MnO_4^-) is reduced to manganese(II) ion (Mn^{2+}) in a redox reaction.

$$MnO_4^-(aq) + 8H^+(aq) + xe^- \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$$

where x is the number of electrons

- Hydrogen ions have to be added on the left-hand side. They pick up the oxygen atoms from the MnO_4^- ion and form water. There are four oxygen atoms in the MnO_4^- ion, so eight hydrogen ions are needed to form four water molecules.

20.8 Half equations of oxidation and reduction (p.64)

- The value of x can be worked out in either of two ways:
  - The equation must be balanced for charge. The charge on the right-hand side is +2. The charge on the left-hand side is:
    $-1 + 8 - x = 7 - x$
    The charge on the left-hand side must be equal to the charge on the right-hand side, i.e. +2. Therefore,
    $7 - x = +2$
    $x = 7 - 2$
    $= 5$
    There are five electrons on the left-hand side, so the correct half equation is:

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

20.8 Half equations of oxidation and reduction (p.64)

- The value of *x* can be worked out in either of two ways:
  - The oxidation number of manganese changes from +7 to +2, which is a change of 5. Therefore, there must be five electrons on the left. The correct half equation is:

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

20.8 Half equations of oxidation and reduction (p.64)

- Follow the procedure below when writing a half equation in acid solution:

  1 Balance the numbers of all atoms except hydrogen and oxygen.

  2 Add H_2O to the side deficient in oxygen to balance the number of oxygen atoms.

  3 Add H^+ to the side deficient in hydrogen to balance the number of hydrogen atoms.

  4 Add e^- to the side deficient in negative charge to balance the charge.

20.8 Half equations of oxidation and reduction (p.64)

Q (Example 20.4)

Write the half equation for the reduction of dichromate ion to chromium(III) ion in acid solution.

A

1	Balance the numbers of all atoms except hydrogen and oxygen.	$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+}$
2	Add H_2O to the side deficient in oxygen to balance the oxygen atoms.	There are 7 oxygen atoms on the left-hand side but none on the right-hand side, so add $7H_2O$ to the right-hand side: $Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$
3	Add H^+ to the side deficient in hydrogen to balance the hydrogen atoms.	There are 14 hydrogen atoms on the right-hand side but none on the left-hand side, so add $14H^+$ to the left-hand side: $Cr_2O_7^{2-} + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$
4	Add e^- to the side deficient in negative charge to balance the charge.	The half equation is: $Cr_2O_7^{2-}(aq) + 14H^+(aq) + xe^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$ The value of x can be worked out in either of two ways: • The charge on the right-hand side is +6, so the charge on the left-hand side is also +6. $(-2) + 14 \times (+1) - x = +6$ $x = 6$ • The oxidation number of chromium in $Cr_2O_7^{2-}$ is +6 and it is +3 in Cr^{3+}. The oxidation number of each chromium atom changes by 3. As there are two chromium atoms, the total change is 6. This means that there must be six electrons on the left-hand side. So the half equation is: $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

20.8 Half equations of oxidation and reduction (p.64)

Q (Example 20.5)

Write the half equation for the oxidation of sulphite ion to sulphate ion.

A

1 Balance the numbers of all atoms except hydrogen and oxygen.	$SO_3^{2-} \longrightarrow SO_4^{2-}$
2 Add H_2O to the side deficient in oxygen to balance the oxygen atoms.	There is one more oxygen atom on the right-hand side than on the left-hand side, so add $1H_2O$ to the left-hand side: $SO_3^{2-} + H_2O \longrightarrow SO_4^{2-}$
3 Add H^+ to the side deficient in hydrogen to balance the hydrogen atoms.	There are 2 hydrogen atoms on the left-hand side but none on the right-hand side, so add $2H^+$ to the right-hand side: $SO_3^{2-} + H_2O \longrightarrow SO_4^{2-} + 2H^+$
4 Add e^- to the side deficient in negative charge to balance the charge.	The half equation is: $SO_3^{2-}(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2H^+(aq) + xe^-$ The value of x can be worked out in either of two ways: • The charge on the left-hand side is −2, so the charge on the right-hand side is also −2. $(-2) + 2 \times (+1) - x = -2$ $x = 2$ • The oxidation number of sulphur in SO_3^{2-} is +4 and it is +6 in SO_4^{2-}. The oxidation number changes by 2. This means that there must be two electrons on the right-hand side. So the half equation is: $SO_3^{2-}(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$

20.8 Half equations of oxidation and reduction (p.64)

Practice 20.6

Write half equations for the following processes:

a) oxidation of iron to iron(III) ion;

$$Fe(s) \longrightarrow Fe^{3+}(aq) + 3e^{-}$$

b) reduction of nitrate ion to nitrogen monoxide;

$$NO_3^{-}(aq) + 4H^{+}(aq) + 3e^{-} \longrightarrow NO(g) + 2H_2O(l)$$

c) reduction of bromate ion (BrO_3^{-}) to bromine (Br_2).

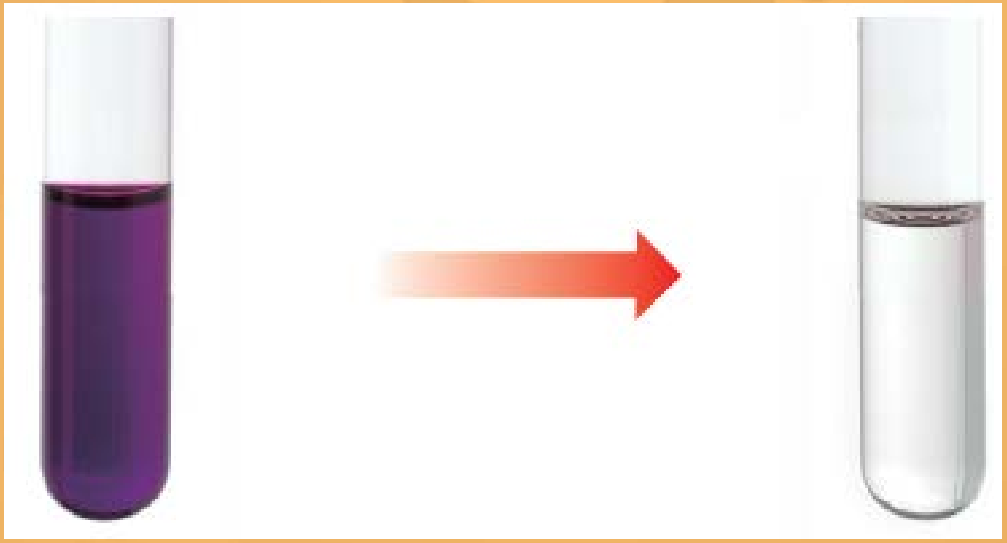
$$2BrO_3^{-}(aq) + 12H^{+}(aq) + 10e^{-} \longrightarrow Br_2(aq) + 6H_2O(l)$$

20.9 Common oxidising agents (p.69)

Acidified potassium permanganate solution

- During a redox reaction, permanganate ions are reduced, under acidic conditions, to manganese(II) ions.

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



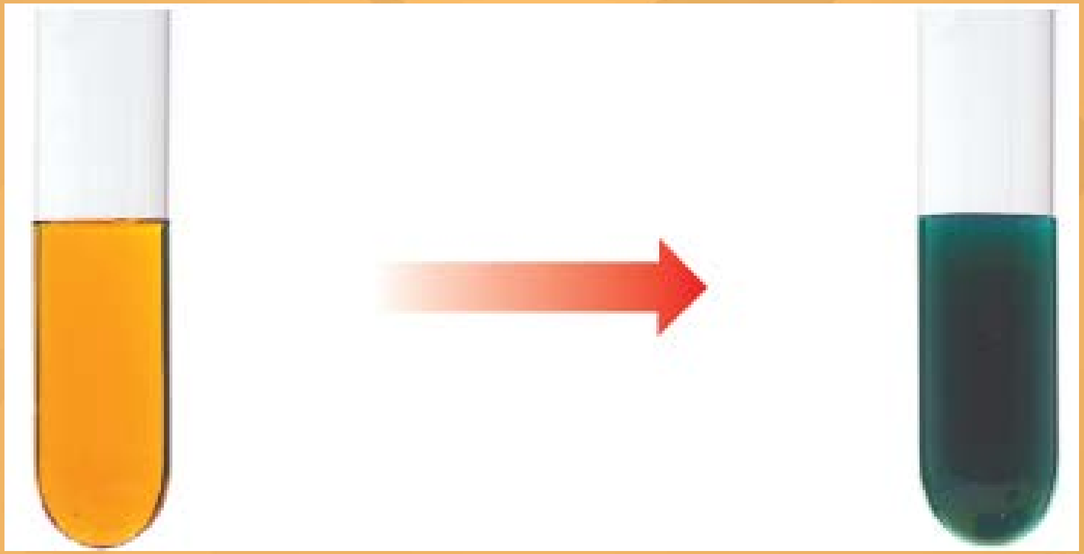
MnO_4^-(aq) ions (left) are reduced to Mn^{2+}(aq) ions (right) during a redox reaction

20.9 Common oxidising agents (p.69)

Acidified potassium dichromate solution

- During a redox reaction, orange dichromate ions are reduced, under acidic conditions, to green chromium(III) ions.

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$



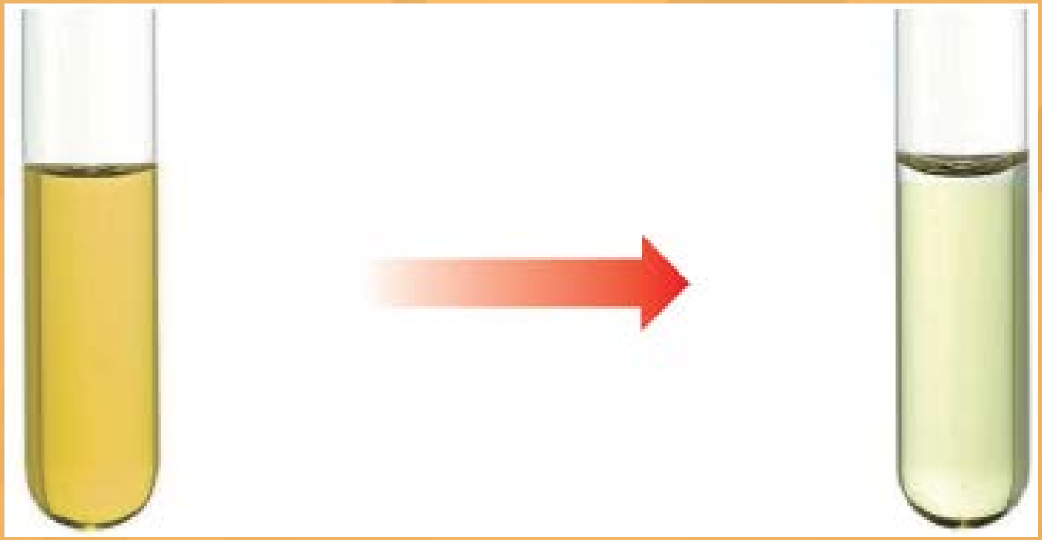
$Cr_2O_7^{2-}$(aq) ions (left) are reduced to Cr^{3+}(aq) ions (right) during a redox reaction

20.9 Common oxidising agents (p.69)

Iron(III) ion

- During a redox reaction, yellow-brown iron(III) ions are reduced to pale green iron(II) ions.

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



Fe^{3+}(aq) ions (left) are reduced to Fe^{2+}(aq) ions (right) during a redox reaction

20.9 Common oxidising agents (p.69)

Aqueous chlorine

- Chlorine is slightly soluble in water and forms a very pale green solution. During a redox reaction, aqueous chlorine is reduced to colourless chloride ions.

$$Cl_2(aq) + 2e^- \longrightarrow 2Cl^-(aq)$$

20.9 Common oxidising agents (p.69)

Dilute nitric acid

- During a redox reaction, colourless dilute nitric acid (about 1.5 mol dm^{-3}) is reduced to colourless nitrogen monoxide gas.

$$NO_3^-(aq) + 4H^+(aq) + 3e^- \longrightarrow NO(g) + 2H_2O(l)$$

- When nitrogen monoxide gas mixes with air, it reacts with oxygen in the air to give brown nitrogen dioxide gas.

