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**CHEMISTRY**

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**Topic Questions**

**Module 5: Physical chemistry and transition elements**

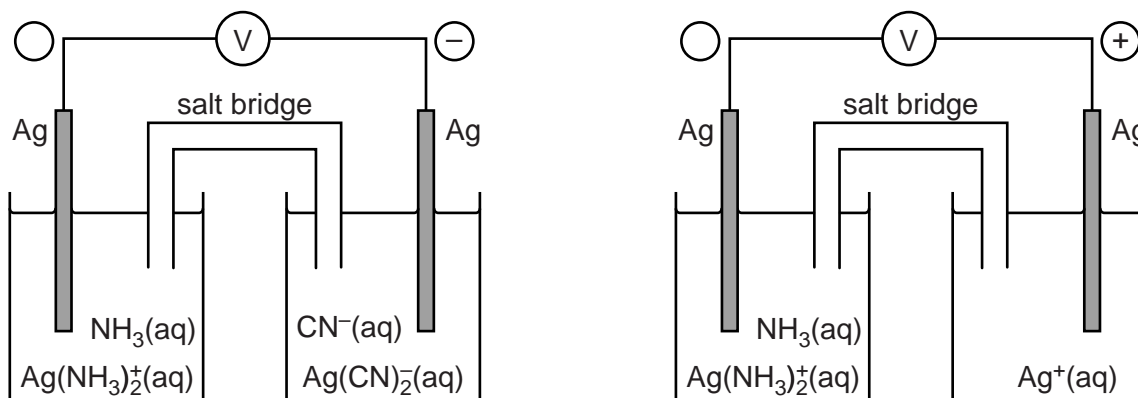
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1 Three redox systems, **C**, **D** and **E** are shown in Table 6.1.

<b>C</b>	$\text{Ag}(\text{NH}_3)_2^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s}) + 2\text{NH}_3(\text{aq})$
<b>D</b>	$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$
<b>E</b>	$\text{Ag}(\text{CN})_2^-(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s}) + 2\text{CN}^-(\text{aq})$

**Table 6.1**

The two cells below were set up in an experiment to compare the standard electrode potentials of redox systems **C**, **D** and **E**. The signs on each electrode are shown.



(a) List the three redox systems in order by adding the labels **C**, **D** and **E** to the table below.

$E^\ominus$	redox system
<b>Most negative</b>	
↑	
<b>Least negative</b>	

[1]



- (b) A standard cell is set up between redox system **D** in **Table 6.1** and a standard hydrogen half-cell. The standard cell potential of redox system **D** is +0.34V.

The cell delivers a current for a length of time.  
 The pH of the solution in the standard hydrogen half-cell decreases.

- (i) What is the pH of the solution in a standard hydrogen half-cell?

pH = ..... [1]

- (ii) Explain, in terms of electrode potentials and equilibrium, why the pH of the solution in the hydrogen half-cell decreases as this cell delivers current.

.....  
 .....  
 .....  
 .....  
 ..... [2]

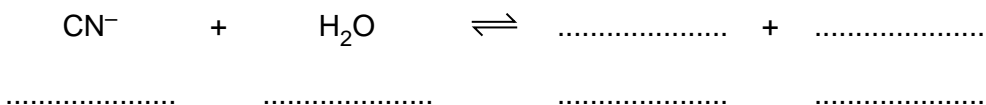
- (iii) Write the equation for the overall cell reaction that takes place in this cell.

..... [1]

- (c) The CN<sup>-</sup> ion is the conjugate base of a very toxic weak acid.

In aqueous solutions of CN<sup>-</sup> ions, an acid–base equilibrium is set up.

- (i) Complete the equation for this equilibrium and label the conjugate acid–base pairs.



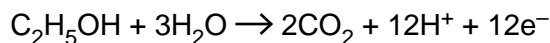
[1]

- (ii) Explain, in terms of equilibrium, why acidic conditions should **not** be used with cells containing CN<sup>-</sup>(aq) ions.

.....  
 .....  
 ..... [1]

- (d) Direct-ethanol fuel cells (DEFCs) are being developed in which the fuel is ethanol rather than hydrogen.

The half-equation for the reaction at the ethanol electrode of the DEFC is shown below:



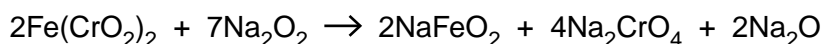
- (i) State **one** important difference between a fuel cell and a modern storage cell.
- .....
- ..... [1]
- (ii) Suggest **one** advantage of using ethanol, rather than hydrogen, in a fuel cell for vehicles.
- .....
- ..... [1]
- (iii) The overall reaction in a DEFC is the same as for the complete combustion of ethanol.
- Write the equation for the overall reaction in a DEFC.
- ..... [1]
- (iv) Deduce the half-equation for the reaction at the oxygen electrode in a DEFC.
- ..... [1]
- (v) Using oxidation numbers, show that oxidation and reduction take place in a DEFC.
- Oxidation: .....
- .....
- Reduction: .....
- ..... [2]

[Total: 13]

- 2 Chromite is the main ore of chromium. The chromium-containing compound in chromite is  $\text{Fe}(\text{CrO}_2)_2$ . 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,  $\text{Na}_2\text{O}_2$ .



