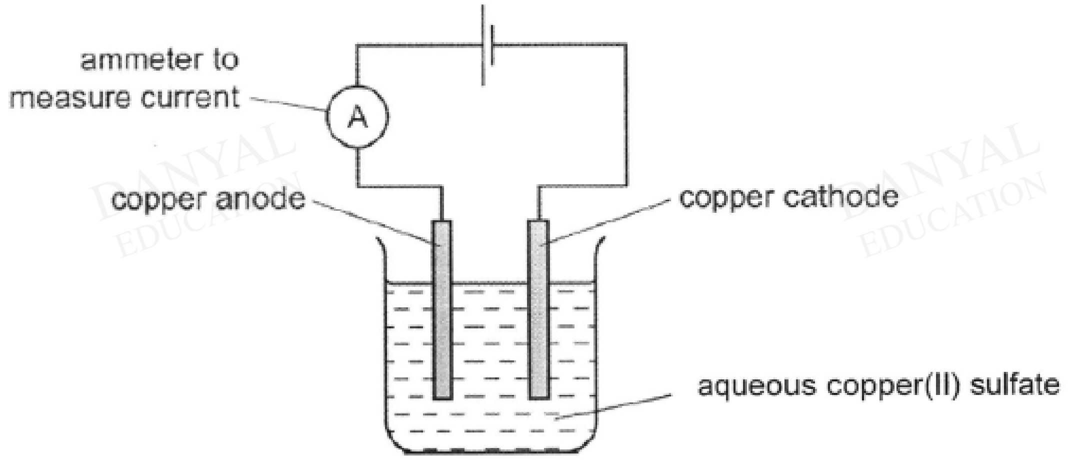
O Level Pure Chemistry Structured

Electrolysis Test 3.0

Q1

B9 A student carries out a series of experiments. In each experiment, he electrolyses aqueous copper(II) sulfate solution using the apparatus shown.



He uses the same concentration of aqueous copper(II) sulfate each time, but changes the current he passes through the solution. He runs each electrolysis for 10 minutes. The student weighs the negative electrode before and after each experiment and works out the mass of copper deposited.

The table shows the results of his experiments.

experiment	time / mins	current / Amps	mass of copper deposited / g
1	10	1.0	0.21
2	10	2.0	0.40
3	10	3.0	0.58
4	10	4.0	0.81

(a) Use data from the table to plot a graph to show the relationship between current and mass of copper deposited.
Draw a straight line of best fit, taking into account all of your plotted points. [3]

(b) Describe the relationship that your graph shows.

……………………………………………………………………………………………

……………………………………………………………………………………………[1]

(c) The student carries out another experiment. He passes an electric current of 6.0 A through a solution of copper(II) sulfate for 5 minutes.

Use your graph and the information above to predict the mass of copper that would form in the experiment.

……………………………………………………………………………………………

……………………………………………………………………………………………[1]

(d) At the end of the electrolysis, the student removes a sample of the electrolyte and puts it in a test tube. He then adds aqueous ammonia dropwise to the sample until there is no more further change. Describe and explain what the student sees.

……………………………………………………………………………………………

……………………………………………………………………………………………[2]

(e) The student carries out another electrolysis using aqueous silver nitrate and silver electrodes. His results are shown in the table.

electrolyte	time / mins	current / Amps	mass of silver deposited / g
aqueous silver nitrate	10	4.0	2.7

(i) Write an ionic half equation for the reaction that happens at the cathode.

……………………………………………………………………………………………[1]

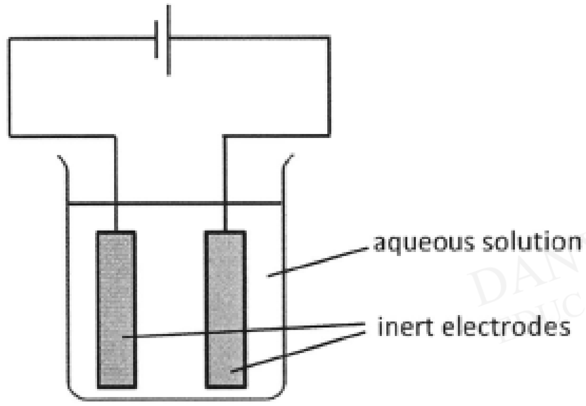
(ii) Carry out calculations to compare the difference in the number of moles of copper and the number of moles of silver that are formed when a current of 4.0 A is used for 10 minutes. Suggest an explanation for the difference in the number of moles of each metal formed.

[2]

Q2

Lead(II) nitrate and copper(II) chloride are both soluble salts.

