Electrode Potentials & Redox
A Level only

Question Paper 1

Level	A Level
Subject	Chemistry
Exam Board	OCR
Module	Physical Chemistry & Transition Elements
Topic	Electrode Potentials & Redox
Paper	A Level only
Booklet	Question Paper 1

Time allowed: **47 minutes**

Score: **/35**

Percentage: **/100**

Grade Boundaries:

A*	A	B	C	D	E
>85%	73%	60%	47%	34%	21%

Question 1

The redox equilibria for a hydrogen–oxygen fuel cell in alkaline solution are shown below.

$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$ $\quad E^\ominus = -0.83\,V$

$^1/_2O_2(g) + H_2O(l) + 2e^- \rightleftharpoons 2OH^-(aq)$ $\quad E^\ominus = +0.40\,V$

What is the equation for the overall cell reaction?

A $H_2(g) + 4OH^-(aq) \rightarrow 3H_2O(l) + {}^1/_2O_2(g)$

B $3H_2O(l) + {}^1/_2O_2 \rightarrow H_2(g) + 4OH^-(aq)$

C $H_2O(l) \rightarrow H_2(g) + {}^1/_2O_2(g)$

D $H_2(g) + {}^1/_2O_2(g)$ $\quad H_2O(l)$ **[1]**

Chromite is the main ore of chromium. The chromium-containing compound in chromite is $Fe(CrO_2)_2$.
The percentage of chromium in a sample of chromite can be determined using the method below.

Step 1

A 5.25 g sample of chromite ore is heated with sodium peroxide, Na_2O_2.

$$2Fe(CrO_2)_2 + 7Na_2O_2 \rightarrow 2NaFeO_2 + 4Na_2CrO_4 + 2Na_2O$$

Water is added to the resulting mixture.
Na_2CrO_4 dissolves in the water forming a solution containing CrO_4^{2-} ions.

Step 2

The mixture from **Step 1** is filtered and the filtrate is made up to 1.00 dm^3 in a volumetric flask.
A 25.0 cm^3 sample of this alkaline solution is pipetted into a conical flask and an excess of aqueous potassium iodide is added.

- A redox reaction takes place between I^- ions, CrO_4^{2-} ions and H_2O.
- In this reaction 1 mol CrO_4^{2-} forms 1.5mol I_2.

Step 3

The resulting mixture is titrated with 0.100 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3(aq)$ to estimate the I_2 present:

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

The average titre of $Na_2S_2O_3(aq)$ is 25.5 cm^3.

(a) In **Step 1** Na_2O and $NaFeO_2$ react with water forming an alkaline solution containing a brown precipitate. This is **not** a redox reaction.

Write equations for:
- the reaction of Na_2O with water
- the reaction of $NaFeO_2$ with water. [2]

(b) Determine the percentage, by mass, of chromium in the ore.
Give your answer to **one** decimal place. [6]

(c) This part refers to **Step 2** of the method.

In the redox reaction between I^- ions, CrO_4^{2-} ions and H_2O:

- CrO_4^{2-} ions, are reduced to chromium(III) ions, Cr^{3+}
- I^- ions are oxidised to iodine, I_2
- Construct an overall equation for the redox reaction and write half equations for the oxidation and reduction.

Overall equation:

Half equations: [3]

[Total: 11 Marks]

Electrochemical cells contain two redox systems, one providing electrons and the other accepting electrons. The tendency to lose or gain electrons is measured using values called standard electrode potentials.

(a) Define the term *standard electrode potential*.

Include all standard conditions in your answer. [2]

(b) The table below shows two redox systems and their standard electrode potentials, $E^{\ominus}$.

Redox system	$E^{\ominus}$/V
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Ag^{+}(aq) + e^- \rightleftharpoons Ag(s)$	+0.80

A standard $Cu^{2+}(aq)$ / Cu(s) half-cell is connected to a standard $Ag^{+}(aq)$ / Ag(s) half-cell. The potential of the cell is measured.

Water is then added to the $Cu^{2+}(aq)$ / Cu(s) half-cell. This changes the position of equilibrium in the half-cell. The cell potential increases.

(i) Write down the equation for the overall cell reaction. [1]

(ii) Explain, in terms of equilibrium, why the cell potential increases. [3]

(c) Hydrogen fuel cells are being developed for powering vehicles.

(i) State **one** advantage of using hydrogen as a fuel compared with conventional fuels. [1]

(i) In vehicles, hydrogen can be stored on the surface of a solid material or within a solid material.

State **one** other way that hydrogen can be stored as a fuel for vehicles. [1]

(d) Aluminium–oxygen cells are being investigated for powering vehicles.

The **reactions** at each electrode are shown below.

$$Al\,(s) + 4OH^-(aq) \rightarrow Al(OH)_4^-(aq) + 3e^-$$

$$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$

(i) The standard electrode potential for the O_2 / OH^- redox system is +0.40V.
The standard cell potential of an aluminium–oxygen cell is 2.71V.

What is the standard electrode potential of the aluminium redox system in this cell? [1]

(ii) Construct the overall cell equation for an aluminium–oxygen cell. [2]

[Total: 11 Marks]

A student carries out an investigation to prepare and analyse a sample of barium ferrate(VI), $BaFeO_4$. The steps in the investigation are shown below.

Step 1

The student adds solid iron(III) oxide to a hot aqueous solution containing an excess of hydroxide ions. The student bubbles chlorine gas through the mixture.
A solution forms containing aqueous ferrate(VI) ions, $FeO_4^{2-}(aq)$, and aqueous chloride ions.

Step 2

The student adds aqueous barium chloride to the resulting solution.
A precipitate of impure barium ferrate(VI) forms.
The precipitate is filtered, washed with distilled water and dried.
The student obtains 0.437 g of impure solid barium ferrate(VI).

Step 3

An excess of acidified aqueous potassium iodide is added to the solid from **step 2**.
The $BaFeO_4$ reacts as shown below, and the impurity does not react. A solution forms containing aqueous iodine, $I_2(aq)$.

$$BaFeO_4(s) + 8H^+(aq) + 3I^-(aq) \rightarrow Fe^{3+}(aq) + Ba^{2+}(aq) + 1\tfrac{1}{2}I_2(aq) + 4H_2O(l)$$

Step 4

The student determines the amount of I_2 formed by carrying out a titration with aqueous sodium thiosulfate, $Na_2S_2O_3(aq)$.

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

26.4 cm^3 of 0.100 $mol\,dm^{-3}$ $Na_2S_2O_3(aq)$ are required to reach the end point.

(a) Construct an equation for the oxidation of iron(III) oxide (**step 1**).

[2]

(b) Write an **ionic** equation for the formation of barium ferrate(VI) (**step 2**).

Include state symbols. [1]

(c) In **step 3**, what is the reducing agent?

Explain your answer in terms of electrons.

reducing agent [2]

(d) The solid sample of barium ferrate(VI) obtained in step 2 is impure.

Determine the percentage, by mass, of barium ferrate(VI) in the 0.437 g of solid formed in **step 2**.

Give your answer to **one** decimal place. [4]

(e) When the solution is not alkaline, ferrate(VI) ions react with water.
The reaction forms a gas with a density of 1.333×10^{-3} g cm^{-3}, measured at room temperature and pressure, and an orange–brown precipitate.

- Determine the formulae of the gas and the precipitate.
- Write an equation for the reaction that takes place.

Gas .

Precipitate

equation [3]

[Total: 12 Marks]