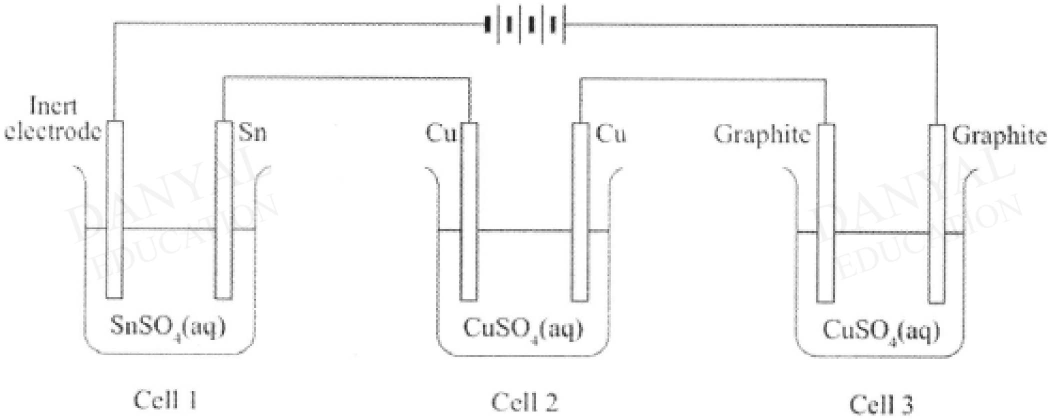
O Level Pure Chemistry Structured

Electrolysis Test 2.0

Q1

Three electrolytic cells are set up in series as shown below.
All of the electrolytes have a concentration of 1.00 mol/dm^3.



(a) Write the half ionic equations with state symbols in the table below to show the reactions that happen in Cells 2 and 3.

cell	electrode	half ionic equation
2	copper anode	
	copper cathode	
3	graphite anode	
	graphite cathode	

[4]

(b) Calculate the mass of copper produced at one of the electrodes in cell 2 if the tin electrode in cell 1 decreased in mass by 0.034 g.

[2]

(c) Describe the colour and the pH of the solutions in cells 2 and 3 after the electrolysis set-up has been running for some time Explain your answers.

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....

....

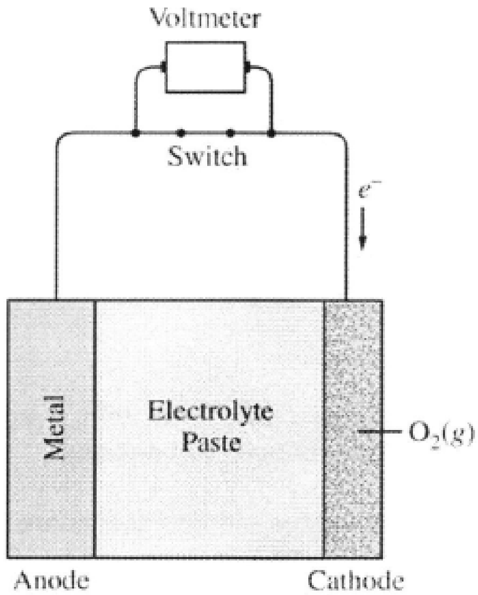
....

.... [4]

[Total: 10]

Q2

Metal-air cells are a relatively new type of portable energy source consisting of a metal anode, an alkaline electrolyte paste that contains water, and a porous cathode membrane that lets in oxygen from the air.



A schematic of the cell is shown above. The half ionic equations that take place at the cathode and at the three possible metal anodes are given in the table below.

electrode	half ionic equation
cathode	O_2 (g) + $2H_2O$ (l) + $4e^-$ → $4OH^-$ (aq)
zinc anode	Zn (s) + $2OH^-$ (aq) → ZnO (s) + H_2O (l) + $2e^-$
sodium anode	2Na (s) + $2OH^-$ (aq) → Na_2O (s) + H_2O (l) + $2e^-$
calcium anode	Ca (s) + $2OH^-$ (aq) → CaO (s) + H_2O (l) + $2e^-$

(a) Early forms of metal-air cells used zinc as the anode. Zinc oxide is produced as the cell operates.

Using the half ionic equations from the table above, write the overall equation that takes place in zinc-air cells.

.... [1]

(b) The electrolyte paste contains OH^- ions. On the diagram of the cell above, draw an arrow to indicate the direction of migration of OH^- ions through the electrolyte as the cell operates.

