VCE CHEMISTRY 2010

YEAR 12 TRIAL EXAM UNIT 4

CONDITION OF SALE:

Time allowed: 90 minutes
Total marks: 76

20 Multiple Choice Questions
7 Short Answer Questions

An Answer Sheet is provided for Section A.
Answer all questions in Section B in the space provided.

To download the Chemistry Data Book please visit the VCAA website:
http://www.vcaa.vic.edu.au/vce/studies/chemistry/chem1_sample_2008.pdf *Page 20*

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Student Name....…………………………..

VCE Chemistry 2010 Year 12 Trial Exam Unit 4

Student Answer Sheet

Instructions for completing test. Use only a 2B pencil. If you make a mistake erase and enter the correct answer. Marks will not be deducted for incorrect answers.

Write your answers to the Short Answer Section in the space provided directly below the question. There are 20 Multiple Choice questions to be answered by circling the correct letter in the table below.

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
Question 13	A	B	C	D	*Question 14*	A	B	C	D
Question 15	A	B	C	D	*Question 16*	A	B	C	D
Question 17	A	B	C	D	*Question 18*	A	B	C	D
Question 19	A	B	C	D	*Question 20*	A	B	C	D

VCE Chemistry 2010 Year 12 Trial Exam Unit 4

Multiple Choice Questions – Section A

Section A consists of 20 multiple-choice questions.
Section A is worth approximately 25 per cent of the marks available.
Choose the response that is **correct** *or* **best answers** *the question.*
Indicate your choice on the answer sheet provided.

Question 1

Ethanol and ethanoic acid react to produce the ester ethyl ethanoate according to the equation

$$CH_3CH_2OH(l) + CH_3COOH(l) \rightleftharpoons CH_3COOCH_2CH_3(l) + H_2O(l)$$

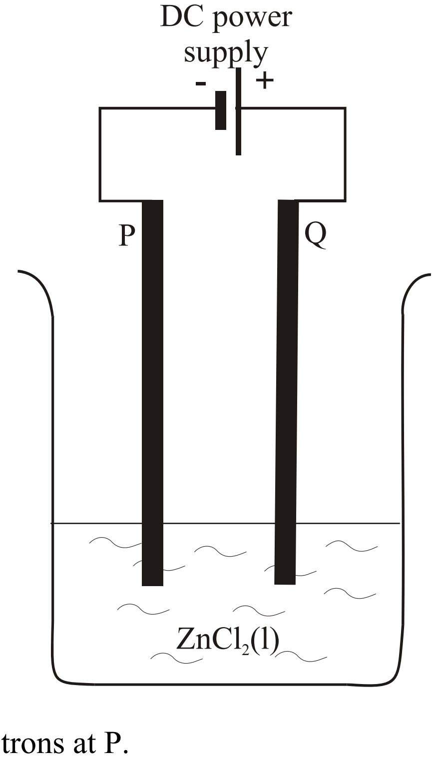
1 mol of ethanol is added to 1 mol of ethanoic acid and allowed to react in the presence of an acid catalyst.

When this reaction is as equilibrium,

A. the concentration fraction will equal 1.
B. the concentration of ethyl ethanoate equals the concentration of ethanol.
C. the rate of production of ethyl ethanoate equals the rate of its hydrolysis.
D. the pH will be 7.

Question 2

In the electrochemical cell represented below



A. chloride ions lose electrons at P.
B. chloride ions gain electrons at Q.
C. chloride ions gain electrons at P.
D. chloride ions lose electrons at Q.

Question 3

An equilibrium system represented by the equation below has $K = 2.3 \text{x} 10^4$ at 30°C and $K = 92$ at 250°C.

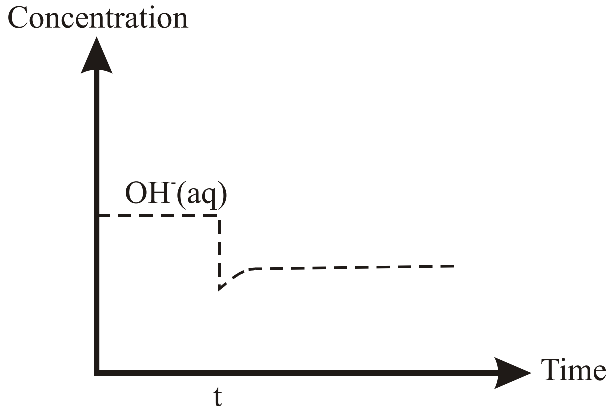
$$2X + Y \rightleftharpoons Z$$

On the basis of the information,

A. $\Delta H < 0$ and the yield of Z increases with decreasing temperature.
B. $\Delta H < 0$ and the yield of Z decreases with decreasing temperature.
C. $\Delta H > 0$ and the yield of Z increases with decreasing temperature.
D. $\Delta H > 0$ and the yield of Z decreases with decreasing temperature.

Question 4

The concentration-time graph below shows how the [OH^-] changes as a result of a change imposed on an equilibrium system. The self-ionisation of water is endothermic.



The graph is consistent with

A. the addition of a small amount of 1 M NaOH(aq) to an aqueous solution of ethanoic acid.
B. cooling pure water.
C. heating pure water.
D. adding 100 mL water to 100 mL of NH_3(aq).

Question 5

When 25 mL of 1.0 M NaOH at 24°C, is added to 25 mL of 1.0 M HCl at 24°C, the temperature of the solution formed is 30°C.