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

20.9 Common oxidising agents (p.69)

Concentrated nitric acid

- During a redox reaction, colourless concentrated nitric acid (about 15 mol dm^{-3}) is reduced to brown nitrogen dioxide gas.

$$NO_3^-(aq) + 2H^+(aq) + e^- \longrightarrow NO_2(g) + H_2O(l)$$

Brown nitrogen dioxide gas is evolved when concentrated nitric acid acts as an oxidising agent in a reaction

20.9 Common oxidising agents (p.69)

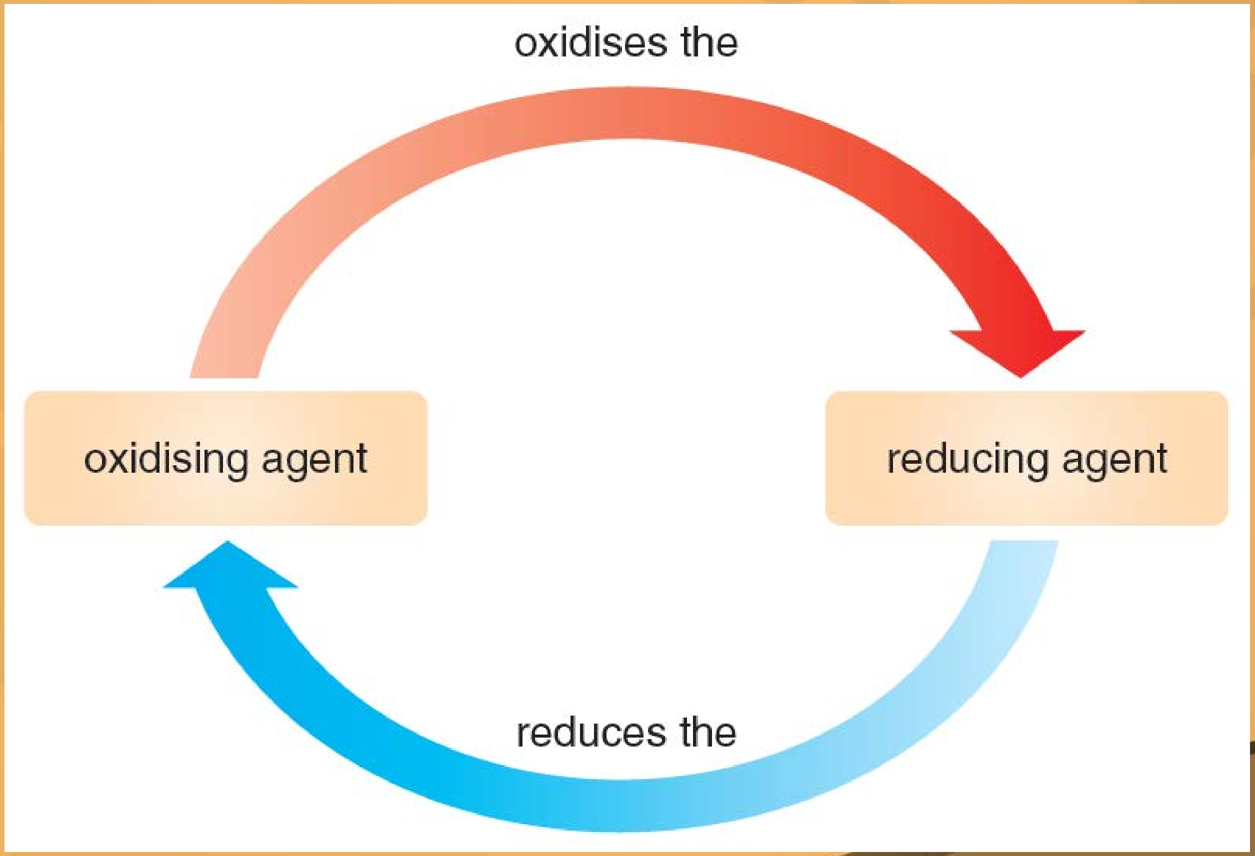
Concentrated sulphuric acid

- During a redox reaction, colourless concentrated sulphuric acid is reduced to colourless sulphur dioxide gas.

$$H_2SO_4(l) + 2H^+(aq) + 2e^- \longrightarrow SO_2(g) + 2H_2O(l)$$

20.10 Common reducing agents (p.72)

- The relationship between oxidising and reducing agents is shown below.



20.10 Common reducing agents (p.72)

Sulphite ion

- During a redox reaction, colourless sulphite ions are oxidised to colourless sulphate ions.

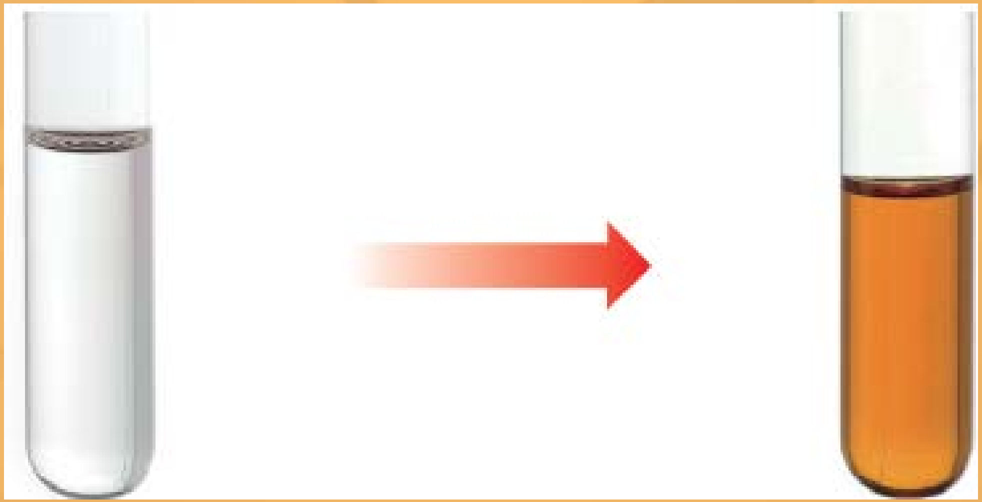
$$SO_3^{2-}(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$$

20.10 Common reducing agents (p.72)

Iodide ion

- During a redox reaction, colourless iodide ions are oxidised to brown iodine.

$$2I^-(aq) \longrightarrow I_2(aq) + 2e^-$$



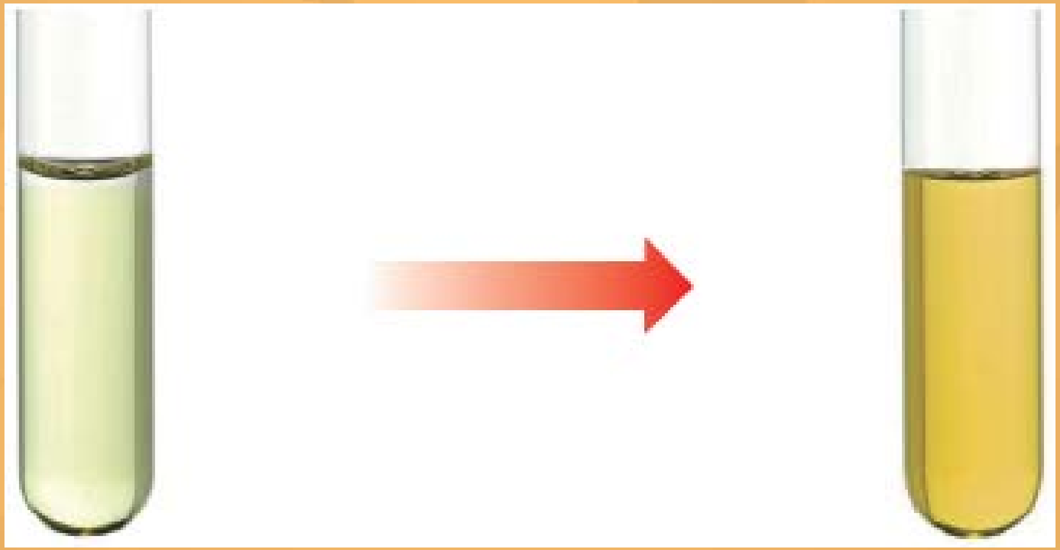
I^-(aq) ions (left) are oxidised to I_2(aq) (right) during a redox reaction

20.10 Common reducing agents (p.72)

Iron(II) ion

- During a redox reaction, pale green iron(II) ions are oxidised to yellow-brown iron(III) ions.

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



Fe^{2+}(aq) ions (left) are oxidised to Fe^{3+}(aq) ions (right) during a redox reaction

20.10 Common reducing agents (p.72)

Zinc

- During a redox reaction, zinc is oxidised to colourless zinc ions.

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

20.10 Common reducing agents (p.72)

Practice 20.7

1 Consider the reaction between solutions containing iron(II) ions and silver ions.

$$Fe^{2+}(aq) + Ag^{+}(aq) \longrightarrow Fe^{3+}(aq) + Ag(s)$$

a) Describe the observations you expect when iron(II) sulphate solution and silver nitrate solution are mixed.

The solution changes from green to yellow / orange / brown.

A grey / black solid forms.

b) State with a reason which chemical species is oxidised in the reaction.

$Fe^{2+}(aq)$ is oxidised.

The oxidation number of Fe increases from +2 to +3.

20.10 Common reducing agents (p.72)

Practice 20.7 (continued)

2 In an experiment, potassium iodide solution is added to acidified potassium dichromate solution until in excess.

a) Identify the oxidising agent and the reducing agent involved.

Oxidising agent: acidified potassium dichromate solution

Reducing agent: potassium iodide solution

b) State the expected observation.

A brown solution forms.

c) Write half equations for the changes involved.

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

$2I^-(aq) \longrightarrow I_2(aq) + 2e^-$

20.11 Balancing redox equations by using half equations (p.74)

- An oxidation half equation can be combined with a reduction half equation to produce an overall redox equation. This is done in two steps:

1 Multiply one or both half equations by integers so that the number of electrons becomes the same in both half equations.

2 Add the two half equations together, and cancel the electrons and anything else that is exactly the same on both sides of the equation.

Investigating redox reactions
Ref.

20.11 Balancing redox equations by using half equations (p.74)

- When aqueous chlorine is added to iron(II) chloride solution, iron(II) ion is oxidised to iron(III) ion and chlorine atoms are reduced to chloride ions.
  The two half equations are:

  $Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + e^-$ oxidation

  $Cl_2(aq) + 2e^- \longrightarrow 2Cl^-(aq)$ reduction

  1 Multiply the first equation by 2 so that the number of electrons becomes the same in both half equations.

  2 Add the two half equations together:

$$2Fe^{2+}(aq) \longrightarrow 2Fe^{3+}(aq) + 2e^-$$
$$Cl_2(aq) + 2e^- \longrightarrow 2Cl^-(aq)$$
$$2Fe^{2+}(aq) + Cl_2(aq) + \cancel{2e^-} \longrightarrow 2Fe^{3+}(aq) + 2Cl^-(aq) + \cancel{2e^-}$$

  Cancel the electrons to obtain the overall redox equation:

$$2Fe^{2+}(aq) + Cl_2(aq) \longrightarrow 2Fe^{3+}(aq) + 2Cl^-(aq)$$

20.11 Balancing redox equations by using half equations (p.74)

Q (Example 20.6)

Acidified permanganate ions oxidise tin(II) ions to tin(IV) ions. In this reaction, acidified permanganate ions are reduced to manganese(II) ions.

The two half equations are:

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

$Sn^{2+}(aq) \longrightarrow Sn^{4+}(aq) + 2e^-$

Deduce the overall redox equation.

A

1 Multiply the first half equation by 2 and the second half equation by 5 so that the number of electrons becomes the same in both.

2 Add the two half equations together:

$$2MnO_4^-(aq) + 16H^+(aq) + 10e^- \longrightarrow 2Mn^{2+}(aq) + 8H_2O(l)$$

$$5Sn^{2+}(aq) \longrightarrow 5Sn^{4+}(aq) + 10e^-$$

$$2MnO_4^-(aq) + 16H^+(aq) + 5Sn^{2+}(aq) + \cancel{10e^-} \longrightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5Sn^{4+}(aq) + \cancel{10e^-}$$

Cancel the electrons to obtain the overall redox equation:

$$2MnO_4^-(aq) + 16H^+(aq) + 5Sn^{2+}(aq) \longrightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5Sn^{4+}(aq)$$

20.11 Balancing redox equations by using half equations (p.74)

Q (Example 20.7)

Acidified dichromate ions oxidise sulphite ions to sulphate ions. In this reaction, acidified dichromate ions are reduced to chromium(III) ions.

The two half equations are:

$SO_3^{2-}(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

Deduce the overall redox equation.

A

1 Multiply the first half equation by 3 so that the number of electrons becomes the same in both.

2 Add the two half equations together:

$$3SO_3^{2-}(aq) + 3H_2O(l) \longrightarrow 3SO_4^{2-}(aq) + 6H^+(aq) + 6e^-$$

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

$$3SO_3^{2-}(aq) + 3H_2O(l) + Cr_2O_7^{2-}(aq) + 14H^+(aq) + \cancel{6e^-} \longrightarrow 3SO_4^{2-}(aq) + 6H^+(aq) + 2Cr^{3+}(aq) + 7H_2O(l) + \cancel{6e^-}$$

20.11 Balancing redox equations by using half equations (p.74)

Q (Example 20.7) (continued)

A

The electrons have now been cancelled, but you are left with an equation that has H^+ and H_2O on both sides in unequal numbers.

- Subtract the $6H^+$ on the right-hand side from the $14H^+$ on the left-hand side to give $8H^+$ on the left-hand side.
- Subtract the $3H_2O$ on the left-hand side from the $7H_2O$ on the right-hand side to give $4H_2O$ on the right-hand side.

The overall redox equation is:

$$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \longrightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$$

20.11 Balancing redox equations by using half equations (p.74)

Practice 20.8

1 Acidified dichromate ions oxidise iodide ions to iodine. In this reaction, acidified dichromate ions are reduced to chromium(III) ions.

The two half equations are:

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

$2I^-(aq) \longrightarrow I_2(aq) + 2e^-$

Deduce the overall redox equation.

1 Multiply the second half equation by 3 so that the number of electrons becomes the same in both equations.

2 Add the two half equations together:

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

$$6I^-(aq) \longrightarrow 3I_2(aq) + 6e^-$$

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6I^-(aq) + \cancel{6e^-} \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3I_2(aq) + \cancel{6e^-}$$

Cancel the electrons to obtain the overall redox equation:

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6I^-(aq) \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3I_2(aq)$$

20.11 Balancing redox equations by using half equations (p.74)

Practice 20.8 (continued)

2 The reaction between acidified potassium permanganate solution and hydrogen peroxide solution involves two changes represented by the following half equations:

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

$H_2O_2(aq) \longrightarrow 2H^+(aq) + O_2(g) + 2e^-$

Deduce the overall redox equation.

1 Multiply the first half equation by 2 and the second half equation by 5 so that the number of electrons becomes the same in both equations.