The diagram shows the apparatus that a student used to electrolyse aqueous solutions of the salts.



(a) Complete the table of information about the electrolysis.

solution	name of products of electrolysis		ionic equation for reaction at each electrode
dilute aqueous lead(II) nitrate	at negative electrode		
	at positive electrode		
concentrated aqueous copper(II) chloride	at negative electrode		
	at positive electrode		

[4]

(b) Complete the table below to show the observations that would be made during the electrolysis.

solution	observations	
dilute aqueous lead(II) nitrate	at negative electrode	
	at positive electrode	
	in electrolyte	
concentrated aqueous copper(II) chloride	at negative electrode	
	at positive electrode	
	in electrolyte	

[3]

(c) 20 cm^3 of 0.5 $mol\,dm^{-3}$ aqueous lead(II) nitrate was mixed with 15 cm^3 of 0.8 $mol\,dm^{-3}$ aqueous copper(II) chloride.

Calculate the maximum mass of precipitate that can be obtained.

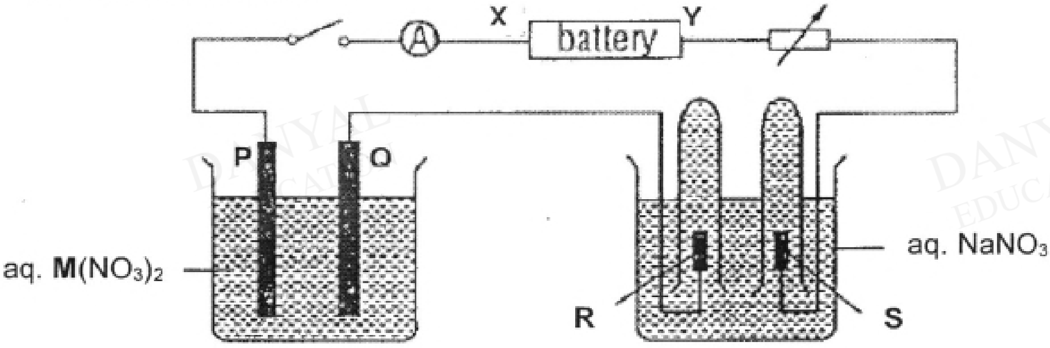
[3]

Q3

The following experiment was set up for the electrolysis of aqueous $M(NO_3)_2$ and aqueous sodium nitrate. Electrode **Q** is made up of metal **M**.

Electrodes **P**, **R** and **S** are made of platinum.

X and **Y** are terminals of the battery.
When the switch was closed, metal **M** began to deposit at electrode **P**, and gas bubbles appeared above **R** and **S**.



(a) Deduce the polarity (positive or negative) of the terminals **X** and **Y** of the battery.

X: ______________________________

Y: ______________________________ [1]

(b) Write an ionic equation for the reaction occurring at electrode **P**. Include state symbols.

______________________________ [1]

(c) After the electrolysis has been carried out for 30 minutes, 200 cm^3 of gas, measured at room temperature and presure, was collected in the test tube at electrode **R**.

(i) Name the gas collected in the test tube at electrode **R**.

______________________________ [1]

(ii) Estimate the volume of gas collected in the test tube at electrode **S**.

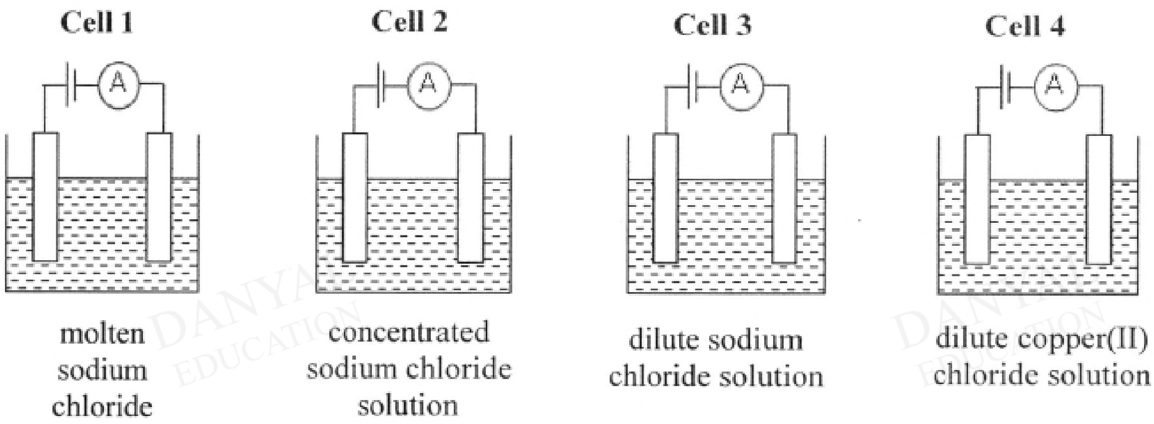
______________________________ [1]

(iii) If the relative atomic mass of metal **M** is 64, calculate the gain in mass of electrode **P**.