[1]

(c) Metal-air cells need to be lightweight for many applications. A fresh zinc-air cell is weighed on an analytical balance before being placed in a hearing aid for use.

(i) As the cell operates, does the mass of the cell increase, decrease, or remain the same? Using the overall equation in **(a)**, explain your answer.

....

....

....

....

.... [2]

(ii) In order to transfer more electrons with a smaller mass, sodium and calcium metals are investigated as potential anodes.

Which of these 2 metals would transfer more electrons if 1 g of metal anode is used? Justify your answer with calculations.

[4]

(e) Draw the 'dot and cross' diagram of sodium oxide formed when sodium anode is used in metal-air cells. Include all the electrons.

[2]

[Total: 10]

Q3

Extraction Process 3

This uses the electrolysis of aqueous solutions of very pure zinc sulfate. The first step in this process is the same as the first step in Extraction Process 2. The second step uses sulfuric acid made from SO_2 collected in the first step. The third step involves the electrolysis of zinc sulfate solution to form pure zinc.

$$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$$

$$ZnO(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(g)$$

$$ZnSO_4(aq) \rightarrow Zn(s) \quad \text{[electrolysis]}$$

The electrolysis of zinc sulfate solution can be carried out by using graphite or zinc as anode. When zinc is used, the anode needs to be replaced frequently.

(c) Zinc sulfate solution is electroysed in Extraction Process 3.

(i) If zinc is used as the anode, write the ionic half-equations, including state symbols, for the reaction occurring at each electrode.

Anode: ....

Cathode: .... [2]

(ii) Explain why the zinc anode needs to be replaced periodically?

....

....

.... [2]

(d) A factory replaced zinc sulfate solution with molten zinc chloride.
Suggest why molten zinc chloride may not be a good choice to replace zinc sulfate solution.

....

.... [1]

Q4

Dilute aqueous sodium chloride forms hydrogen and oxygen during electrolysis. Fig. 8.1 shows an electrolytic cell used for the process.

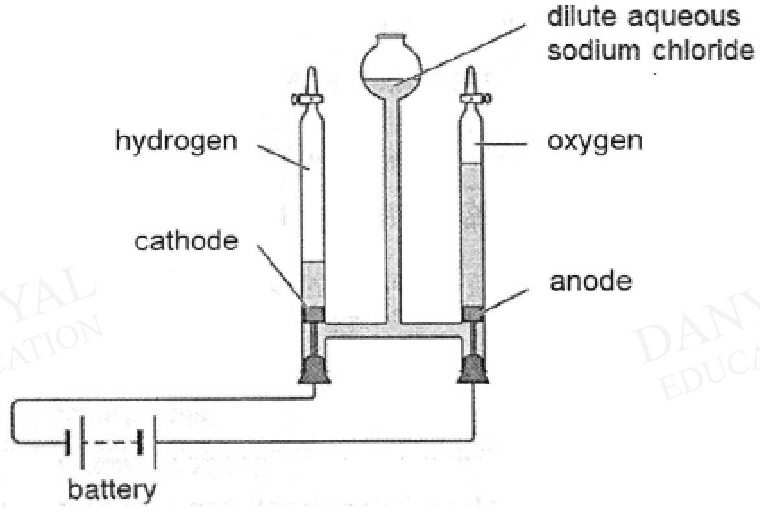


Fig. 8.1

(a) Write ionic equations for the reactions at the cathode and anode.

cathode: ....

anode: ....[2]

(b) The gases are collected and their volumes are measured. In theory, the ratio of hydrogen and oxygen should be 2:1.

Oxygen is more soluble than hydrogen in water. This changes the ratio of gases that are collected.

(i) Using the transfer of electrons, explain why the theoretical ratio of hydrogen to oxygen is 2:1.

....

....[1]

(b) (ii) Explain how and why the solubility of oxygen affects the ratio of hydrogen to oxygen that is collected.

....

....

....

....[2]

(iii) The deviation from the expected ratio is more obvious at the start of electrolysis, but becomes less noticeable after the electrolysis has been running for some time.

Suggest a reason why this happens.

....

....

....

....[2]

(c) Describe and explain the changes to the concentration of sodium chloride during the electrolysis.

....

....[1]

(d) The same apparatus can be used to electrolyse concentrated aqueous sodium chloride.

State one difference between the products obtained from the electrolysis of concentrated and dilute aqueous sodium chloride.

....

....[1]

(e) Suggest why inert electrodes are used for the electrolysis.

....