2 Add the two half equations together:

$$2MnO_4^-(aq) + 16H^+(aq) + 10e^- \longrightarrow 2Mn^{2+}(aq) + 8H_2O(l)$$

$$5H_2O_2(aq) \longrightarrow 10H^+(aq) + 5O_2(g) + 10e^-$$

$$2MnO_4^-(aq) + 16H^+(aq) + 5H_2O_2(aq) + \cancel{10e^-} \longrightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 10H^+(aq) + 5O_2(g) + \cancel{10e^-}$$

20.11 Balancing redox equations by using half equations (p.74)

Practice 20.8 (continued)

The electrons have now been cancelled, but you are left with an equation that has H^+ on both sides in unequal numbers.

Subtract the $10H^+$ on the right-hand side from the $16H^+$ on the left-hand side.

The overall redox equation is:

$$2MnO_4^-(aq) + 6H^+(aq) + 5H_2O_2(aq) \longrightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5O_2(g)$$

20.12 Balancing redox equations by using changes in oxidation numbers (p.77)

- The underlying principle is that the increase in the oxidation number of an element in one reactant must be equal to the decrease in the oxidation number of an element in the other reactant.

20.12 Balancing redox equations by using changes in oxidation numbers (p.77)

The rules are as follows:

1 Identify the elements which undergo a change in oxidation number — one whose oxidation number increases (reducing agent) and the other whose oxidation number decreases (oxidising agent).

2 Calculate the change in oxidation number per atom. Multiply this number with the number of atoms undergoing the change.

3 Equate the increase in oxidation number with decrease in oxidation number on the reactant side by inserting an appropriate coefficient before the chemical formula of each reactant containing an element with a change in oxidation number.

4 Balance the equation with respect to all other atoms except hydrogen and oxygen.

5 For reactions taking place in acid solutions, add H^+ to the side deficient in positive charge.

6 Add H_2O to the appropriate side to balance the number of oxygen atoms.

20.12 Balancing redox equations by using changes in oxidation numbers (p.77)

- When aqueous chlorine is added to iron(II) chloride solution, the iron(II) ions are oxidised to iron(III) ions and the chlorine atoms are reduced to chloride ions.

  The skeleton equation is:

  $Fe^{2+}(aq) + Cl_2(aq) \longrightarrow Fe^{3+}(aq) + Cl^-(aq)$

1 The oxidation number of Fe increases while that of Cl decreases.

2 The oxidation number of Fe increases from +2 to +3, i.e. an increase of one.
The oxidation number of Cl decreases from 0 to –1, i.e. a total decrease of two for two Cl atoms.

3 Equate the increase in oxidation number with decrease in oxidation number on the reactant side by adding a coefficient of '2' before $Fe^{2+}(aq)$.
$2Fe^{2+}(aq) + Cl_2(aq) \longrightarrow Fe^{3+}(aq) + Cl^-(aq)$

4 Balance the equation by adding a coefficient of '2' before $Fe^{3+}(aq)$ and $Cl^-(aq)$.
The overall redox equation is:
$2Fe^{2+}(aq) + Cl_2(aq) \longrightarrow 2Fe^{3+}(aq) + 2Cl^-(aq)$

20.12 Balancing redox equations by using changes in oxidation numbers (p.77)

Q (Example 20.8)

Acidified permanganate ions oxidise tin(II) ions to tin(IV) ions. In this reaction, acidified permanganate ions are reduced to manganese(II) ions.

The skeleton equation is:

$$MnO_4^-(aq) + Sn^{2+}(aq) \longrightarrow Mn^{2+}(aq) + Sn^{4+}(aq)$$

Deduce the overall redox equation.

20.12 Balancing redox equations by using changes in oxidation numbers (p.77)

Q (Example 20.8) (continued)

A

1 Identify the elements which undergo a change in oxidation number.	The oxidation number of Mn decreases while that of Sn increases.
2 Calculate the change in oxidation number.	The oxidation number of Mn decreases from +7 to +2, i.e. a decrease of five. The oxidation number of Sn increases from +2 to +4, i.e. an increase of two.
3 Equate the increase in oxidation number with the decrease in oxidation number on the reactant side.	To equate the change in oxidation numbers, $2MnO_4^-$ and $5Sn^{2+}$ are needed. In this way, the total change in oxidation number is ten for both. $2MnO_4^-(aq) + 5Sn^{2+}(aq) \longrightarrow Mn^{2+}(aq) + Sn^{4+}(aq)$
4 Balance the equation with respect to all other atoms except hydrogen and oxygen.	Add a coefficient of '2' before $Mn^{2+}(aq)$ and a coefficient of '5' before $Sn^{4+}(aq)$. $2MnO_4^-(aq) + 5Sn^{2+}(aq) \longrightarrow 2Mn^{2+}(aq) + 5Sn^{4+}(aq)$
5 Add H^+ to the side deficient in positive charge.	Total charge on the left-hand side = 2 x (–1) + 5 x (+2) = +8 Total charge on the right-hand side = 2 x (+2) + 5 x (+4) = +24 Add $16H^+$ to the left-hand side. $2MnO_4^-(aq) + 5Sn^{2+}(aq) + 16H^+(aq) \longrightarrow 2Mn^{2+}(aq) + 5Sn^{4+}(aq)$
6 Add H_2O to balance the number of oxygen atoms.	Add $8H_2O$ to the right-hand side. The overall redox equation is: $2MnO_4^-(aq) + 5Sn^{2+}(aq) + 16H^+(aq) \longrightarrow 2Mn^{2+}(aq) + 5Sn^{4+}(aq) + 8H_2O(l)$

20.12 Balancing redox equations by using changes in oxidation numbers (p.77)

Q (Example 20.9)

Acidified dichromate ions oxidise sulphite ions to sulphate ions. In this reaction, acidified dichromate ions are reduced to chromium(III) ions.

The skeleton equation is:

$$Cr_2O_7^{2-}(aq) + SO_3^{2-}(aq) \longrightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$$

Deduce the overall redox equation.

20.12 Balancing redox equations by using changes in oxidation numbers (p.77)

Q (Example 20.9) (continued)

A

1 Identify the elements which undergo a change in oxidation number.	The oxidation number of Cr decreases while that of S increases.
2 Calculate the change in oxidation number.	The oxidation number of Cr decreases from +6 to +3, i.e. a total decrease of six for two Cr atoms. The oxidation number of S increases from +4 to +6, i.e. an increase of two.
3 Equate the increase in oxidation number with the decrease in oxidation number on the reactant side.	To equate the change in oxidation numbers, $3SO_3^{2-}$ is needed. In this way, the total change in oxidation number is six for both. $Cr_2O_7^{2-}(aq) + 3SO_3^{2-}(aq) \longrightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$
4 Balance the equation with respect to all other atoms except hydrogen and oxygen.	Add a coefficient of '2' before $Cr^{3+}(aq)$ and a coefficient of '3' before $SO_4^{2-}(aq)$. $Cr_2O_7^{2-}(aq) + 3SO_3^{2-}(aq) \longrightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq)$
5 Add H^+ to the side deficient in positive charge.	Total charge on the left-hand side = (–2) + 3 x (–2) = –8 Total charge on the right-hand side = 2 x (+3) + 3 x (–2) = 0 Add $8H^+$ to the left-hand side. $Cr_2O_7^{2-}(aq) + 3SO_3^{2-}(aq) + 8H^+(aq) \longrightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq)$
6 Add H_2O to balance the number of oxygen atoms.	Add $4H_2O$ to the right-hand side. The overall redox equation is: $Cr_2O_7^{2-}(aq) + 3SO_3^{2-}(aq) + 8H^+(aq) \longrightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq) + 4H_2O(l)$

20.12 Balancing redox equations by using changes in oxidation numbers (p.77)

Practice 20.9

1 Balance the following redox equations:

a) $Zn(s) + \underline{\ 2\ }V^{3+}(aq) \longrightarrow Zn^{2+}(aq) + \underline{\ 2\ }V^{2+}(aq)$

b) $2IO_3^-(aq) + \underline{\ 5\ }H_2O_2(aq) + \underline{\ 2\ }H^+(aq) \longrightarrow I_2(aq) + \underline{\ 5\ }O_2(g) + \underline{\ 6\ }H_2O(l)$

20.12 Balancing redox equations by using changes in oxidation numbers (p.77)

Practice 20.9 (continued)

2 Acidified dichromate ions oxidise iron(II) ions to iron(III) ions.
In this reaction, acidified dichromate ions are reduced to chromium(III) ions.
The skeleton equation is:

$$Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \longrightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$$

Deduce the overall redox equation.

20.12 Balancing redox equations by using changes in oxidation numbers (p.77)

Practice 20.9 (continued)

1 Identify the elements which undergo a change in oxidation number.	The oxidation number of Cr decreases while that of Fe increases.
2 Calculate the change in oxidation number.	The oxidation number of Cr decreases from +6 to +3, i.e. a total decrease of six for two Cr atoms. The oxidation number of Fe increases from +2 to +3, i.e. an increase of one.
3 Equate the increase in oxidation number with the decrease in oxidation number on the reactant side.	To equate the change in oxidation numbers, $6Fe^{2+}$ is needed. In this way, the total change in oxidation number is six for both. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) \longrightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$
4 Balance the equation with respect to all other atoms except hydrogen and oxygen.	Add a coefficient of '2' before $Cr^{3+}(aq)$ and a coefficient of '6' before Fe^{3+}. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) \longrightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq)$
5 Add H^+ to the side deficient in positive charge.	Total charge on the left-hand side = (–2) + 6 x (+2) = +10 Total charge on the right-hand side = 2 x (+3) + 6 x (+3) = +24 Add $14H^+$ to the left-hand side. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) + 14H^+(aq) \longrightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq)$
6 Add H_2O to balance the number of oxygen atoms.	Add $7H_2O$ to the right-hand side. The overall redox equation is: $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) + 14H^+(aq) \longrightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(l)$

20.13 Deducing half equations from the overall redox equation (p.81)

- Half equations can be deduced from the overall redox equation because the latter shows the chemical formulae of all the reactants and products. Consider the reaction between iodine and thiosulphate ions:

  $$I_2(aq) + 2S_2O_3^{2-}(aq) \longrightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

  The oxidation number of iodine decreases so iodine is reduced. The half equation is:

  $$I_2(aq) + 2e^- \longrightarrow 2I^-(aq)$$

  Thus the thiosulphate ions must be oxidised in the reaction. The number of electrons on the right-hand side of the half equation for the oxidation of thiosulphate ions must be the same as that on the left-hand side of the half equation for the reduction of iodine. Hence the half equation for the oxidation of thiosulphate ions is:

  $$2S_2O_3^{2-}(aq) \longrightarrow S_4O_6^{2-}(aq) + 2e^-$$

20.13 Deducing half equations from the overall redox equation (p.81)

Practice 20.10

Rewrite each of the following overall redox equations as two half equations. In each case, identify which chemical species is oxidised and which is reduced.

a) $2I^-(aq) + Cl_2(aq) \longrightarrow I_2(aq) + 2Cl^-(aq)$

$Cl_2(aq) + 2e^- \longrightarrow 2Cl^-(aq)$

$2I^-(aq) \longrightarrow I_2(aq) + 2e^-$

$I^-(aq)$ is oxidised.

$Cl_2(aq)$ is reduced.

b) $HSO_3^-(aq) + 2Fe^{3+}(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2Fe^{2+}(aq) + 3H^+(aq)$

$HSO_3^-(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 3H^+(aq) + 2e^-$

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

HSO_3^- is oxidised.

$Fe^{3+}(aq)$ is reduced.

20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

- All the chemical species on the left-hand side of the E.C.S. can accept electrons and be reduced. Hence they are all oxidising agents.
- All the chemical species on the right-hand side of the E.C.S. can lose electrons and be oxidised. Hence they are all reducing agents.
- The oxidising power of oxidising agents increases down the series while the reducing power of reducing agents decreases down the series.

All oxidising agents should oxidise the reduced form of any chemical species above them in the electrochemical series.

