Lisachem Materials

Chemistry Physics Biology
Psychology

VCE CHEMISTRY 2007
SUPPLYING AND USING ENERGY UNIT 4

Time allowed: 50 minutes
Total marks: 40

SECTION A
Contains 12 multiple choice questions

SECTION B
4 Extended response questions

A data sheet and multiple choice answer sheet are provided. Answer extended response questions in the space provided. Use the marks and time allowed as a guide to how much time you should spend answering each question.

Lisachem Materials
PO Box 2018, Hampton East, Victoria, 3188
Ph: (03) 9598 4564 Fax: (03) 8677 1725
Email: orders@lisachem.com.au

						relative atomic number symbol name relative atomic mass	1 H Hydrogen 1.0										2 He Helium 4.0
3 Li Lithium 6.9	4 Be Beryllium 9.0											5 B Boron 10.8	6 C Carbon 12.0	7 N Nitrogen 14.0	8 O Oxygen 16.0	9 F Fluorine 19.0	10 Ne Neon 20.2
11 Na Sodium 23.0	12 Mg Magnesium 24.3											13 Al Aluminium 27.0	14 Si Silicon 28.1	15 P Phosphorus 31.0	16 S Sulfur 32.1	17 Cl Chlorine 35.5	18 Ar Argon 39.9
19 K Potassium 39.1	20 Ca Calcium 40.1	21 Sc Scandium 44.9	22 Ti Titanium 47.9	23 V Vanadium 50.9	24 Cr Chromium 52.0	25 Mn Manganese 54.9	26 Fe Iron 55.9	27 Co Cobalt 58.9	28 Ni Nickel 58.7	29 Cu Copper 63.6	30 Zn Zinc 65.4	31 Ga Gallium 69.7	32 Ge Germanium 72.6	33 As Arsenic 74.9	34 Se Selenium 79.0	35 Br Bromine 79.9	36 Kr Krypton 83.8
37 Rb Rubidium 85.5	38 Sr Strontium 87.6	39 Y Yttrium 88.9	40 Zr Zirconium 91.2	41 Nb Niobium 92.9	42 Mo Molybdenum 95.9	43 Tc Technetium 98.1	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9	46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3
55 Cs Caesium 132.9	56 Ba Barium 137.3	57 La Lanthanum 138.9	72 Hf Hafnium 178.5	73 Ta Tantalum 180.9	74 W Tungsten 183.8	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.1	79 Au Gold 197.0	80 Hg Mercury 200.6	81 Tl Thallium 204.4	82 Pb Lead 207.2	83 Bi Bismuth 209.0	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (261)	105 Ha **Hahnium** (262)	106 Sg Seaborgium (266)	107 Ns Neilsbohrium (264)	108 Hs Hassium (269)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (272)	111 Rg Roentgenium (272)	112 Uub Ununbium (277)		114 Uuq Ununquadium (289)				

Lanthanide series	58 Ce Cerium 140.1	59 Pr Praseodymium 140.9	60 Nd Neodymium 144.2	61 Pm Promethium (145)	62 Sm Samarium 150.3	63 Eu Europium 152.0	64 Gd Gadolinium 157.2	65 Tb Terbium 158.9	66 Dy Dysprosium 162.5	67 Ho Holmium 164.9	68 Er Erbium 167.3	69 Tm Thulium 168.9	70 Yb Ytterbium 173.0	71 Lu Lutetium 175.0
Actinide series	90 Th Thorium 232.0	91 Pa Protactinium 231.0	92 U Uranium 238.0	93 Np Neptunium 237.1	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (254)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (255)	103 Lr Lawrencium (256)

Data Sheet VCE Chemistry 2007 Supplying & Using Energy Unit 4

Physical Constants

F = 96 500 C mol^{-1}
R = 8.31 JK^{-1} mol^{-1}
V_m (STP) = 22.4 L mol^{-1}
V_m (SLC) = 24.5 L mol^{-1}
Specific heat of water = 4.18 J g^{-1} ^{o}C^{-1}