[3]

Q4

The diagrams show the electrolysis of four different electrolytes, using graphite electrodes.



(a) Explain why sodium metal can be extracted from molten sodium chloride solution using **Cell 1**. [1]

____

____

(b) The gases produced at the electrodes in **Cell 2** and **Cell 3** are collected and their volumes are measured.

(i) Name the gas formed at the anode of **Cell 3** and describe a test to identify the gas. [1]

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____

(ii) Explain why the gases collected at the anode in **Cells 2** and **Cell 3** are different. [2]

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____

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(iii) The volume of gas collected at the cathode in **Cell 3** is twice the volume of gas collected at the anode. By giving equations for the reactions that take place at the two electrodes, explain why this is so. [2]

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____

____

(c) (i) Explain what happens to the ions as they arrive at the cathode in **Cell 4**. [2]

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____

____

____

(ii) Explain why the reading of the current meter in **Cell 4** decreases when a few drops of lead (II) nitrate solution were added to the electrolyte. [2]

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Q5

(a) The results of an experiment involving the electrolysis of aqueous copper(II) sulfate using copper electrodes are shown below.

experiment	temperature of electrolyte / °C	current used / A	duration of electrolysis / s	mass of copper formed at the cathode / g
1	20	1.0	1000	0.329
2	20	2.0	1000	0.658
3	20	2.0	2000	1.320
4	25	2.0	2000	1.320
5	30	1.0	1000	0.329

Table 5.1

(i) State one industrial application of this electrolysis.

....[1]

(ii) John predicted that the rate of reaction is faster at a higher temperature so the mass of copper formed at the cathode should be greater for the same duration of electrolysis.

Comment on the accuracy of John's prediction using the information from **Table 5.1**.

....

....[1]

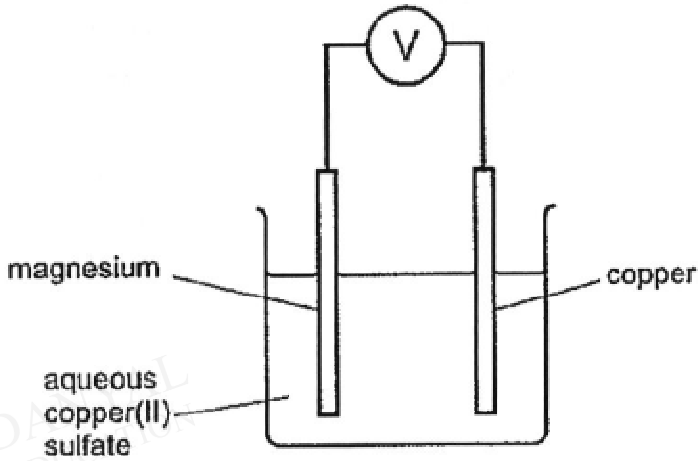
(iii) Explain why the copper(II) sulfate solution remains blue throughout the experiment.

....

....

[2]

(b) The diagram shows a simple electrochemical cell.



(i) Write down the ionic half equations, including state symbols, for the reaction at the electrodes.

cathode ....

anode ....[2]

The magnesium metal is then replaced with silver and some unknown metals **X**, **Y**, and **Z**. The voltmeter readings for the different set ups are as shown below.

metal	voltmeter reading / V
magnesium	+2.70
silver	-0.46
X	+0.48
Y	+0.78
Z	-0.70

Table 5.2

(ii) How does **Table 5.2** show that copper is more reactive than silver?

....[1]

(iii) Which combination of metals in **Table 5.2** gives the highest voltage?