20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

Half equation								
Oxidised form						**Reduced form**		
$Li^{+}(aq)$			+	e^{-}	$\rightleftharpoons$	$Li(s)$		
$K^{+}(aq)$			+	e^{-}	$\rightleftharpoons$	$K(s)$		
$Ca^{2+}(aq)$			+	$2e^{-}$	$\rightleftharpoons$	$Ca(s)$		
$Na^{+}(aq)$			+	e^{-}	$\rightleftharpoons$	$Na(s)$		
$Mg^{2+}(aq)$			+	$2e^{-}$	$\rightleftharpoons$	$Mg(s)$		
$Al^{3+}(aq)$			+	$3e^{-}$	$\rightleftharpoons$	$Al(s)$		
$Zn^{2+}(aq)$			+	$2e^{-}$	$\rightleftharpoons$	$Zn(s)$		
$Fe^{2+}(aq)$			+	$2e^{-}$	$\rightleftharpoons$	$Fe(s)$		
$Pb^{2+}(aq)$			+	$2e^{-}$	$\rightleftharpoons$	$Pb(s)$		
$2H^{+}(aq)$			+	$2e^{-}$	$\rightleftharpoons$	$H_2(g)$		
$H_2SO_4(l)$	+	$2H^{+}(aq)$	+	$2e^{-}$	$\rightleftharpoons$	$SO_2(g)$	+	$2H_2O(l)$
$Cu^{2+}(aq)$			+	$2e^{-}$	$\rightleftharpoons$	$Cu(s)$		
$O_2(g)$	+	$2H_2O(l)$	+	$4e^{-}$	$\rightleftharpoons$	$4OH^{-}(aq)$		
$I_2(aq)$			+	$2e^{-}$	$\rightleftharpoons$	$2I^{-}(aq)$		
$Fe^{3+}(aq)$			+	e^{-}	$\rightleftharpoons$	$Fe^{2+}(aq)$		
$Ag^{+}(aq)$			+	e^{-}	$\rightleftharpoons$	$Ag(s)$		
$NO_3^{-}(aq)$	+	$2H^{+}(aq)$	+	e^{-}	$\rightleftharpoons$	$NO_2(g)$	+	$H_2O(l)$
$NO_3^{-}(aq)$	+	$4H^{+}(aq)$	+	$3e^{-}$	$\rightleftharpoons$	$NO(g)$	+	$2H_2O(l)$
$Br_2(aq)$			+	$2e^{-}$	$\rightleftharpoons$	$2Br^{-}(aq)$		
$Cr_2O_7^{2-}(aq)$	+	$14H^{+}(aq)$	+	$6e^{-}$	$\rightleftharpoons$	$2Cr^{3+}(aq)$	+	$7H_2O(l)$
$Cl_2(aq)$			+	$2e^{-}$	$\rightleftharpoons$	$2Cl^{-}(aq)$		
$MnO_4^{-}(aq)$	+	$8H^{+}(aq)$	+	$5e^{-}$	$\rightleftharpoons$	$Mn^{2+}(aq)$	+	$4H_2O(l)$
$F_2(g)$			+	$2e^{-}$	$\rightleftharpoons$	$2F^{-}(aq)$		

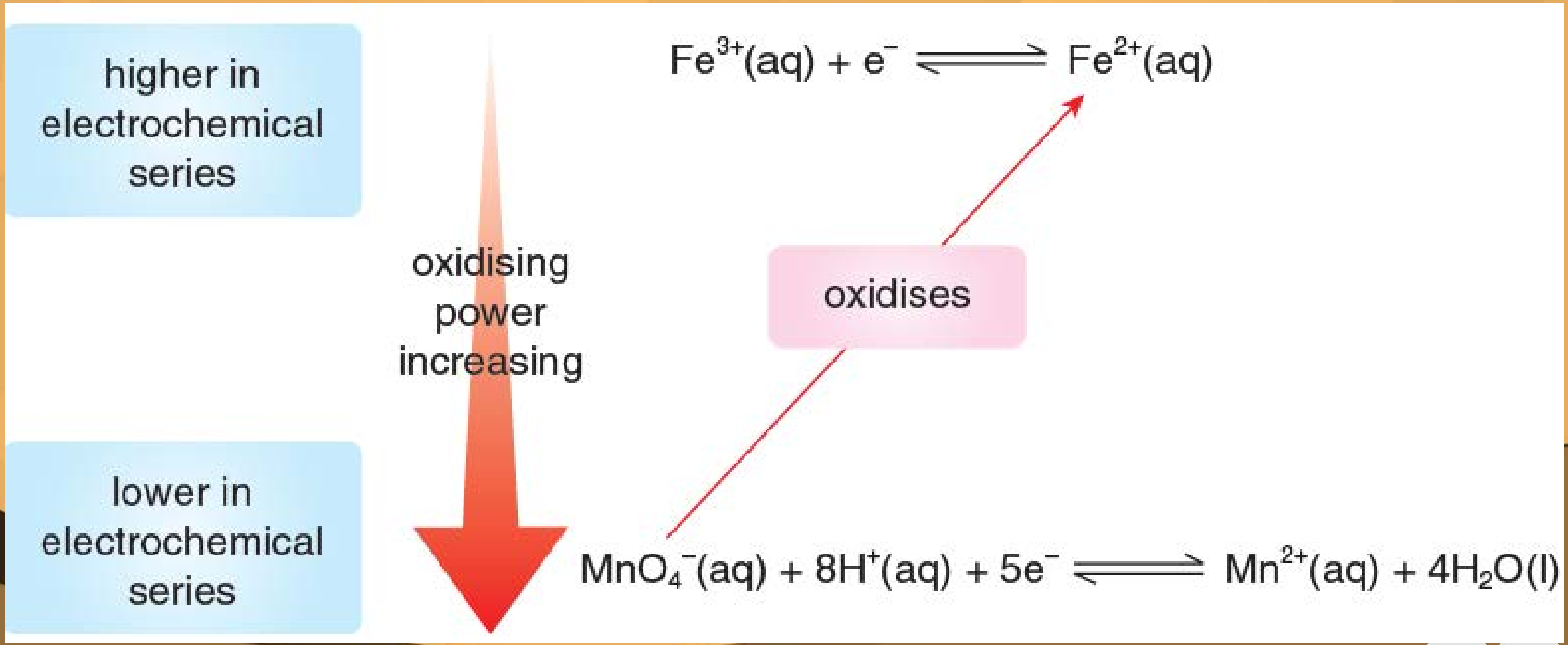
oxidising power increasing

reducing power increasing

20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

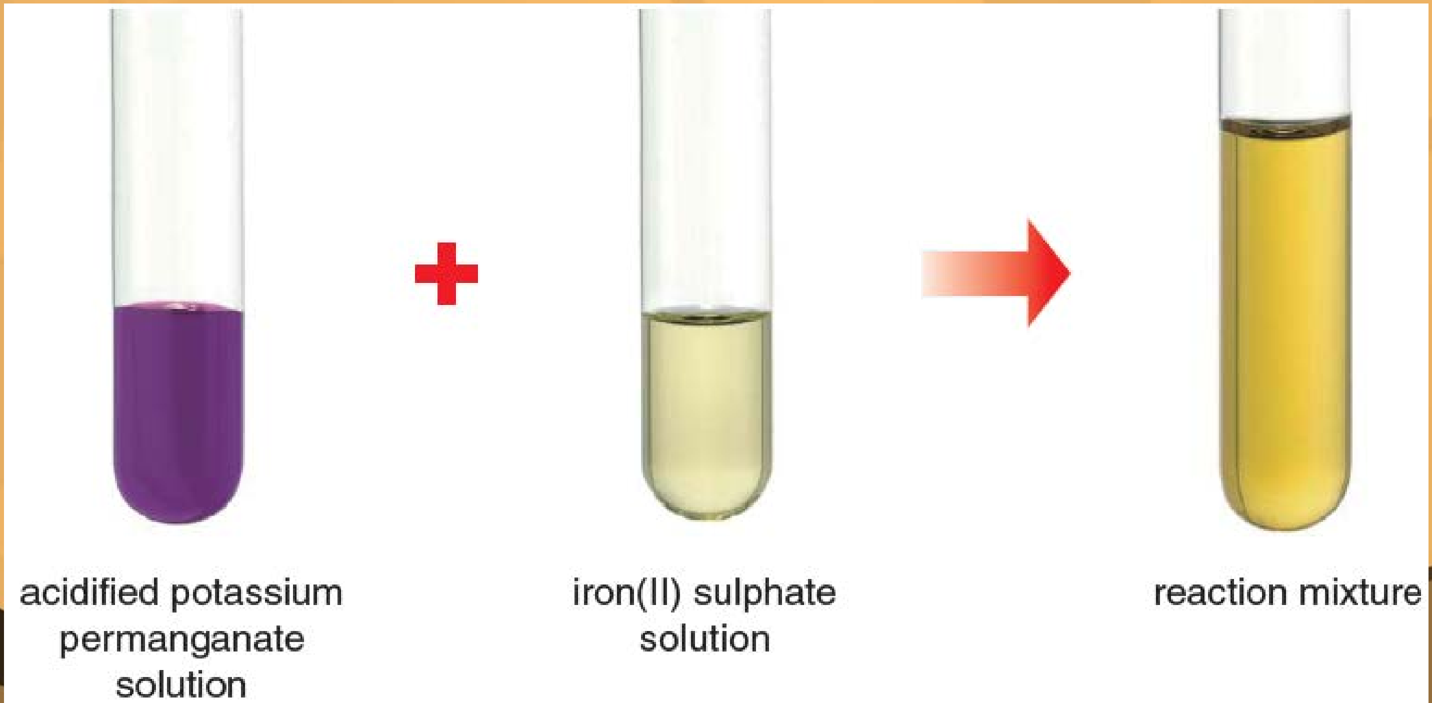
Will acidified potassium permanganate solution react with iron(II) sulphate solution?

- As permanganate ion is a stronger oxidising agent than iron(III) ion, it will oxidise iron(II) ion to iron(III) ion.



20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

- In this reaction, the purple permanganate ions are reduced to pale pink manganese(II) ions and the pale green iron(II) ions are oxidised to yellow-brown iron(III) ions.



20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

- The following two half equations represent the changes that occur:

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

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

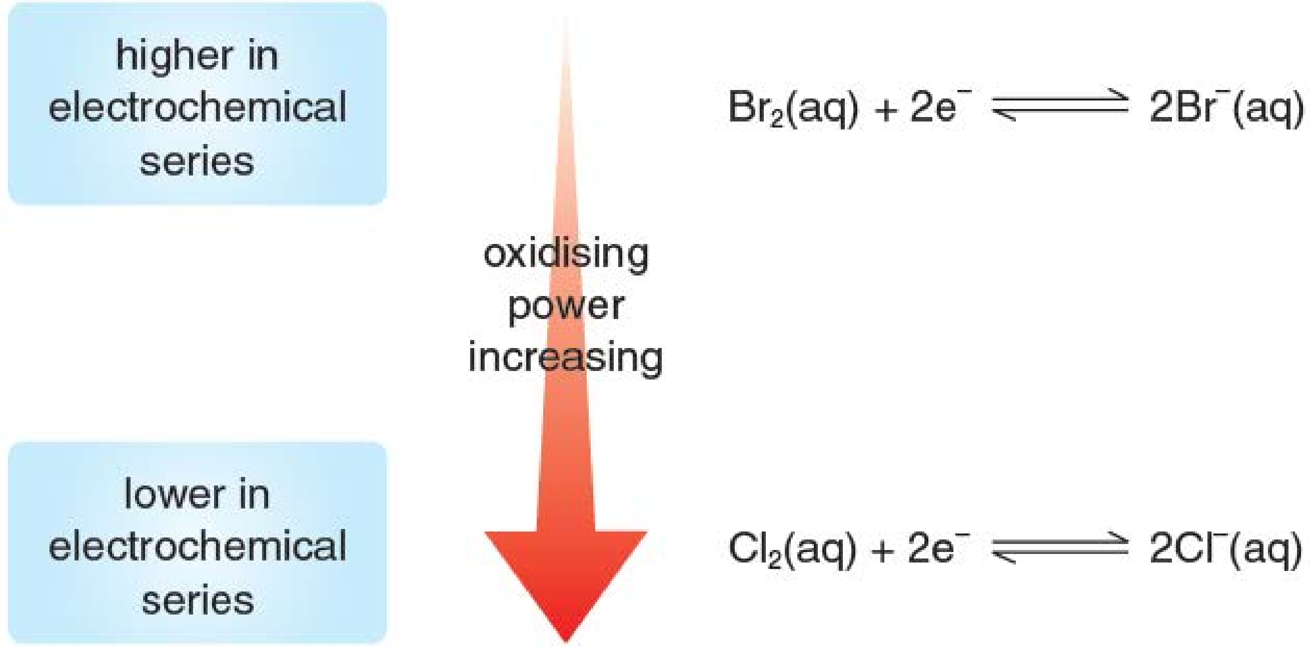
  Combination of the half equations gives the redox equation for the reaction:

  $MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \longrightarrow Mn^{2+}(aq) + 4H_2O(l) + 5Fe^{3+}(aq)$

20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

Will aqueous bromine react with potassium chloride solution?

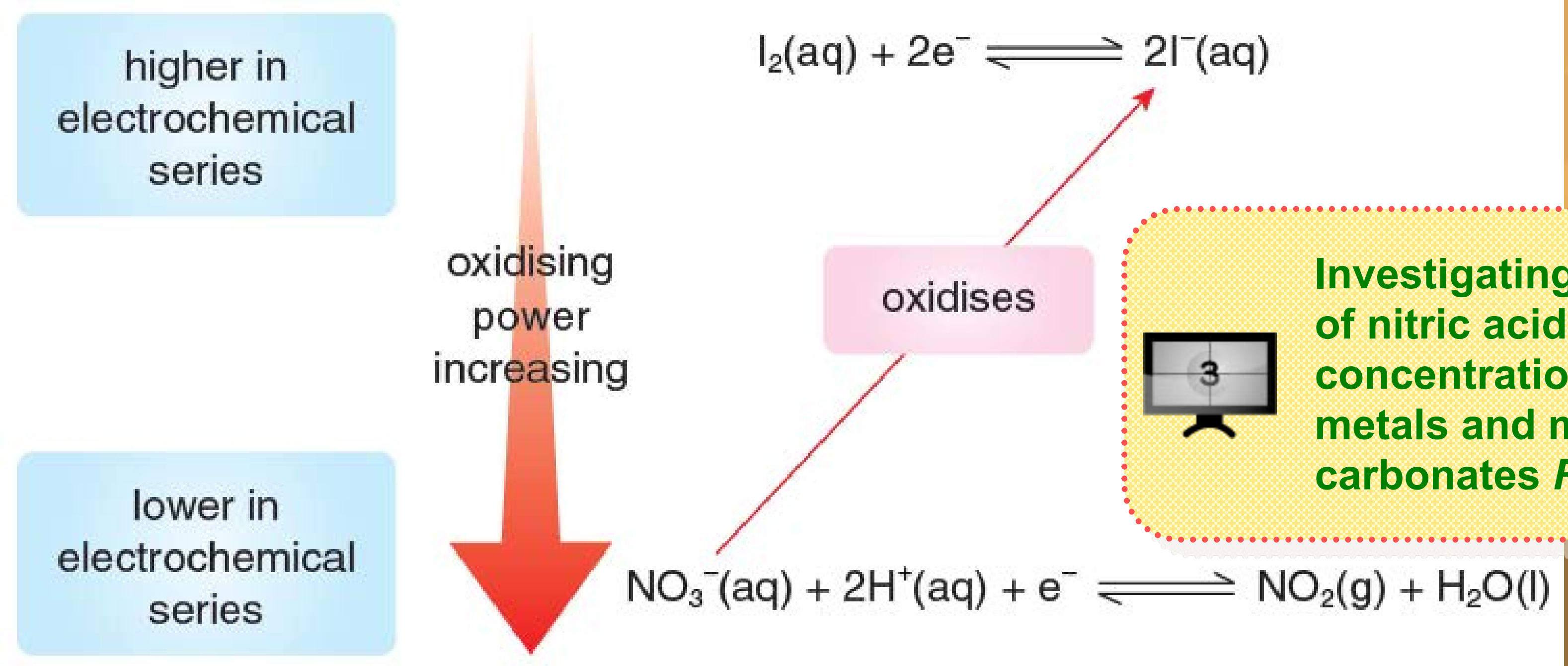
- As aqueous bromine is a weaker oxidising agent than aqueous chlorine, it cannot oxidise chloride ion.
  Hence there is no reaction between aqueous bromine and potassium chloride solution.