Ideal gas equation

$pV = nRT$

The Electrochemical Series

		$E°$ in volt
$F_2(g) + 2e^-$	$\rightarrow 2F^-(aq)$	+ 2.87
$H_2O_2(aq) + 2H^+(aq) + 2e^-$	$\rightarrow 2H_2O(l)$	+ 1.77
$Au^+(aq) + e^-$	$\rightarrow Au(s)$	+ 1.68
$MnO_4^-(aq) + 8H^+(aq) + 5e^-$	$\rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+ 1.50
$Cl_2(g) + 2e^-$	$\rightarrow 2Cl^-(aq)$	+ 1.36
$O_2(g) + 4H^+(aq) + 4e^-$	$\rightarrow 2H_2O(l)$	+ 1.23
$Br_2(l) + 2e^-$	$\rightarrow 2Br^-(aq)$	+ 1.09
$Ag^+(aq) + e^-$	$\rightarrow Ag(s)$	+ 0.80
$Fe^{3+}(aq) + e^-$	$\rightarrow Fe^{2+}(aq)$	+ 0.77
$I_2(s) + 2e^-$	$\rightarrow 2I^-(aq)$	+ 0.54
$O_2(g) + 2H_2O(l) + 4e^-$	$\rightarrow 4OH^-(aq)$	+ 0.40
$Cu^{2+}(aq) + 2e^-$	$\rightarrow Cu(s)$	+ 0.34
$CO_2(g) + 8H^+(aq) + 8e^-$	$\rightarrow CH_4(g) + 2H_2O(l)$	+ 0.17
$S(s) + 2H^+(aq) + 2e^-$	$\rightarrow H_2S(g)$	+ 0.14
$2H^+(aq) + 2e^-$	$\rightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^-$	$\rightarrow Pb(s)$	- 0.13
$Sn^{2+}(aq) + 2e^-$	$\rightarrow Sn(s)$	- 0.14
$Ni^{2+}(aq) + 2e^-$	$\rightarrow Ni(s)$	- 0.23
$Co^{2+}(aq) + 2e^-$	$\rightarrow Co(s)$	- 0.28
$Fe^{2+}(aq) + 2e^-$	$\rightarrow Fe(s)$	- 0.44
$Zn^{2+}(aq) + 2e^-$	$\rightarrow Zn(s)$	- 0.76
$2H_2O(l) + 2e^-$	$\rightarrow H_2(g) + 2OH^-(aq)$	- 0.83
$Mn^{2+}(aq) + 2e^-$	$\rightarrow Mn(s)$	- 1.03
$Al^{3+}(aq) + 3e^-$	$\rightarrow Al(s)$	- 1.67
$Mg^{2+}(aq) + 2e^-$	$\rightarrow Mg(s)$	- 2.34
$Na^+(aq) + e^-$	$\rightarrow Na(s)$	- 2.71
$Ca^{2+}(aq) + 2e^-$	$\rightarrow Ca(s)$	- 2.87
$K^+(aq) + e^-$	$\rightarrow K(s)$	- 2.93
$Li^+(aq) + e$	$\rightarrow Li(s)$	- 3.02

Student Name………………………………………

VCE Chemistry 2007 Supplying and Using Energy Year 12 Unit 4

SECTION A

MULTIPLE CHOICE ANSWER SHEET

Instructions:

For each question choose the response that is correct or best answers the question.
Circle the chosen response on this answer sheet.
Only circle **one** response for each question.

Question 1.	A	B	C	D
Question 2.	A	B	C	D
Question 3.	A	B	C	D
Question 4.	A	B	C	D
Question 5.	A	B	C	D
Question 6.	A	B	C	D
Question 7.	A	B	C	D
Question 8.	A	B	C	D
Question 9.	A	B	C	D
Question 10.	A	B	C	D
Question 11.	A	B	C	D
Question 12.	A	B	C	D

VCE Chemistry 2007 Supplying and Using Energy Unit 4

SECTION A - [12 marks, 15 minutes]

This section contains 12 multiple choice questions.
For each question choose the response that is correct or best answers the question.
Indicate your answer on the answer sheet provided.
*(Choose only **one** answer for each question.)*

Question 1
When a piece of iron wire is placed in an aqueous solution of iron(III) sulfate;

A. the iron wire will act as an oxidant;
B. the iron wire will act as a reductant;
C. the iron(III) ions will be oxidised;
D. no significant reaction will occur.