....[1]

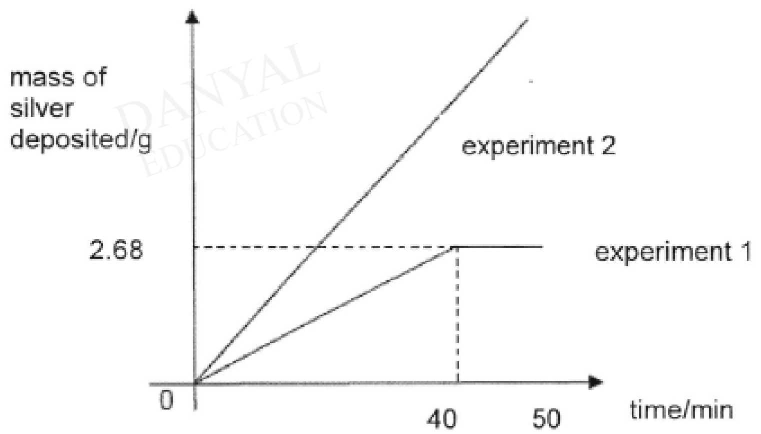
[Total: 10]

Q5

A pupil carried out two separate experiments using different electrodes in the laboratory.

In each experiment, he electrolysed 2.00 dm^3 of aqueous silver nitrate containing 2.68 g of silver ions. The same amount of current was passed in both experiments and the increase in mass of the cathode was weighed every 5 minutes for 50 minutes.

The diagram below shows the results of the two experiments.



(a) Describe how the mass of silver deposited at the cathode changes with time in each experiment.

....

....

.... [2]

(b) Write the half equation for the formation of silver at the cathode for both experiments.

.... [1]

(c) Carbon electrodes were used in experiment 1. Using this knowledge, explain the shape of the graph in experiment 1.

....

....

....

.... [2]

(d) Write the half equation for the reaction at the anode of experiment 1.

.... [1]

(e) Suggest another material for the electrodes used in experiment 2 and explain the shape of the graph.

....

....

.... [2]

(f) A circuit was connected as shown in Fig. 2.1 and a current passed through it for a period of time.

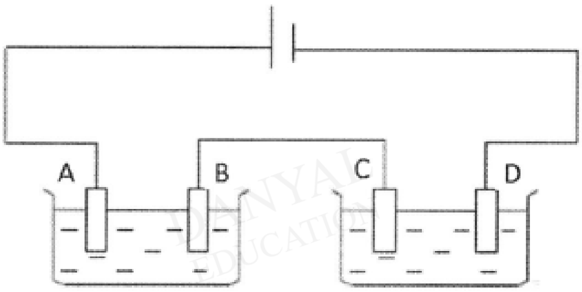


Fig. 2.1

Given that 12.8 g of copper and 14.0 g of cerium were deposited at electrodes **B** and **D** respectively, calculate the charge on a cerium ion.

[2]

[Total: 10 marks]

Answers

Electrolysis Test 2.0

Q1

(a)

cell	electrode	half ionic equation
2	copper anode	$Cu\ (s) \rightarrow Cu^{2+}\ (aq) + 2e$
	copper cathode	$Cu^{2+}\ (aq) + 2e \rightarrow Cu\ (s)$
3	graphite anode	$4OH^-\ (l) \rightarrow O_2\ (g) + 2H_2O\ (l) + 4e$
	graphite cathode	$Cu^{2+}\ (aq) + 2e \rightarrow Cu\ (s)$

[4]

(b) $Sn\ (s) \rightarrow Sn^{2+}\ (aq) + 2e$

Number of moles of Sn = 0.034/ 119 = **0.00028571** [1]

Every 2 moles of electrons, 1 mole of Sn is oxidised and 1 moles of Cu is formed.

Mass of Cu = 0.00028571 x 64 = **0.0183 g** [1]

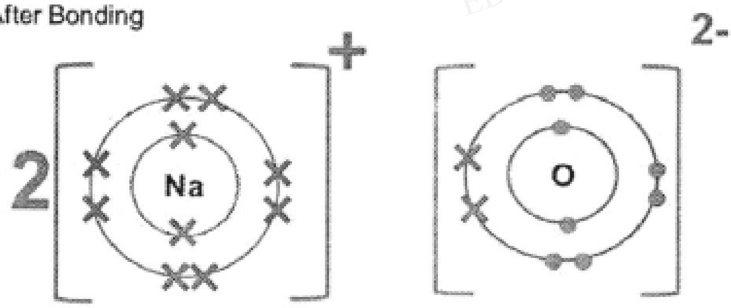
(c) The electrolyte remains blue in cell 2 and the blue colour fades in cell 3; [1]

positive Cu electrode (anode) is oxidized to replace the Cu^{2+} ions in the electrolyte in cell 2 [1]

pH does not change in cell 2 and decreases in cell 3; [1]

solution becomes acidic as hydroxide ions are preferentially discharged at the anode, leaving behind H+ ions in the electrolyte. [1]