20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

Will concentrated nitric acid react with potassium iodide solution?

- As concentrated nitric acid is a stronger oxidising agent than aqueous iodine, it will oxidise iodide ion to iodine.



Investigating the action of nitric acid of different concentrations on metals and metal carbonates *Ref.*

20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

- In this reaction, the concentrated nitric acid is reduced to brown nitrogen dioxide and the colourless iodide ions are oxidised to brown iodine.

A brown gas is evolved when concentrated nitric acid reacts with potassium iodide solution

20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

- The following two half equations represent the changes that occur:

  $$2I^-(aq) \longrightarrow I_2(aq) + 2e^-$$

  $$NO_3^-(aq) + 2H^+(aq) + e^- \longrightarrow NO_2(g) + H_2O(l)$$

  Combination of the half equations gives the redox equation for the reaction:

  $$2I^-(aq) + 2NO_3^-(aq) + 4H^+(aq) \longrightarrow I_2(aq) + 2NO_2(g) + 2H_2O(l)$$

20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

Practice 20.11

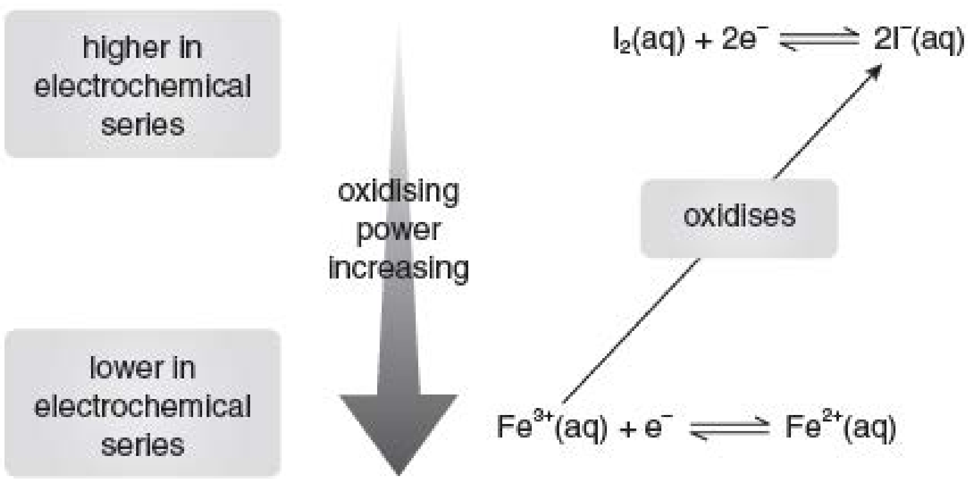
a) Based on the electrochemical series, predict whether each of the following mixtures would react. Explain your answers.
i) iron(III) sulphate solution + potassium iodide solution
ii) copper + dilute hydrochloric acid
iii) zinc + dilute nitric acid

b) For each mixture that reacts,
i) state the expected observations;
ii) deduce the overall redox equation.

20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

Practice 20.11 (continued)

a) and b) i) Look at the relative positions of the relevant half equations in the electrochemical series. As iron(III) ion is a stronger oxidising agent than iodine, it will oxidise iodide ion to iodine.



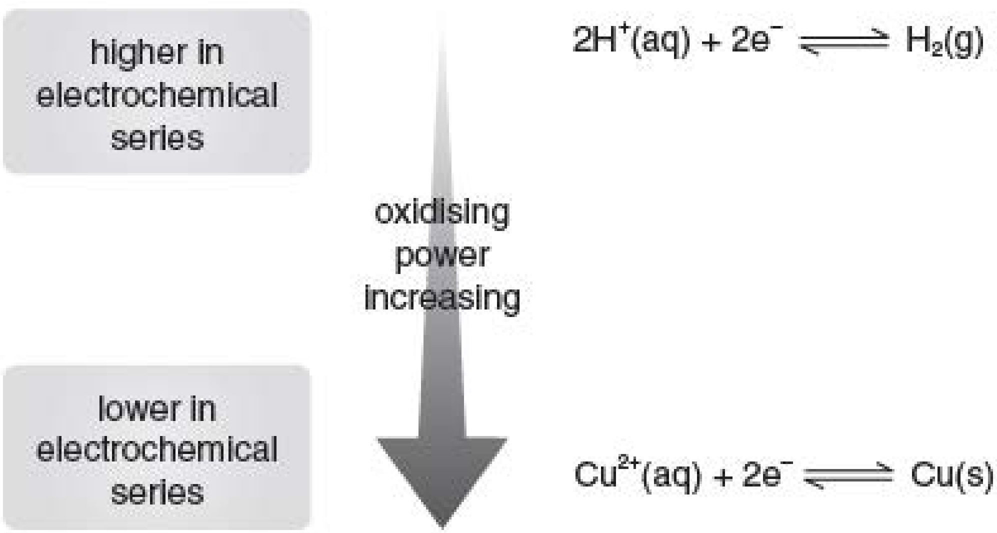
In this reaction, the yellow-brown iron(III) ions are reduced to pale green iron(II) ions. The colourless iodide ions are oxidised to brown iodine. Hence a brown solution results.

$$2Fe^{3+}(aq) + 2I^-(aq) \longrightarrow 2Fe^{2+}(aq) + I_2(aq)$$

20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

Practice 20.11 (continued)

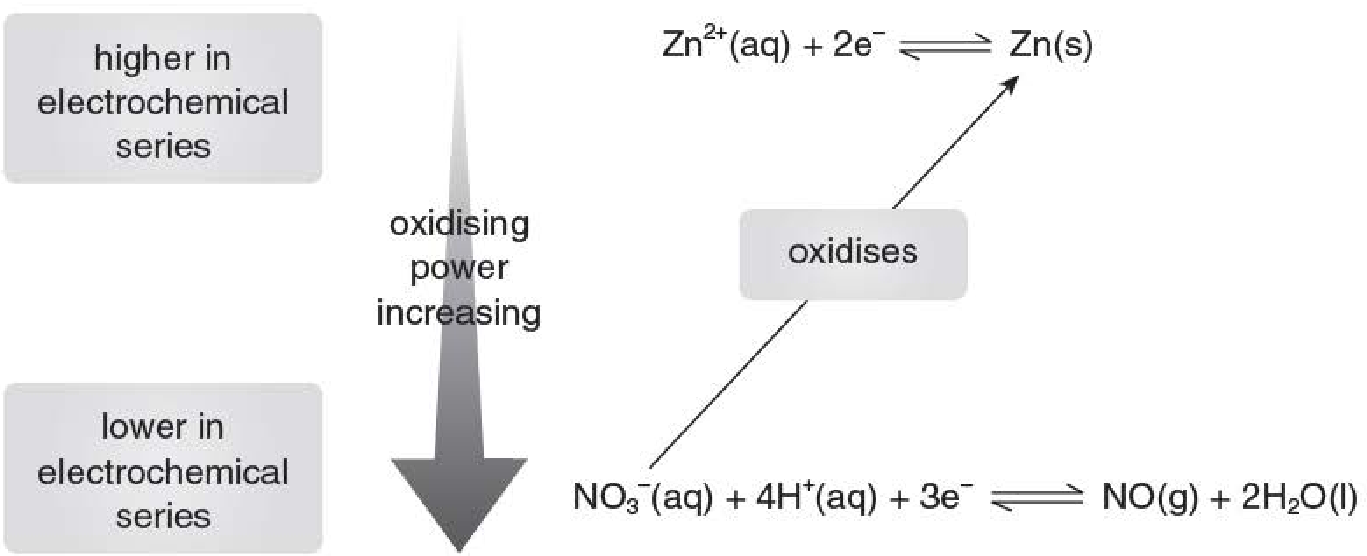
ii) Look at the relative positions of the relevant half equations in the electrochemical series. As $H^+(aq)$ ion is a weaker oxidising agent than copper(II) ion, it cannot oxidise copper. Hence there is no reaction between copper and dilute hydrochloric acid.



20.14 How are the relative oxidising and reducing power of chemical species related to their positions in the electrochemical series (p.82)

Practice 20.11 (continued)

iii) Look at the relative positions of the relevant half equations in the electrochemical series. As dilute nitric acid is a stronger oxidising agent than zinc ion, it will oxidise zinc to zinc ion.



20.15 Chlorine as an oxidising agent (p.87)

- Chlorine is a strong oxidising agent. During a redox reaction, it is reduced to chloride ions.

$$Cl_2(aq) + 2e^- \longrightarrow 2Cl^-(aq)$$

Ranking halogens according to their oxidising power ***Ref.***

20.15 Chlorine as an oxidising agent (p.87)

Reaction with halide ions

- When aqueous chlorine is added to potassium bromide solution, a yellow-brown colour appears because bromine is formed.
- Chlorine is a stronger oxidising agent than bromine. It oxidises bromide ions to bromine.

$$Cl_2(aq) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$$

Aqueous chlorine reacts with potassium bromide solution to form bromine

20.15 Chlorine as an oxidising agent (p.87)

- When aqueous chlorine is added to potassium iodide solution, a brown colour appears because iodine is formed.
- Chlorine is a stronger oxidising agent than iodine. It oxidises iodide ions to iodine.

$$Cl_2(aq) + 2I^-(aq) \longrightarrow 2Cl^-(aq) + I_2(aq)$$

Aqueous chlorine reacts with potassium iodide solution to form iodine

20.15 Chlorine as an oxidising agent (p.87)

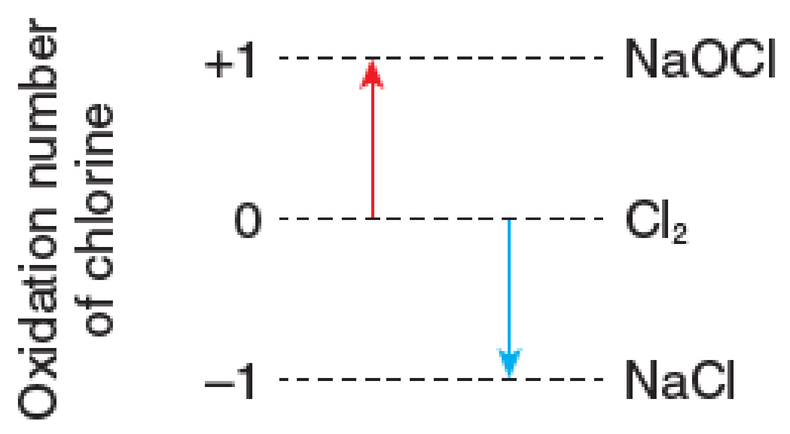
Reaction with alkalis

With alkalis at room temperature

- Chlorine reacts with cold sodium hydroxide solution to give a mixture of sodium chloride and sodium hypochlorite.

$$\underset{0}{Cl_2}(aq) + 2NaOH(aq) \longrightarrow Na\underset{-1}{Cl}(aq) + NaO\underset{+1}{Cl}(aq) + H_2O(l)$$

- Chlorine itself, being an element, has an oxidation number of 0. In sodium chloride, it has an oxidation number of –1. In sodium hypochlorite, it has an oxidation number of +1.



The oxidation numbers of Cl in NaCl and NaOCl are –1 and +1 respectively

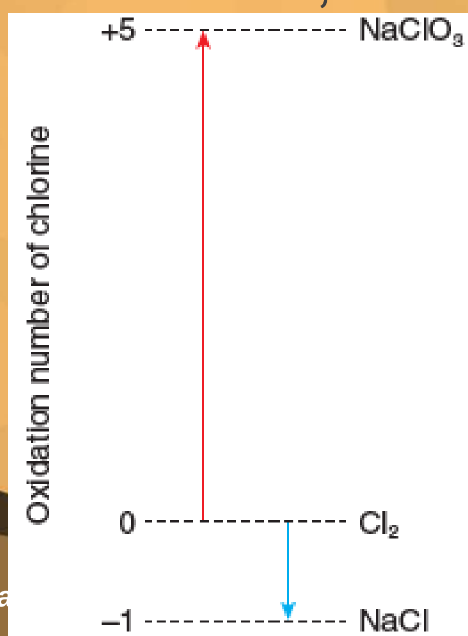
20.15 Chlorine as an oxidising agent (p.87)

With hot alkalis

- Chlorine reacts with hot and concentrated sodium hydroxide solution to give a mixture of sodium chloride and sodium chlorate.

$$\underset{0}{Cl_2}(aq) + 6NaOH(aq) \longrightarrow 5Na\underset{-1}{Cl}(aq) + Na\underset{+5}{Cl}O_3(aq) + 3H_2O(l)$$

- In sodium chlorate, chlorine has an oxidation number of +5.



The oxidation numbers of Cl in NaCl and $NaClO_3$ are –1 and +5 respectively

20.15 Chlorine as an oxidising agent (p.87)

- These reactions of chlorine with sodium hydroxide solution are examples of **disproportionation (歧化作用)**. This term is applied to a reaction in which one element has been both oxidised and reduced, so that its oxidation number goes up and down at the same time.