Question 2
The mass of a metallic object increased by 0.440 g as a result of being electroplated with nickel using an aqueous nickel(II) cyanide solution. What electrical charge was required to produce this plating of nickel?

A. 1450 C.
B. 725 C.
C. 1.50×10^{-2} C.
D. 360 C.

Question 3
A primary cell is;

A. a cell that produces a constant voltage through out its life;
B. a galvanic cell that can be recharged;
C. a galvanic cell that cannot be recharged;
D. an electrolytic cell that cannot be recharged.

Question 4
Compared to energy produced from fossil fuels, nuclear energy;

A. is a renewable resource;
B. will produce more energy from a given mass of fuel;
C. will produce less energy from a given mass of fuel;
D. is cheaper to set up and operate, making it cheaper to the consumer.

Question 5
An electric current is passed through an aqueous solution containing aluminium sulfate, tin(II) sulfate and zinc sulfate, using platinum electrodes. The initial product formed at the cathode would be;

A. hydrogen gas;
B. aluminium metal;
C. zinc metal;
D. tin metal.

Question 6

In an experimental investigation a group of students found that 1680 J of energy were required to increase the temperature of 126.3 g of ethanol by 5.68 ºC. What is the specific heat capacity for ethanol?

A. -296 J $^{\circ}C^{-1}$ g^{-1}.
B. -2.34 J $^{\circ}C^{-1}$ g^{-1}.
C. 75.5 J $^{\circ}C^{-1}$ g^{-1}.
D. 2.34 J $^{\circ}C^{-1}$ g^{-1}.

Question 7

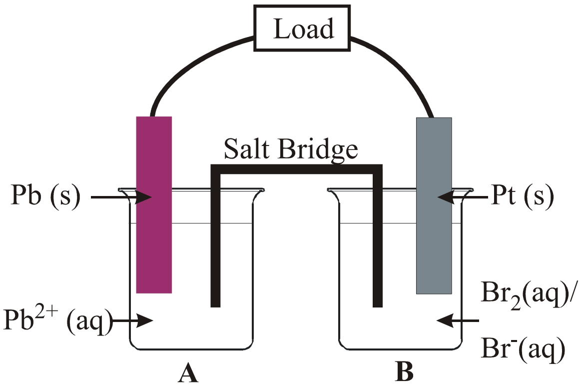
The reaction that occurs in a methane-oxygen fuel cell when it is producing energy can be represented by the chemical equation;

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

The process that occurs at the positive electrode of this cell would involve

A. the reduction of methane;
B. the oxidation of methane;
C. the reduction of oxygen;
D. the oxidation of oxygen.

Questions 8 and 9 refer to the following galvanic cell.



Question 8

In the galvanic cell shown in the diagram, the electrons will flow through the external circuit from the;

A. platinum electrode to lead electrode and anions will move from the salt bridge into half-cell A;
B. lead electrode to platinum electrode and anions will move from the salt bridge into half-cell A;
C. lead electrode to platinum electrode and anions will move from the salt bridge into half-cell B;
D. platinum electrode to lead electrode and anions will move from the salt bridge into half-cell B.

Question 9

The chemical half-equation for the reaction occurring in half-cell B when this cell is discharging would be;

A. $2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$;
B. $Pt(s) \rightarrow Pt^{2+}(aq) + 2e^-$;
C. $Pt(s) + 2e^- \rightarrow Pt^{2-}(aq)$;
D. $Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$.

Question 10

The electrolyte most commonly used to produce sodium in a Down's cell is;

A. a molten mixture of calcium chloride and sodium chloride;
B. a molten mixture of sodium chloride and magnesium chloride;
C. molten sodium chloride;
D. a concentrated aqueous solution containing a mixture calcium chloride and sodium chloride.

Question 11

When a 0.145 g piece of aluminium wire was burnt in a bomb calorimeter, with a calibration factor of 5980 J ºC^{-1}, a 0.76 ºC temperature increase occurred. What is the enthalpy change for the combustion of aluminium represented by the chemical equation?