....[1]

Answers

Electrolysis Test 3.0

Q1

9a	Correct axes labelled with units [1] Correct plotting of all points [1] Draw best-fit line through origin [1]
b	Mass of copper formed increases proportionally to the increase in current.
c	Student extrapolates the graph to show a mass of 1.2g obtained when a current of 6.0 A is used. In 10 mins, 1.2 g of copper is produced. In 5 mins, 0.6 g of copper is produced. [1]
d	A blue precipitate is formed which dissolves in excess aqueous ammonia to form a dark blue solution. [1] There is no change in the electrolyte. / There is copper(II) ions present in the solution. [1]
ei	$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$
eii	no. of moles of copper = 0.81 / 64 = 0.0127 mol (3 s.f.) no. of moles of silver = 2.7 / 108 = 0.0250 mol [1 m for both correct moles of copper and silver] $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$ From the calculation, no. of moles of copper is about half that of the no. of moles of silver. Each copper(II) ion needs two electrons to form a copper atom whereas each silver ion needs one electron to form a silver atom. [1] [explanation – 1m]

Q2

(a)

solution	name of products of electrolysis		ionic equation for reaction at each electrode
dilute aqueous lead(II) nitrate	at negative electrode	hydrogen (gas)	$2H^+ + 2e^- \rightarrow H_2$ (g)
	at positive electrode	oxygen (gas) and water	$4OH^-$ (aq) $\rightarrow 2H_2O$ (aq) + O_2 (g) + 4 e^-
concentrated aqueous copper(II) chloride	at negative electrode	copper (solid)	Cu^{2+} (aq) + $2e^- \rightarrow Cu$ (s)
	at positive electrode	chlorine (gas)	$2Cl^-$ (aq) $\rightarrow Cl_2$ (g) + $2e^-$

1 mark for each correct row

(b)

solution	observations	
dilute aqueous lead(II) nitrate	at negative electrode	Effervescence / bubbles
	at positive electrode	Effervescence / bubbles
	in electrolyte	No change observed
concentrated aqueous copper(II) chloride	at negative electrode	Pink / reddish brown deposit
	at positive electrode	Effervescence / bubbles
	in electrolyte	Blue colour fades/ becomes lighter lighter

1 mark for every 2 correct answers

(c)

n[$Pb(NO_3)_2$] = 0.5 x (20/1000) = 0.01 mol
n[$CuCl_2$] = 0.8 x (15/1000) = 0.012 mol [1]

Mole ratio of $Pb(NO_3)_2$: $CuCl_2$ = 1 : 1
n($PbCl_2$) = 0.01 mol [1]

m($PbCl_2$) = 0.01 x (207 + 71) = 2.78 g [1]

Q3

(a) **X**: negative **Y:** positive [1]

(b) $M^{-2+}(aq) + 2e^- \rightarrow M(s)$ [1]

(c) **(i)** hydrogen [1]

(ii) 100 cm^3 [1]

(iii) $2H^+(aq) + 2e^- \rightarrow H_2(g)$ [3]
No. of moles of hydrogen gas = 200/24000 = 0.0083333 [1]
No. of moles of electrons = 0.01667 [1]
No. of moles of M = 0.01667 /2 = 0.0083333
Mass of M = 0.0083333 × 64 = 0.533 g [1]

Q4

a		The molten sodium chloride contains **Na^+ ions** [½],which are the only cations present, and will **accept electrons/be discharged**[½] at the cathode of Cell 1.	1
b	i	Oxygen gas [½]. It relights a glowing splint [½].	1
	ii	Cell 2 contains concentrated sodium chloride. The **chloride ions will be preferentially discharged** [½] as **chlorine gas** [½], instead of the hydroxide ions. Cell 3 contains dilute sodium chloride and the **hydroxide ions are preferentially discharged** [½] instead, forming **oxygen gas** [½].	2
	iii	Cathode: $2H^+(aq) + 2e \rightarrow H_2(g)$ or $4H^+(aq) + 4e \rightarrow 2H_2(g)$ **[½]** Anode $4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e$ [½] For the same number of electrons [½], the volume of hydrogen gas produced is twice the volume of oxygen gas produced. [½]	2
c	i	**H^+ and Cu^{2+} ions** [½] both arrive at the cathode in Cell 4. Since **copper is less reactive than hydrogen in the reactivity series** [½], it will be **preferentially discharged** [½] instead at the cathode, forming **copper metal** [½].	2
	ii	An insoluble salt of $PbCl_2$ [1] is formed It does not contain delocalised ions [1] to conduct electricity. OR it removes Cl^- ions [1]	2

Q5

a)(i) Purification of copper; (Reject: electroplating)

a(ii) Comparing experiment 3 and 4 OR 1 and 5, when the temperature increases for the same current used and duration of electrolysis, the mass of copper formed remains the same.

a(iii)
For every copper atom that is oxidised at the anode, a copper(II) ion is discharged/reduced at the cathode;
There is no net change in concentration of copper(II) ions throughout the experiment;

b(i)
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$;
$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$;

b(ii)
The voltmeter reading is negative.

b(iii) Magnesium and Z