20.16 Hydrogen halides as reducing agents (p.91)

- Hydrogen halides can act as reducing agents. The strength of the hydrogen halides as reducing agents is in the order: HI > HBr > HCl.

- This trend can be seen in the reactions of solid sodium halides with concentrated sulphuric acid.

20.16 Hydrogen halides as reducing agents (p.91)

Reaction of sodium halides with concentrated sulphuric acid

- When concentrated sulphuric acid is added to a solid sodium halide, hydrogen halide is formed first.

$$NaX(s) + H_2SO_4(l) \longrightarrow NaHSO_4(s) + HX(g)$$

- The hydrogen halide produced might then reduce the concentrated sulphuric acid.

20.16 Hydrogen halides as reducing agents (p.91)

Sodium chloride (solid)

- Sodium chloride reacts with concentrated sulphuric acid to give steamy fumes of hydrogen chloride.

$$NaCl(s) + H_2SO_4(l) \longrightarrow NaHSO_4(s) + HCl(g)$$

- This is not a redox reaction because no oxidation number has changed. The hydrogen chloride is too weak a reducing agent to reduce the sulphuric acid.

20.16 Hydrogen halides as reducing agents (p.91)

Sodium bromide (solid)

- When concentrated sulphuric acid is added to sodium bromide, steamy fumes of hydrogen bromide and brown fumes of bromine are given off. Colourless sulphur dioxide is also formed.

Steamy fumes of hydrogen bromide and brown fumes of bromine are produced when concentrated sulphuric acid reacts with sodium bromide

20.16 Hydrogen halides as reducing agents (p.91)

- First, sodium hydrogensulphate and hydrogen bromide are produced.

$$NaBr(s) + H_2SO_4(l) \longrightarrow NaHSO_4(s) + HBr(g)$$

- Hydrogen bromide is a strong enough reducing agent to reduce the sulphuric acid to sulphur dioxide.

$$\underset{-1}{2HBr(g)} + \underset{+6}{H_2SO_4(l)} \longrightarrow \underset{0}{Br_2(g)} + \underset{+4}{SO_2(g)} + 2H_2O(l)$$

- The oxidation number of sulphur decreases from +6 to +4 while that of bromine increases from –1 to 0.
- Only a little hydrogen bromide is found. The principal products are bromine and sulphur dioxide.

20.16 Hydrogen halides as reducing agents (p.91)

Sodium iodide (solid)

- When concentrated sulphuric acid is added to sodium iodide, steamy fumes of hydrogen iodide and a black solid of iodine are produced.
- Yellow solid sulphur may be seen and the bad egg smell of hydrogen sulphide gas is present. Colourless sulphur dioxide gas is also evolved.

Steamy fumes of hydrogen iodide and a black solid of iodine are formed when concentrated sulphuric acid reacts with sodium iodide

20.16 Hydrogen halides as reducing agents (p.91)

Investigating the properties of dilute and concentrated sulphuric acid *Ref.*

- First, hydrogen iodide is produced.

$$NaI(s) + H_2SO_4(l) \longrightarrow NaHSO_4(s) + HI(g)$$

- Hydrogen iodide is a stronger reducing agent than hydrogen bromide. The sulphuric acid is reduced not only to Sulphur dioxide, but further to sulphur (oxidation number = 0), and even to hydrogen sulphide (oxidation number of sulphur = –2).

$$2H\underset{-1}{I}(g) + H_2\underset{+6}{S}O_4(l) \longrightarrow \underset{0}{I}_2(s) + \underset{+4}{S}O_2(g) + 2H_2O(l)$$

$$6H\underset{-1}{I}(g) + H_2\underset{+6}{S}O_4(l) \longrightarrow 3\underset{0}{I}_2(s) + \underset{0}{S}(s) + 4H_2O(l)$$

$$8H\underset{-1}{I}(g) + H_2\underset{+6}{S}O_4(l) \longrightarrow 4\underset{0}{I}_2(s) + H_2\underset{-2}{S}(g) + 4H_2O(l)$$

20.16 Hydrogen halides as reducing agents (p.91)

- Hydrogen iodide is oxidised so easily that only a trace of it is found.
- Concentrated sulphuric acid acts as an oxidising agent in its reactions with sodium bromide and sodium iodide.

20.16 Hydrogen halides as reducing agents (p.91)

Practice 20.12

1 A student investigated the chemistry of chlorine and the halide ions. The student made the following observations in two tests.

Test	Observation
1 Add aqueous chlorine to potassium bromide solution.	The colourless solution turned yellow-brown.
2 Add silver nitrate solution to potassium chloride solution.	The colourless solution produced a white precipitate.

a) i) Identify the chemical species responsible for the yellow-brown colour in *Test 1*. Bromine

ii) Write the ionic equation for the reaction that occurred.

$Cl_2(aq) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$

b) i) Identify the chemical species responsible for the white precipitate in *Test 2*. Silver chloride

ii) Write the ionic equation for the reaction that occurred.

$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$

20.16 Hydrogen halides as reducing agents (p.91)

Practice 20.12 (continued)

2 a) Identify TWO sulphur-containing reduction products formed when concentrated sulphuric acid oxidises hydrogen iodide.

Any two of the following:
- SO_2
- S
- H_2S

b) For each of the reduction products identified in (a), write an equation to illustrate its formation from concentrated sulphuric acid.

Any two of the following:
- $2HI(g) + H_2SO_4(l) \longrightarrow I_2(s) + SO_2(g) + 2H_2O(l)$
- $6HI(g) + H_2SO_4(l) \longrightarrow 3I_2(s) + S(s) + 4H_2O(l)$
- $8HI(g) + H_2SO_4(l) \longrightarrow 4I_2(s) + H_2S(g) + 4H_2O(l)$

20.17 Aqueous sulphur dioxide as a reducing agent (p.93)

- Sulphur dioxide is a colourless toxic gas with a choking acidic smell.
- Aqueous sulphur dioxide is a good reducing agent. Sulphur dioxide dissolves in water to form sulphurous acid.

$$SO_2(g) + H_2O(l) \rightleftharpoons H_2SO_3(aq)$$

- The reducing properties of aqueous sulphur dioxide are due to the presence of sulphite ions (SO_3^{2-}) from sulphurous acid.
- In a redox reaction, the sulphite ions are oxidised to sulphate ions (SO_4^{2-}). The oxidation number of sulphur increases from +4 to +6.

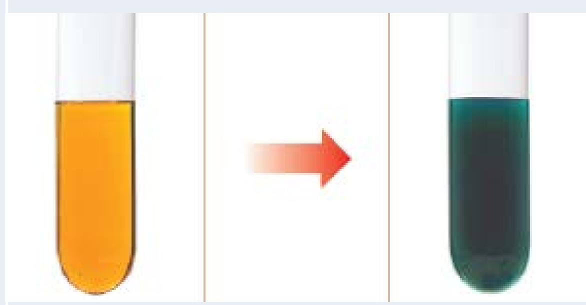
$$\underset{+4}{SO_3^{2-}}(aq) + H_2O(l) \longrightarrow \underset{+6}{SO_4^{2-}}(aq) + 2H^+(aq) + 2e^-$$

20.17 Aqueous sulphur dioxide as a reducing agent (p.93)

- Sulphur dioxide reduces the following solutions:
  - acidified potassium permanganate solution;
  - acidified potassium dichromate solution;
  - iron(III) sulphate solution;
  - aqueous bromine.

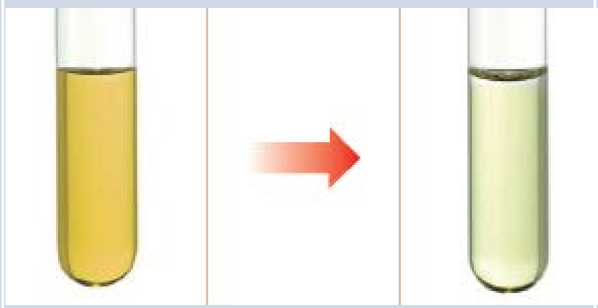
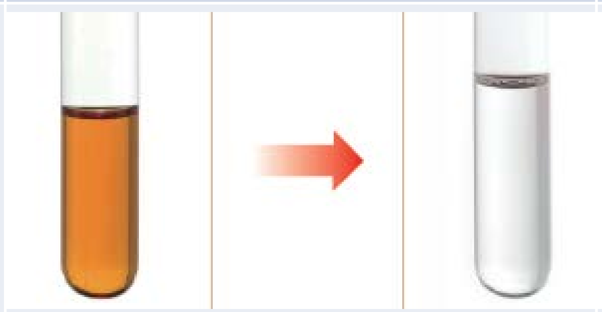
20.17 Aqueous sulphur dioxide as a reducing agent (p.93)

- The colour change as sulphur dioxide gas is bubbled into different solutions.

Sulphur dioxide gas is bubbled into	Colour change	Equation
acidified potassium permanganate solution		$5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \longrightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$
acidified potassium dichromate solution		$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \longrightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$

20.17 Aqueous sulphur dioxide as a reducing agent (p.93)

- The colour change as sulphur dioxide gas is bubbled into different solutions.

Sulphur dioxide gas is bubbled into	Colour change	Equation
iron(III) sulphate solution		$SO_3^{2-}(aq) + 2Fe^{3+}(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2Fe^{2+}(aq) + 2H^+(aq)$
aqueous bromine		$SO_3^{2-}(aq) + Br_2(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2Br^-(aq) + 2H^+(aq)$

20.17 Aqueous sulphur dioxide as a reducing agent (p.93)

Practice 20.13

1 Sodium sulphite solution is added to acidified potassium permanganate solution until in excess.

a) State the expected observation;

Purple acidified potassium permanganate solution turns colourless.

b) Write the redox equation for the reaction.

$$5SO_3^{2-}(aq) + 2MnO_4^{-}(aq) + 6H^{+}(aq) \longrightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$$

2 Explain how sodium sulphite and sodium sulphate can be distinguished by using concentrated nitric acid.

Sodium sulphite is oxidised to sodium sulphate by concentrated nitric acid.

There is NO reaction between sodium sulphate and concentrated nitric acid.

Key terms (p.96)

oxidation	氧化作用	oxidation number	氧化數
reduction	還原作用	Stock system	司托克系統
redox reaction	氧化還原反應	breathalyser	呼氣分析儀
oxidising agent	氧化劑	disproportionation	歧化作用
reducing agent	還原劑		

Summary (p.97)

1 The following table summarises the definitions of oxidation and reduction.

Defined in terms of	Oxidation	Reduction
gain and loss of oxygen	gain of oxygen	loss of oxygen
loss and gain of hydrogen	loss of hydrogen	gain of hydrogen
loss and gain of electron(s)	loss of electron(s)	gain of electron(s)
changes in oxidation numbers	increase in oxidation number	decrease in oxidation number

2 Oxidation and reduction always takes place together. The combined process is called a redox reaction.

3 A metal high in the electrochemical series is a strong reducing agent while its ion is a weak oxidizing agent.

4 The oxidation number of an element is an imaginary charge assigned to it according to a set of rules.

Summary (p.97)

5 The following table summarises the characteristics of oxidising agents and reducing agents.

Oxidising agent	Reducing agent
causes oxidation	causes reduction
becomes reduced itself	becomes oxidised itself
loses oxygen	gains oxygen
gains hydrogen	loses hydrogen
gains electron(s)	loses electron(s)
oxidation number of an element in it decreases	oxidation number of an element in it increases

6 Redox equations are balanced by using:
a) half equations; and
b) changes in oxidation numbers.

Summary (p.97)

7 In a given electrochemical series,

a) the oxidising power of oxidising agents increases down the series;

b) the reducing power of reducing agents decreases down the series.

8 All oxidising agents should oxidise the reduced form of any chemical species above them in the electrochemical series.

Summary (p.97)

9 a) The following table summarises the oxidising properties of chlorine.

	Property	Reaction	Equation
Chlorine	oxidising property	aqueous chlorine oxidises bromide ions to bromine	$Cl_2(aq) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$
		aqueous chlorine oxidises iodide ions to iodine	$Cl_2(aq) + 2I^-(aq) \longrightarrow 2Cl^-(aq) + I_2(aq)$
	being oxidised and reduced simultaneously (disproportionation)	reaction with cold and dilute sodium hydroxide solution	$Cl_2(g) + 2NaOH(aq) \longrightarrow NaCl(aq) + NaOCl(aq) + H_2O(l)$
		reaction with hot and conc. sodium hydroxide solution	$3Cl_2(g) + 6NaOH(aq) \longrightarrow 5NaCl(aq) + NaClO_3(aq) + 3H_2O(l)$

Summary (p.97)

9 b) The following table summarises the reactions of sodium halides with concentrated sulphuric acid.

Reaction of conc. sulphuric acid with	Reactions that occur	Equations
sodium bromide (solid)	hydrogen bromide is first produced; hydrogen bromide reduces the sulphuric acid to sulphur dioxide	$NaBr(s) + H_2SO_4(l) \longrightarrow NaHSO_4(s) + HBr(g)$ $2HBr(s) + H_2SO_4(l) \longrightarrow Br_2(g) + SO_2(g) + 2H_2O(l)$
sodium iodide (solid)	hydrogen iodide is first produced; hydrogen iodide reduces the sulphuric acid to sulphur dioxide, sulphur and hydrogen sulphide	$NaI(s) + H_2SO_4(l) \longrightarrow NaHSO_4(s) + HI(g)$ $2HI(g) + H_2SO_4(l) \longrightarrow I_2(s) + SO_2(g) + 2H_2O(l)$ $6HI(g) + H_2SO_4(l) \longrightarrow 3I_2(s) + S(s) + 4H_2O(l)$ $8HI(g) + H_2SO_4(l) \longrightarrow 4I_2(s) + H_2S(g) + 4H_2O(l)$

Summary (p.97)

9 c) The following table summarises the reducing properties of aqueous sulphur dioxide.