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

A. -3400 kJ mol^{-1}.
B. -4.5 kJ mol^{-1}.
C. -840 kJ mol^{-1}.
D. -1700 kJ mol^{-1}.

Question 12

The lead-acid battery is an example of secondary cells. The reaction that occurs when the battery is recharged can be represented by the chemical equation;

$$2PbSO_4(s) + 2H_2O(l) \rightarrow PbO_2(s) + Pb(s) + 4H^+(aq) + 2SO_4^{2-}(aq)$$

Which one of the following statements is **incorrect** about what occurs when the battery is discharging?

A. The lead acts as the reductant.
B. The plate coated with the PbO_2 will be the cathode.
C. The pH of the solution will decrease.
D. The concentration of the sulfate ions in the electrolyte will decrease.

End of Section A

SECTION B - [28 marks, 35 minutes]

This section contains four questions, numbered 1 to 4.
All questions should be answered in the spaces provided.
The mark allocation and approximate time that should be spent on each question are given.

Question 1 - [6 marks, 8 minutes]
The electrical calibration data for a solution calorimeter is shown below.

Volume of solution	100.0 mL
Initial Temperature	18.27 ºC
Voltage	6.10 V
Current	2.13 A
Time	10.0 minutes
Final Temperature	32.93 ºC

a. Determine the calibration factor for this calorimeter.

[2 marks]

b. When 0.370 g of zinc powder was added to 100.0 mL of 1.0 M aqueous copper(II) sulfate solution in the calorimeter, a 2.32 ºC temperature increase occurred.

i. Write an appropriate chemical equation for the reaction between the zinc and the copper(II) solution.

[1 mark]

ii. Calculate the change in enthalpy for this reaction.

[3 marks]

Question 2 - [9 marks, 11 minutes]

In an investigation of an electrolytic reaction a group of students used a 4.0 V power supply to pass a current of 0.44 A through a 100 mL sample of 1.0 M aqueous zinc sulfate solution for 60.0 minutes using unreactive electrodes. During the electrolysis the students observed a gas being produced at one of the electrodes, and a solid deposited on the other.

a. Write an appropriate chemical half-equation to represent the reaction that would occur at the anode.

[1 mark]

b. Write an appropriate chemical half-equation to represent the reaction that would occur at the cathode.

[1 mark]

c. Calculate the charge passed through the solution.

[1 mark]

d. What would be the expected mass of solid deposited on the electrode at the completion of the electrolysis?

[2 marks]

e. What volume of gas at SLC should be produced at the other electrode?

[2 marks]

f. How would the pH of the solution at the end of the electrolysis compare with the initial pH of the solution?

[1 mark]

g. Another group of students, using similar apparatus and conditions, allowed the electrolysis to continue for 5 hours before the power supply was turned off at the end of the school day. The electrodes were left in the solution until the students returned and when they measured the mass of the electrode containing the deposit they found it to be significantly less than they expected. What is one explanation for their observation?

[1 mark]

Question 3 - [8 marks, 10 minutes]

a. A model car uses a small hydrogen-oxygen fuel cell to supply its electrical energy. Assume that the cell operates at standard conditions and uses an acidic electrolyte.

   i. Write an appropriate chemical half-equation for the reaction that occurs at the anode.

[1 mark]

   ii. Write an appropriate chemical half-equation for the reaction that occurs at the cathode.

[1 mark]

   iii. What voltage would be produced by a single cell?

[1 mark]

   iv. What current would be produced if a cell consumed 4.9 mL of hydrogen at SLC over a 60 s period?

[3 marks]

b. It is widely proposed that hydrogen should be used as a fuel for vehicles as it would produce no greenhouse gas emissions.

  i. What is one advantage of using a hydrogen-oxygen fuel cell over an internal combustion engine using hydrogen as its fuel in such vehicles?

[1 mark]

  ii. What is one disadvantage of using a hydrogen-oxygen fuel cell over an internal combustion engine using hydrogen as its fuel in such vehicles?

[1 mark]

Question 4 - [5 marks, 6 minutes]
Methanol can be used as a fuel, and its combustion reaction can be represented by the thermochemical equation;

$$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l) \qquad \Delta H = -1524 \text{ kJ mol}^{-1}$$

a. What mass of methanol must be burnt to produce 260 kJ of energy?