Q2

(a)	$2Zn (s) + O_2 (g) \rightarrow 2ZnO (s)$	[1]
(b)	Arrow should point left in the electrolyte	[1]
(c)	Mass decreases	[1]
	Oxygen in the air reacts with zinc in the cell to form zinc oxide which weighs more than the original mass of zinc	[1]
(d)	For Na:	
	Number of moles of Na = 1 / 23 = **0.0434**	[1]
	Number of moles of electrons = **0.0434**	[1]
	For Ca:	
	Number of moles of Ca = 1 / 40 = **0.025**	
	Number of moles of electrons = 2 x 0.025 = **0.05**	[1]
	The cell with **calcium anode will transfer more electrons**	[1]
(e)	After Bonding 	[2]

Q3

(c)	**(i)**	Anode: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ Cathode: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	1 1
	(ii)	Zinc is the reactive anode and will be oxidised / ionised to from Zn^{2+} ions (don't accept zinc oxide is formed, without stating the oxidation of zinc) and hence needs to be replaced	1 1
(d)		A high amount of energy is required to maintain the molten zinc chloride, increasing the cost of production of zinc.	1

Q4

(a)	cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$; anode: $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$; A: if all state symbols missing R: wrong state symbol(s), partial state symbols	[1] [1]
(b)(i)	For every mole of water electrolysed, 4 mole of electrons are transferred from hydroxide ions to hydrogen ions to form 2 mole of hydrogen gas and 1 mole of oxygen gas;	[1]
(b)(ii)	The hydrogen to oxygen ratio increases/ oxygen to hydrogen ratio decreases; Less oxygen collected/ more oxygen dissolves in solution compared to hydrogen;	[1] [1]
(b)(iii)	Solution becomes saturated with oxygen; Oxygen no longer dissolves in solution and is collected as gas; A: oxygen cannot dissolve into the solution anymore	[1] [1]
(c)	Concentration of sodium chloride increases as water is being used up; A: water is being removed from solution, OH^- and H^+ ions are removed/used up/discharged from the solution	[1]
(d)	Chlorine gas produced at the anode for concentrated sodium chloride compared to oxygen gas and water for dilute sodium chloride; A: chloride gas produced at anode instead of oxygen gas, aqueous NaOH produced in the electrolysis of concentrated NaCl but not for dilute NaCl	[1]
(e)	To prevent the electrodes from reacting with the electrolytes or the products formed/ prevent the electrodes from dissolving into the solution; A: prevent reaction with solution/ions of solution, so they do not take part in reactions A: if students assume electrolysis of conc. NaCl then accept prevention of reaction with chlorine only (R: chloride)	[1]

Q5

B2 (a)	Expt 1- mass of cathode increases until 2.68g and stops increasing thereafter. Expt 2- mass of cathode kept increasing even after 2.68g.	**2**
(b)	$Ag^+(aq) + e^- \rightarrow Ag(s)$	**1**
(c)	Ag ions present in the solution discharges at the cathode. At the **anode, hydroxide ions are discharged** and **no new Ag ions are released into the solution**. Hence, once all the Ag ions in the solution are discharged, no more Ag can be formed.	**2**
(d)	$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$	**1**
(e)	**Silver electrodes** are used in experiment 2. The silver anode dissolves into the solution, **replenishing the silver ions** that have been discharged at the cathode, hence the increasing mass of silver deposited.	**2**
(f)	No of moles of Cu = 12.8g ÷ 64 = 0.2 mol No of moles of Cerium = 14.0g ÷ 140 = 0.1mol Amt of electrons to discharge Cu= 0.2 x 2 = 0.4 mole Charge of Cerium = +4 OR No of moles of Cu = 12.8g ÷ 64 = 0.2 mol No of moles of Cerium = 14.0g ÷ 140 = 0.1mol Amt of electrons to discharge Cu= 0.2 x 1 = 0.2 mole Charge of Cerium = +2	**1** **1**