	Reaction	Equation
Aqueous sulphur dioxide	reduces permanganate ions to manganese(II) ions	$5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \longrightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$
	reduces dichromate ions to chromium(III) ions	$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \longrightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$
	reduces iron(III) ions to iron(II) ions	$SO_3^{2-}(aq) + 2Fe^{3+}(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2Fe^{2+}(aq) + 2H^+(aq)$
	reduces bromine to bromide ions	$SO_3^{2-}(aq) + Br_2(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2Br^-(aq) + 2H^+(aq)$

Unit Exercise (p.99)

Note: Questions are rated according to ascending level of difficulty (from 1 to 5):

 **question targeted at level 3 and above;**

 **question targeted at level 4 and above;**

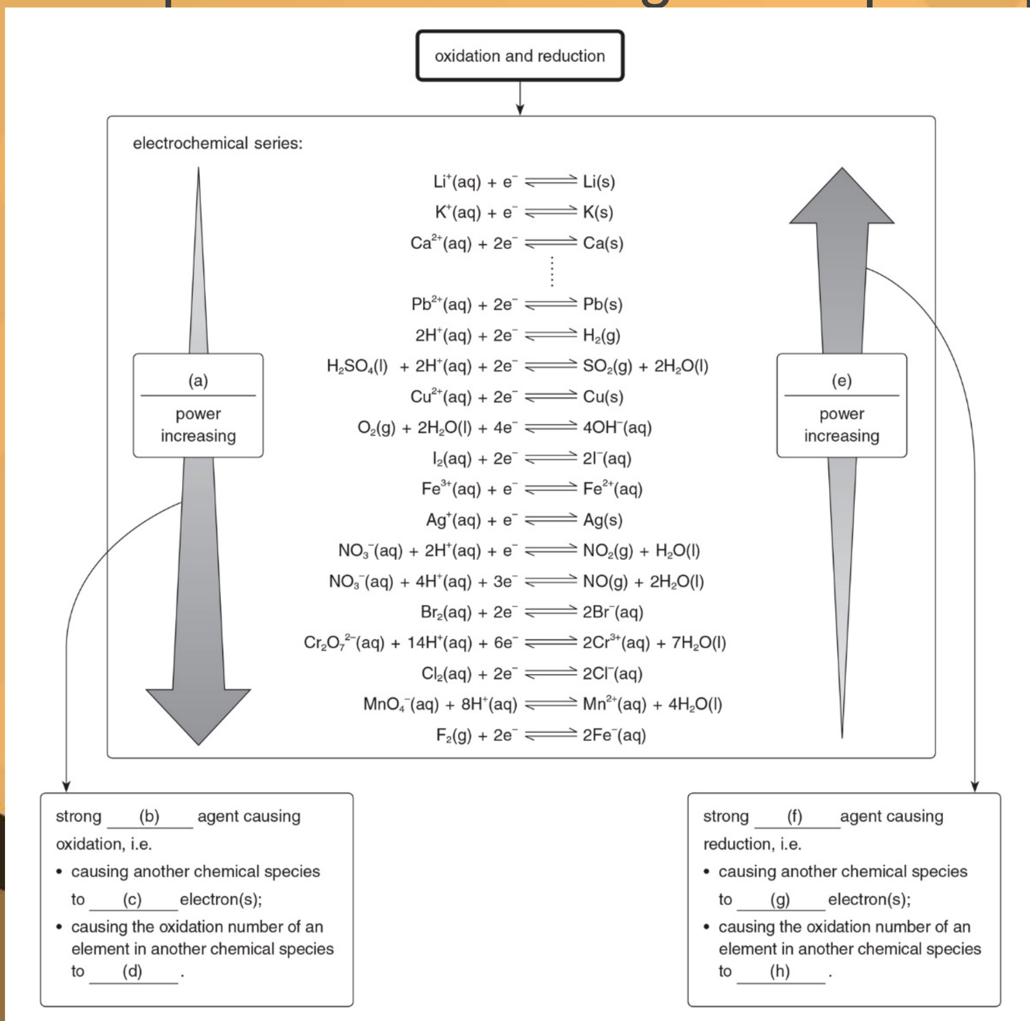
 **question targeted at level 5.**

‘ * ’ indicates 1 mark is given for effective communication.

Unit Exercise (p.99)

PART I KNOWLEDGE AND UNDERSTANDING

1 Complete the following concept map.



a) oxidising
b) Oxidising
c) lose
d) Increase
e) Reducing
f) Reducing
g) gain
h) decrease

Unit Exercise (p.99)

PART II MULTIPLE CHOICE QUESTIONS

2 Which of the following chemical species contains molybdenum (Mo) with its HIGHEST oxidation number?

A $MoCl_5$
B MoO_4^{2-}
C Mo_2S_3
D Mo_6Cl_{12}

Explanation:

Option	Chemical species	Oxidation number of Mo
A	$MoCl_5$	+5
B	MoO_4^{2-}	+6
C	Mo_2S_3	+3
D	Mo_6Cl_{12}	+2

Answer: B

Unit Exercise (p.99)

3 Which of the following statements concerning zinc is correct?

A It forms a soluble oxide when placed in NH_3(aq).
B It acts as a reducing agent when placed in HCl(aq).
C It undergoes oxidation when placed in $MgCl_2$(aq).
D It forms an acidic solution when placed in hot H_2O(l).

(HKDSE, Paper 1A, 2017, 11)

Answer: B

Unit Exercise (p.99)

4 The reaction between iron(III) ion and tin(II) ion may be represented by the ionic equation below.

$$2Fe^{3+}(aq) + Sn^{2+}(aq) \longrightarrow 2Fe^{2+}(aq) + Sn^{4+}(aq)$$

This reaction is classified as a redox reaction because

A the iron(III) ion is oxidised and the tin(II) ion acts as an oxidising agent.

B the iron(III) ion is oxidised and the tin(II) ion acts as a reducing agent.

C the iron(III) ion is reduced and the tin(II) ion acts as a reducing agent.

D the iron(III) ion is reduced and the tin(II) ion acts as an oxidising agent.

Answer: C

Unit Exercise (p.99)

5 Which of the following reactions of hydrochloric acid is a redox reaction?

A $2HCl(aq) + Ca(s) \longrightarrow H_2(g) + CaCl_2(aq)$
B $2HCl(aq) + K_2S(s) \longrightarrow H_2S(g) + 2KCl(aq)$
C $2HCl(aq) + CaO(s) \longrightarrow CaCl_2(aq) + H_2O(l)$
D $2HCl(aq) + K_2CO_3(s) \longrightarrow CO_2(g) + H_2O(l) + 2KCl(aq)$

Explanation:
In the reaction between hydrochloric acid and calcium, the oxidation number of hydrogen decreases from +1 to 0 while that of calcium increases from 0 to +2.

Answer: A

Unit Exercise (p.99)

6 Which of the following is NOT a redox reaction?

A $2AgBr(s) \longrightarrow 2Ag(s) + Br_2(g)$
B $SO_2(g) + 2H_2S(g) \longrightarrow 3S(s) + 2H_2O(l)$
C $2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$
D $Ca(HCO_3)_2(aq) \longrightarrow CaCO_3(s) + H_2O(l) + CO_2(g)$

Answer: D

Unit Exercise (p.99)

7 How many electrons are required when the following half equation is balanced using the smallest possible integers?

$$BrO_3^- + H^+ + e^- \longrightarrow Br_2 + H_2O$$

A 2
B 5
C 10
D 12

Explanation:

1	Balance the numbers of all atoms except hydrogen and oxygen.	$2BrO_3^- \longrightarrow Br_2$
2	Add H_2O to the side deficient in oxygen to balance the oxygen atoms.	$2BrO_3^- \longrightarrow Br_2 + 6H_2O$
3	Add H^+ to the side deficient in hydrogen to balance the hydrogen atoms.	$2BrO_3^- + 12H^+ \longrightarrow Br_2 + 6H_2O$
4	Add e^- to the side deficient in negative charge to balance the charge.	The half equation is: $2BrO_3^- + 12H^+ + xe^- \longrightarrow Br_2 + 6H_2O$ The value of x can be worked out in either of two ways: • The charge on the right-hand side is 0, so the charge on the left-hand side is also 0. $2 \times (-1) + 12 \times (+1) - x = 0$, $x = 10$ • The oxidation number of Br in BrO_3^- is +5 and it is 0 in Br_2. The oxidation number of each Br atom changes by 5. As there are two Br atoms, the total change is 10. This means that there must be ten electrons on the left-hand side.

Answer: C

Unit Exercise (p.99)

8 Consider the following chemical equation:

$$3Ni(OH)_2(s) + xHCl(aq) + yAuCl_4^-(aq) \longrightarrow 3NiCl_4(aq) + yAu(s) + zCl^-(aq) + 6H_2O(l)$$

Which of the following combinations is correct?

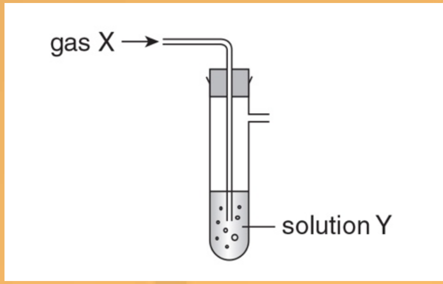
	x	y	z
A	4	2	2
B	6	2	2
C	4	3	3
D	6	3	3

Answer: B

(HKDSE, Paper 1A, 2017, 15)

Unit Exercise (p.99)

9 Gas X is bubbled steadily into solution Y as shown in the diagram below:



In which of the following combinations would NOT have a visible change in solution Y?

	gas X	solution Y
A	$Cl_2(g)$	$KI(aq)$
B	$O_2(g)$	$FeSO_4(aq)$
C	$CO_2(g)$	acidified $KMnO_4(aq)$
D	$SO_2(g)$	acidified $Na2Cr_2O_7(aq)$

Answer: C

(HKDSE, Paper 1A, 2016, 13)

Unit Exercise (p.99)

10 Astatine (At) is below iodine in Group VII of the Periodic Table.

Which of the following statements is most likely to be correct?

A NaAt(aq) reacts with $AgNO_3$(aq) to form a yellow solution.
B NaAt(s) and concentrated sulphuric acid react to form astatine.
C Astatine and KCl(aq) react to form KAt(aq) and chlorine.
D NaAt(aq) and dilute sulphuric acid react to give steamy fumes of HAt(g).

Answer: B

Unit Exercise (p.99)

10 (continued)

Explanation:

Option A — NaCl(aq), NaBr(aq) and NaI(aq) react with $AgNO_3$(aq) to form precipitates. Hence NaAt(aq) should react with $AgNO_3$(aq) to form a precipitate.

Option B — NaBr(s) and NaI(s) react with concentrated sulphuric acid to form bromine and iodine respectively. Hence it is likely that NaAt(s) reacts with concentrated sulphuric acid to form astatine.

Option C — Astatine should be a weaker oxidising agent than chlorine. Hence there is NO reaction between astatine and KCl(aq).

Option D — NaCl(aq), NaBr(aq) and NaI(aq) do not react with dilute sulphuric acid. Hence there should be NO reaction between NaAt(aq) and dilute sulphuric acid.

Unit Exercise (p.99)

11 A solid gives a brick-red flame in a flame test and reacts with concentrated sulphuric acid to produce steamy fumes, but no other gases. The solid could be

A calcium chloride.
B calcium bromide.
C sodium chloride.
D sodium bromide.

Explanation:
Calcium chloride reacts with concentrated sulphuric acid to produce steamy fumes (hydrogen chloride gas).

Answer: A

Unit Exercise (p.99)

12 Part of the electrochemical series is shown below.

higher in electrochemical series

$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$

$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$

$Br_2(aq) + 2e^- \rightleftharpoons 2Br^-(aq)$

$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$

lower in electrochemical series

Which of the following oxidising agents CANNOT oxidise $Fe^{2+}(aq)$ ion to $Fe^{3+}(aq)$ ion?

(1) $Zn^{2+}(aq)$
(2) $Br_2(aq)$
(3) $MnO_4^-(aq)$

A (1) only
B (2) only
C (1) and (3) only
D (2) and (3) only

Explanation:

(2) and (3) $Br_2(aq)$ and $MnO_4^-(aq)$ are stronger oxidising agents than $Fe^{3+}(aq)$. Hence they can oxidise $Fe^{2+}(aq)$ ion to $Fe^{3+}(aq)$ ion.

Answer: A

Unit Exercise (p.99)

13 Cl_2(aq) and $S_2O_3^{2-}$(aq) ion react according to the following unbalanced equation:

$Cl_2(aq) + S_2O_3^{2-}(aq) + HO(l) \rightarrow SO_4^{2-}(aq) + H^+(aq) + Cl^-(aq)$

Which of the following statements is / are correct?

(1) The oxidation number of sulphur in $S_2O_3^{2-}$(aq) ion is +3.

(2) One of the half equations of this reaction is

$S_2O_3^{2-}(aq) + 5H_2O(l) \longrightarrow 2SO_4^{2-}(aq) + 10H^+(aq) + 8e^-$

(3) Cl_2(aq) is reduced by $S_2O_3^{2-}$(aq) ion in the reaction.

A (1) only

B (2) only

C (1) and (3) only

D (2) and (3) only

Explanation:

(1) The oxidation number of sulphur in $S_2O_3^{2-}$(aq) ion is +2.

(3) The oxidation number of chlorine decreases from 0 to –1. Thus, Cl_2(g) is reduced in the reaction.