[2 marks]

b. Assuming all of this energy was absorbed by a 1.0 L water sample, by how much would its temperature increase?

[2 marks]

c. Why is the energy released by the reaction described by the thermochemical equation,

$$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g) \qquad \Delta H = -1348 \text{ kJ mol}^{-1}$$

less than that for the reaction in a. above?

[1 mark]

END OF TASK

Suggested Answers VCE Chemistry 2007 Supplying & Using Energy Unit 4

SECTION A [1 mark per question.]

Q1 B. From the electrochemical series the appropriate half-equations are:

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ E° = +0.77

$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ E° = -0.44

When iron wire is placed in a solution containing iron(III) ions, the **iron will act as the reductant and be oxidised** while the **iron(III) ions will act as the oxidant and be reduced**, yielding the overall reaction

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

Q2 A. The redox half-equation for the reduction of nickel(II) ions to nickel metal is

$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$

$n(Ni) = m / M = 0.440 / 58.7 = 7.50\times10^{-3}$ mol

$n(e^-) = 2\times n(Ni) = 2 \times 7.5\times10^{-3} = 1.50\times10^{-2}$ mol

$Q = n \times F = 1.5\times10^{-2} \times 96500 =$ **1450 C**

Q3 C. A **primary cell is a galvanic cell that cannot be recharged** by the application of an external power supply. This is because the products formed when the cell discharges migrate away from the electrodes and may be consumed in other side reactions that occur within the cell.

Q4 B. Nuclear energy produces more energy than fossil fuels for a given mass of reactant. This is because the production of energy from fossil fuels involves chemical reactions which involve the making and breaking of chemical bonds and the subsequent rearrangement of electrons, whereas nuclear energy involves the rearrangement of protons and neutrons to produce different elements and is a more energetic process.

Q5 D. In electrochemical cells reduction always occurs at the cathode, therefore it is necessary to identify the strongest oxidant from the materials present in the cell using the electrochemical series. The appropriate chemical half-equations are

$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$ E° = -0.14

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ E° = -0.76

$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ E° = -0.83

$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$ E° = -1.67

The **tin(II) ions are the strongest oxidant therefore will be the first to be reduced to tin metal** at the cathode.

Q6 D. The specific heat of a substance is the amount of energy required to increase the temperature of 1.0 g of the substance by 1 °C, without any change in state.

1680 J increased the temperature of the sample by 5.68 °C

Therefore to increase the temperature by 1 °C, 1680/5.68 = 296 J is required.

Therefore to change the temperature of 1g of ethanol (the heat capacity).

$C_p = E / m = 296 / 126.3 = 2.34$ J °C^{-1} g^{-1}

Q7 C. A fuel cell is a galvanic cell and reduction occurs at the cathode which is positively charged. In the reaction;

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

oxygen is the oxidant and is being reduced.

The chemical half-equation for this reaction will be;

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

Q8 B. The appropriate chemical half-equations from the electrochemical series are;

$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$ E° = +1.09

$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$ E° = -0.13

Since the **lead is the stronger reductant it will be oxidised and the electrons will move from the lead electrode to the platinum electrode** through the external circuit.

As **Pb^{2+} ions are being formed in half-cell A, anions will need to flow from the salt bridge into this half-cell** to maintain ionic charge neutrality.

Q9 D. The Br_2 is the stronger oxidant and will be reduced when the cell is discharging.

$\mathbf{Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)}$

Q10 A. The Down's cell is used to produce sodium, which is a strong reductant that will be oxidised by water, therefore aqueous solutions of electrolyte cannot be used in this cell. **The electrolyte must be in the molten state**. Pure sodium chloride has a melting temperature of 801 °C, but when **mixed with calcium chloride** a mixture with a melting temperature of about 600 °C can be obtained. This is the electrolyte commonly used to produce sodium in this cell. The sodium ions are a stronger oxidant compared to the calcium ions and will therefore be reduced before the calcium ions.