Water is added to the resulting mixture.

$\text{Na}_2\text{CrO}_4$  dissolves in the water forming a solution containing  $\text{CrO}_4^{2-}$  ions.

**Step 2**

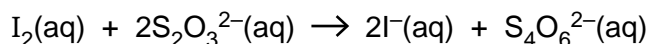
The mixture from **Step 1** is filtered and the filtrate is made up to  $1.00 \text{ dm}^3$  in a volumetric flask.

A  $25.0 \text{ cm}^3$  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  $\text{I}^-$  ions,  $\text{CrO}_4^{2-}$  ions and  $\text{H}_2\text{O}$ .
- In this reaction 1 mol  $\text{CrO}_4^{2-}$  forms 1.5 mol  $\text{I}_2$ .

**Step 3**

The resulting mixture is titrated with  $0.100 \text{ mol dm}^{-3}$  sodium thiosulfate,  $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$  to estimate the  $\text{I}_2$  present:



The average titre of  $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$  is  $25.5 \text{ cm}^3$ .

- (a) In **Step 1**  $\text{Na}_2\text{O}$  and  $\text{NaFeO}_2$  react with water forming an alkaline solution containing a brown precipitate. This is **not** a redox reaction.

Write equations for:

- the reaction of  $\text{Na}_2\text{O}$  with water
- the reaction of  $\text{NaFeO}_2$  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  $\text{I}^-$  ions,  $\text{CrO}_4^{2-}$  ions and  $\text{H}_2\text{O}$ :

- $\text{CrO}_4^{2-}$  ions, are reduced to chromium(III) ions,  $\text{Cr}^{3+}$
- $\text{I}^-$  ions are oxidised to iodine,  $\text{I}_2$
- Construct an overall equation for the redox reaction and write half equations for the oxidation and reduction.

Overall equation:

Half equations:

[3]

[Total: 11]



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

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

Water is then added to the  $\text{Cu}^{2+}(\text{aq})/\text{Cu}(\text{s})$  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.

.....  
.....  
.....  
.....

(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]

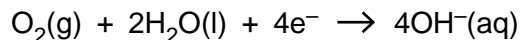
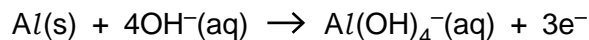
(ii) In vehicles, hydrogen can be stored on the surface of a solid material or within a solid material.

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.



(i) The standard electrode potential for the  $\text{O}_2/\text{OH}^{\text{-}}$  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?

standard electrode potential = ..... V [1]

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

[2]

[Total: 11]



- 4 A student carries out an investigation to prepare and analyse a sample of barium ferrate(VI),  $\text{BaFeO}_4$ . The steps in the investigation are shown below.

**Step 1**

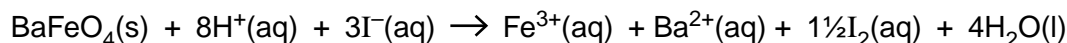
The student adds solid iron(III) oxide to a hot aqueous solution containing an excess of hydroxide ions. The student bubbles chlorine gas through the mixture.  
A solution forms containing aqueous ferrate(VI) ions,  $\text{FeO}_4^{2-}(\text{aq})$ , and aqueous chloride ions.

**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).

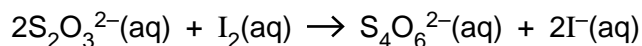
**Step 3**

An excess of acidified aqueous potassium iodide is added to the solid from **step 2**.  
The  $\text{BaFeO}_4$  reacts as shown below, and the impurity does not react. A solution forms containing aqueous iodine,  $\text{I}_2(\text{aq})$ .



**Step 4**

The student determines the amount of  $\text{I}_2$  formed by carrying out a titration with aqueous sodium thiosulfate,  $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$ .



26.4 cm<sup>3</sup> of 0.100 mol dm<sup>-3</sup>  $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$  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* .....

*explanation* .....

.....

.....

[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.

percentage of barium ferrate(VI) = ..... % [4]



(e) When the solution is not alkaline, ferrate(VI) ions react with water. The reaction forms a gas with a density of  $1.333 \times 10^{-3} \text{ g cm}^{-3}$ , 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]