Answer: D

Unit Exercise (p.99)

14 Which of the following statements concerning the reaction of hot concentrated sulphuric acid with copper are correct?

(1) A colourless gas is evolved.
(2) A blue solution is formed.
(3) One mole of $H_2SO_4(l)$ requires two moles of electrons for reduction.

A (1) and (2) only
B (1) and (3) only
C (2) and (3) only
D (1), (2) and (3)

Explanation:
(1) Sulphur dioxide gas is produced in the reaction between copper and concentrated sulphuric acid.
(2) Concentrated sulphuric acid oxidises copper to form copper(II) ions. A blue solution results.
(3) The half equation for the reduction of $H_2SO_4(l)$ is shown below.
$H_2SO_4(l) + 2H^+(aq) + 2e^- \rightleftharpoons SO_2(g) + 2H_2O(l)$
Thus, one mole of $H_2SO_4(l)$ requires two moles of electrons for reduction.

Answer: D

Unit Exercise (p.99)

15 What would be observed when a few drops of concentrated nitric acid is added to KI(aq)?

(1) A brown solution is formed.
(2) A brown precipitate is formed.
(3) A reddish brown gas is released.

A (1) and (2) only
B (1) and (3) only
C (2) and (3) only
D (1), (2) and (3)

Answer: B

(HKDSE, Paper 1A, 2017, 23)

Unit Exercise (p.99)

16 Which of the following reagents can convert iron(II) ions to iron(III) ions?

(1) Aqueous chlorine
(2) Dilute nitric acid
(3) Dilute sulphuric acid

A (1) and (2) only
B (1) and (3) only
C (2) and (3) only
D (1), (2) and (3)

Explanation:
Aqueous chlorine and dilute nitric acid are oxidising agents.

Answer: A

(HKDSE, Paper 1A, 2017, 23)

Unit Exercise (p.99)

PART III STRUCTURED QUESTIONS

17 Which of the following best describes each of the reactions represented by the equations below?

oxidation reduction redox not oxidation or reduction

a) $2Br^-(aq) \longrightarrow Br_2(aq) + 2e^-$

oxidation (1)

b) $Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$

not oxidation or reduction (1)

c) $Cu^{2+}(aq) + e^- \longrightarrow Cu^+(aq)$

reduction (1)

d) $K_2SO_3(aq) + H_2O_2(aq) \longrightarrow K_2SO_4(aq) + H_2O(l)$

redox (1)

e) $Na_2CO_3(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$

not oxidation or reduction (1)

Unit Exercise (p.99)

18 A redox reaction occurs when tin(II) ion reacts with iodate ion.

$$3Sn^{2+}(aq) + IO_3^-(aq) + 6H^+(aq) \longrightarrow 3Sn^{4+}(aq) + I^-(aq) + 3H_2O(l)$$

Indicate each of the following for the reaction:

a) the substance being oxidised;
$Sn^{2+}(aq)$ (1)

b) the substance being reduced;
$IO_3^-(aq)$ (1)

c) the oxidising agent;
$IO_3^-(aq)$ (1)

d) the reducing agent.
$Sn^{2+}(aq)$ (1)

Unit Exercise (p.99)

19 Find the oxidation numbers of the underlined elements in the following chemical species.

a) $H_2\underline{S}$	–2 (1)
b) $\underline{O}F_2$	+2 (1)
c) $Ca\underline{C}_2$	–1 (1)
d) $\underline{Mn}O_4^{2-}$	+6 (1)
e) $\underline{N}H_2OH$	–1 (1)
f) $\underline{V}O^{2+}$	+4 (1)
g) $\underline{N}_2O$	+1 (1)
h) $\underline{Cr}(H_2O)_4Br_2^+$	+3 (1)

Unit Exercise (p.99)

20 a) What is mean by term 'oxidation'?

Any one of the following:

- Gain of oxygen (1)
- Loss of hydrogen (1)
- Loss of electrons (1)
- Increase in oxidation number (1)

b) State, with a reason, which chemical species is oxidised in the reaction below.

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

Zn(s) is oxidised because of loss of electrons / increase in oxidation number. (1)

c) State, with a reason, which chemical species is reduced in the reaction below.

$Hg^{2+}(aq) + Sn^{2+}(aq) \longrightarrow Hg(s) + Sn^{4+}(aq)$

$Hg^{2+}(aq)$ is reduced because of gain of electrons / decrease in oxidation number. (1)

Unit Exercise (p.99)

21 The smelting of iron occurs in a blast furnace, as represented by the following equation:

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$$

In terms of oxidation numbers, explain whether the reaction involves oxidation and reduction.

Yes

Oxidation: <u>Oxidation number of C increases</u> from +2 to +4. (1)

Reduction: <u>Oxidation number of Fe decreases</u> from +3 to 0. (1)

Unit Exercise (p.99)

22 Consider each of the experiments below.

i) State the expected observation.

ii) Write the ionic equation for the reaction.

a) Adding aqueous chlorine to potassium bromide solution

A yellow-brown colour appears. (1)

$Cl_2(aq) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$ (1)

b) Adding concentrated nitric acid to copper

Any one of the following:

- The copper dissolves. (1)
- A brown gas is given off. (1)
- A blue solution forms. (1)
- $Cu(s) + 2NO_3^-(aq) + 4H^+(aq) \longrightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$ (1)

c) Adding sodium sulphite solution to aqueous bromine

The yellow-brown aqueous bromine turns colourless. (1)

$SO_3^{2-}(aq) + Br_2(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2Br^-(aq) + 2H^+(aq)$ (1)

Unit Exercise (p.99)

23 For the following experiment, state the expected observation, and write the chemical equation(s) for the reaction(s) involved.

Adding sodium sulphite solution to acidified potassium dichromate solution until in excess

(HKDSE, Paper 1B, 2015, 2(b))

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Unit Exercise (p.99)

24 An acidified solution containing orange $Cr_2O_7^{2-}(aq)$ ions reacts with zinc in a redox reaction to form a solution containing $Zn^{2+}(aq)$ ions and blue $Cr^{2+}(aq)$ ions.

a) What is the oxidation number of chromium in the $Cr_2O_7^{2-}(aq)$ ion?

+6 (1)

b) The unbalanced half equation for the change involving $Cr_2O_7^{2-}(aq)$ ions is shown below.

$Cr_2O_7^{2-}(aq) + H^+(aq) + e^- \longrightarrow Cr^{2+}(aq) + H_2O$

Balance this half equation.

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 8e^- \longrightarrow 2Cr^{2+}(aq) + 7H_2O(l)$ (1)

c) Deduce the ionic equation for the reaction.

$Zn(s) + Cr_2O_7^{2-}(aq) + 14H^+(aq)$
$\longrightarrow 4Zn^{2+}(aq) + 2Cr^{2+}(aq) + 7H_2O(l)$ (1)

Unit Exercise (p.99)

25 H_2O_2(aq) can be used in the restoration of old paintings. It makes them lighter by converting black lead(II) sulphide in the paints into white lead(II) sulphate. The unbalanced equation for this reaction is

$$PbS(s) + H_2O_2(aq) \longrightarrow PbSO_4(s) + H_2O(l)$$

a) In terms of oxidation numbers, explain whether the reaction involves oxidation and reduction.

Yes

Oxidation: Oxidation number of S increases from –2 to +6. (1)

Reduction: Oxidation number of O decreases from –1 to –2. (1)

b) Balance the equation for the reaction.

$$PbS(s) + 4H_2O_2(aq) \longrightarrow PbSO_4(s) + 4H_2O(l) \quad (1)$$

Unit Exercise (p.99)

26 When sodium hydroxide solution is added to potassium dichromate solution, the following change takes place.

$$Cr_2O_7^{2-}(aq) + 2OH^-(aq) \longrightarrow 2CrO_4^{2-}(aq) + H_2O(l)$$

Decide whether this is a redox reaction. Explain your answer in terms of oxidation numbers.

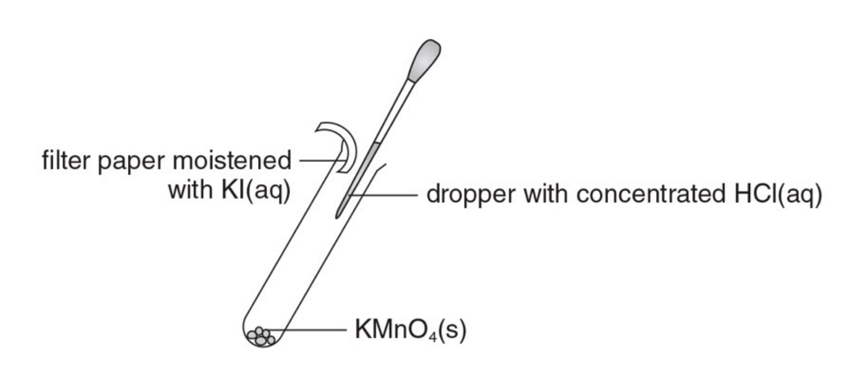
$$Cr_2O_7^{2-}(aq) + 2OH^-(aq) \longrightarrow 2CrO_4^{2-}(aq) + H_2O(l)$$

	$Cr_2O_7^{2-}$	OH^-	CrO_4^{2-}	H_2O
Cr	+6		+6	
O	−2	−2	−2	−2
H		+1		+1

The oxidation number of all the elements remain unchanged. Hence this is NOT a redox reaction.

Unit Exercise (p.99)

27 Refer to the experimental set-up as shown below.



a) HCl is a strong acid. What is meant by the term ‘strong acid’?

b) When concentrated HCl(aq) is dropped into $KMnO_4$(s), a yellowish green gas is formed.

i) What is the yellowish green gas?

ii) Explain whether the reaction forming the yellowish green gas is a redox reaction.

Unit Exercise (p.99)

27 (continued)

c) With the aid of an ionic equation, state the expected observation when the yellowish green gas reaches the filter paper.

d) In consideration of laboratory safety, explain where the experiment should be performed.

(HKDSE, Paper 1B, 2018, 8)

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Unit Exercise (p.99)

28 Chlorine forms a series of oxides, some of which are listed below.

dichlorine monoxide Cl_2O chlorine dioxide ClO_2

dichlorine hexoxide Cl_2O_6

a) Deduce the systematic name of chlorine dioxide using the oxidation number of chlorine.

Chlorine(IV) oxide (1)

b) Chlorine dioxide dissolves in water to form a solution which eventually forms a mixture of two acids.

$$6ClO_2(g) + 3H_2O(l) \longrightarrow 5HClO_3(aq) + HCl(aq)$$

By giving appropriate oxidation numbers, explain why this is a disproportionation reaction.

$$6ClO_2(g) + 3H_2O(l) \longrightarrow 5HClO_3(aq) + HCl(aq)$$

oxidation number of Cl +4 +5 –1 (1)

Chlorine is both oxidised and reduced. Thus, this is a disproportionation reaction. (1)

Unit Exercise (p.99)

29 Solutions of potassium chloride and potassium iodide can be distinguished by using aqueous bromine.

a) State what would be observed when aqueous bromine is added separately to solutions of the halides.

There is no observable change when aqueous bromine is added to potassium chloride solution.
A brown solution results when aqueous bromine is added to potassium iodide solution. (1)

b) Balance the equation for the reaction.

$Br_2(aq) + 2I^-(aq) \longrightarrow 2Br^-(aq) + I_2(aq)$ (1)

c) What do the reactions with bromine show about the relative oxidising abilities of chlorine, bromine and iodine?

The oxidising ability of bromine is lower than that of chlorine but higher than that of iodine. (1)

Unit Exercise (p.99)

30 Nitrogen monoxide (NO) is formed when nitrate ions are reduced by silver in acidic solution. The overall equation for the reaction is shown below.

$$3Ag(s) + NO_3^-(aq) + 4H^+(aq) \longrightarrow 3Ag^+(aq) + NO(g) + 2H_2O(l)$$

a) Give the oxidation numbers of silver and nitrogen in both the reactants and products.
silver: from __0__ to __+1__. (1)
nitrogen: from __+5__ to __+2__. (1)

b) Give the two half equations that can be combined to give the ionic equation shown above.
$Ag(s) \longrightarrow Ag^+(aq) + e^-$ (1)
$NO_3^-(aq) + 4H^+(aq) + 3e^- \longrightarrow NO(g) + 2H_2O(l)$ (1)

Unit Exercise (p.99)

31 In the laboratory, hydrogen chloride can be prepared by reacting sodium chloride with concentrated sulphuric acid.

Some students tried to prepare hydrogen iodide by reacting sodium iodide with concentrated sulphuric acid. When concentrated sulphuric acid was added to sodium iodide, solid sulphur and a black solid were formed. Hydrogen sulphide gas was also produced.

a) Identify the black solid.
Iodine (1)

Unit Exercise (p.99)

31 (continued)

b) Consider the formation of hydrogen sulphide from concentrated sulphuric acid.

State the change in the oxidation number of sulphur.

From +6 to –2 (1)

c) Hydrogen iodide is NOT obtained from the reaction between sodium iodide and concentrated sulphuric acid. Explain why.

HI is a reducing agent which reduces H_2SO_4. / H_2SO_4 is an oxidising agent which oxidises HI. (1)

Unit Exercise (p.99)

32 Sulphur is burnt in a gas jar of oxygen. The gas formed is sulphur dioxide.

a) The gas is tested with a piece of filter paper soaked in acidified potassium dichromate solution.

i) State the colour change observed.

From orange to green (1)

ii) Write the ionic equation for the reaction involved.

$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \longrightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$ (1)

b) Sulphur dioxide gas is passed into a solution containing iron(III) ions.

i) State the colour change observed.

From yellow-brown to pale green (1)

ii) Write the ionic equation for the reaction involved.

$SO_3^{2-}(aq) + 2Fe^{3+}(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2Fe^{2+}(aq) + 2H^+(aq)$ (1)