Q11 A. $E = CF \times \Delta T = 5980 \times 0.76 = 4.5\times10^3$ J = 4.5 kJ

$n(Al) = m / M = 0.145 / 27.0 = 5.37\times10^{-3}$ mol

5.37×10^{-3} mol of Al releases 4.5 kJ of energy

1 mol of Al releases $4.5 / 5.37\times10^{-3} = 840$ kJ of energy

$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

The equation is for 4 mol of Al, therefore 4 × 840 = 3360 kJ = 3400 kJ

The reaction is exothermic therefore $\Delta H < 0$

$\mathbf{\Delta H = -3400\ kJ\ mol^{-1}}$

Q12 C. The chemical equation for the discharge reaction is the reverse, therefore

$PbO_2(s) + Pb(s) + 4H^+(aq) + 2SO_4^{2-}(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$

In this reaction the Pb is being oxidised so is the reductant, therefore A is correct.

The PbO_2 is the oxidant, therefore is being reduced and reduction occurs at the cathode, B is correct.

H^+ ions are being consumed, therefore $[H^+]$ will decrease leading to an increase in the pH ($pH = -\log_{10}[H^+]$), therefore **C is incorrect**.

SO_4^{2-} ions are being consumed, therefore the concentration of these in the electrolyte solution will decrease, D is correct.

SECTION B

Question 1 - [6 marks, 8 minutes]

a. The calibration factor (CF) for a calorimeter is the amount of energy require to change the temperature of the calorimeter and its contents by 1 °C.

$E = V \times I \times t$ (time in seconds)

$E = 6.10 \times 2.13 \times (10.0 \times 60) = 7796$ J = 7800 J **[1 mark]**

$\Delta T = 32.93 - 18.27 = 14.66$ °C

$CF = E / \Delta T = 7800 / 14.66 =$ **532 J °C^{-1}** **[1 mark]**

b. i. An appropriate chemical equation can be either an ionic equation or a full equation. **[1 mark]**

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

$\mathbf{Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)}$

ii. Firstly must determine if either reactant is in excess and which is the limiting one.

$n(Zn) = m / M = 0.370 / 65.4 = 5.66\times10^{-3}$ mol

$n(Cu^{2+}) = c \times V = 1.0 \times (100.0/1000) = 1.0\times10^{-1}$ mol

Clearly the Zn is the limiting reactant. **[1 mark]**

$E = CF \times \Delta T = 532 \times 2.32 = 1.23\times10^{3}$ J = 1.23 kJ **[1 mark]**

5.66×10^{-3} mol of Zn releases 1.23 kJ

1.0 mol of Zn releases $1.23/5.66\times10^{-3} = 217$ kJ mol^{-1}

The temperature increases so the reaction is exothermic, thus $\Delta H < 0$

$\mathbf{\Delta H = -217}$ **kJ mol^{-1} [1 mark]**

Question 2 - [9 marks, 11 minutes]

a. This is an electrolytic cell and as with all electrochemical cells oxidation occurs at the anode. The only reductant in the electrolyte is water, therefore this must be oxidised.

$\mathbf{2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-}$. **[1 mark]**

b. Reduction will occur at the cathode. The oxidants present in the electrolyte are zinc ions and water. The zinc ions are the stronger oxidants therefore these will be reduced first.

$\mathbf{Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)}$ **[1 mark]**

c. $Q = I \times t$ (Time in seconds)

$Q = 0.44 \times (60.0\times60) = 1580$ C = **1600 C [1 mark]**

d. The zinc is deposited on the cathode: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

$n(e^-) = Q / F = 1600 / 96500 = 1.66\times10^{-2}$ mol **[1 mark]**

$n(Zn) = ½\, n(e^-) = ½ \times 1.66\times10^{-2} = 8.3\times10^{-3}$ mol

$m(Zn) = n \times M = 8.3\times10^{-3} \times 65.4 =$ **0.54 g [1 mark]**

e. Oxygen is produced at the anode: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

From d. $n(e^-) = 1.66\times10^{-2}$ mol

$n(O_2) = ¼\, n(e^-) = ¼ \times 1.66\times10^{-2} = 4.1\times10^{-3}$ mol **[1 mark]**

$V(O_2) = n \times V_m = 4.1\times10^{-3} \times 25.4 =$ **0.10 L [1 mark]**

f. The oxidation reaction that occurs at the anode produces H^+ ions, therefore compared with the initial concentration of H^+ ions the final concentration will be higher. The **pH of the solution after the electrolysis will therefore be lower than its original pH. [1 mark]**

g. Some possible explanations could include: **[1 mark allocated]**

The **H^+ ions formed at the anode during the reaction reacted with some of the zinc deposited on the cathode**, thereby reducing the mass.

As the reaction proceeds the concentration of the zinc ions near the cathode decreases and as a result **some of the water (or H^+ ions) present could have been reduced** instead of the zinc ions.

Some of the zinc deposited on the cathode physically detached from the electrode.

Question 3 - [8 marks, 10 minutes]

a. i. Oxidation at the anode, therefore the hydrogen will be oxidised.

$\mathbf{H_2(g) \rightarrow 2H^+(aq) + 2e^-}$ **[1 mark]**

ii. Reduction at the cathode, therefore the oxygen is reduced

$\mathbf{O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O}$ **[1 mark]**

iii. From the electrochemical series the standard potentials for the half-reactions are:

$H_2(g) \rightarrow 2H^+(aq) + 2e^-$ $E° = 0.00$

$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O$ $E° = +1.23$

The difference between the standard potentials give the voltage produced by the cell when it is operating under standard conditions.

Gases	101.3 kPa
Solutions	1.00 M
Temperature	25 °C

V = 1.23 – 0 = **1.23 V [1 mark]**

iv. $n(H_2) = V / V_m = (4.9/1000) / 24.5 = 2.0\times10^{-4}$ mol **[1 mark]**

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

$n(e^-) = 2 \times n(H_2) = 2 \times 2.0\times10^{-4} = 4.0\times10^{-4}$ mol

$Q = n(e^-) \times F = 4.0\times10^{-4} \times 96500 = 39$ C **[1 mark]**

I = Q / t = 39 / 60 = **0.64 A [1 mark]**

b. i. Possible answers include: **[Total marks allocated = 1 mark]**

Fuel calls are more efficient than internal combustion engines therefore more usable energy is available for a given amount of hydrogen gas.

Fuel cells produce electricity directly therefore would produce quieter vehicles compared to internal combustion engines.

ii. Possible answers include: **[Total marks allocated = 1 mark]**

Fuel cells tend to be more expensive than internal combustion engines.

Fuel cells require some expensive auxiliary equipment to operate them.

Current fuel cells require the hydrogen gas to be free from impurities that can cause side reactions on the electrode surfaces reducing their efficiency, whereas internal combustion engines can burn hydrogen mixed with other gases.

Question 4 - [5 marks, 6 minutes]

a. From the thermochemical equation

$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$ $\Delta H = -1524$ kJ mol^{-1}

2 mol of CH_3OH will release 1524 kJ of energy

1 mol of CH_3OH will release 1524/2 = 762 kJ of energy **[1 mark]**

$n(CH_3OH) = 260 / 762 = 0.34$ mol

$M(CH_3OH) = 12.0 + 4\times1.0 + 16.0 = 32.0$ g mol^{-1}

$m(CH_3OH) = n \times M = 0.34 \times 32.0 =$ **11 g [1 mark]**

b. The heat capacity for water is 4.184 J mL^{-1} °C^{-1}

Therefore the energy required to increase 1.0 L of water by 1 °C is

4.184 × 1000 = 4184 J = 4.184 kJ **[1 mark]**

$\Delta T = 260 / 4.184 =$ **62 °C [1 mark]**

c. The two thermochemical equations are:

$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$ $\Delta H = -1524$ kJ mol^{-1}

$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$ $\Delta H = -1348$ kJ mol^{-1}

The difference between these two reactions is the state of the water produced. In the first the water is in the liquid state whereas in the second it is in the gaseous state. **Energy is required to change a liquid into a gas, therefore the energy released by the second reaction is less than that of the first**. **[1 mark]**

The difference between the two reactions:

$4H_2O(l) \rightarrow 4H_2O(g)$ $\Delta H = -1348 - (-1524) = 176$ kJ mol^{-1}

$H_2O(l) \rightarrow H_2O(g)$ $\Delta H = +44$ kJ mol^{-1}