Which combination of reactants, both initially at 24°C, might also be expected to produce a temperature of 30°C when mixed?

A. 25 mL of 2.0 M NaOH and 25 mL of 2.0 M HCl.
B. 50 mL of 2.0 M NaOH and 50 mL of 2.0 M HCl.
C. 50 mL of 1.0 M NaOH and 50 mL of 1.0 M HCl.
D. 100 mL of 2.0 M NaOH and 100 mL of 2.0 M HCl.

Question 6

The self-ionisation constant of pure water, K_w, at 45°C is 4.0×10^{-14}.
On the basis of this information, in pure water at 15°C

A. the pH is < 7
B. the pH > 7
C. $[H_3O^+] > [OH^-]$
D. $[H_3O^+] < [OH^-]$

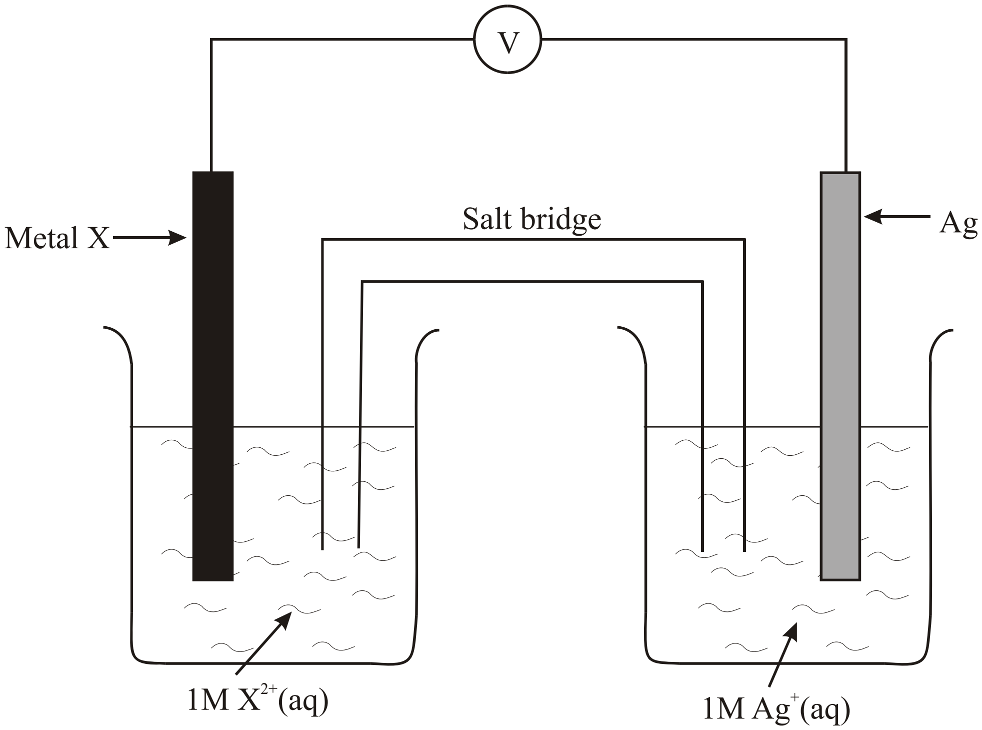
Question 7

During an electrolysis of 1 M $H_2SO_4(aq)$, 2.0 g H_2 was produced at one electrode. The mass of gas produced at the other electrode would be expected to be

A. 8.0 g.
B. 16 g.
C. 32 g.
D. 64 g.

Question 8

Consider the electrochemical cell represented below



If, at 25°C, the voltmeter on this cell shows a reading of 0.94 V, then X is most likely to be

A. copper.
B. tin.
C. iron.
D. zinc.

Question 9

Consider the reaction represented by the equilibrium equation given below

$$2X_2(g) + 3Y_2(g) \rightleftharpoons 2X_2Y_3$$

In a particular investigation, a mixture of 1.0 mol X_2 and 6.0 mol Y_2 was allowed to reach equilibrium.

The amount of X_2Y_3 present at equilibrium could be

A. 0.5 mol.
B. 1.0 mol.
C. 2.0 mol.
D. 4.0 mol.

Question 10

At 200°C, nitrogen oxide reacts with oxygen to form nitrogen dioxide as follows:

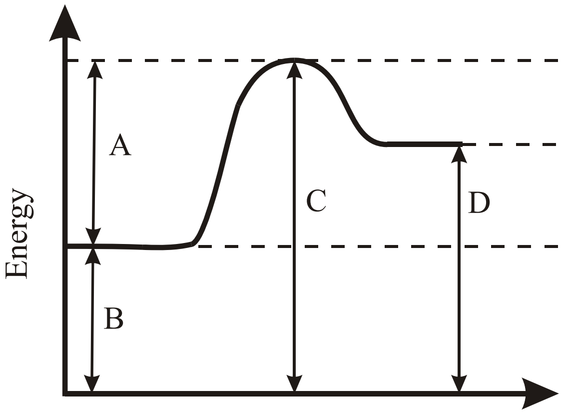
$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g) \quad K_c = 3 \text{ x } 10^6$$

If a one litre mixture of these three gases contains 0.10 mol NO, 0.10 mol NO_2 and 0.010 mol O_2, the reaction is

A. at equilibrium.
B. not at equilibrium and the pressure increases as it moves to equilibrium.
C. not at equilibrium and pressure decreases as it moves to equilibrium.
D. not at equilibrium and the rate of the forward reaction is slower than the rate of the reverse reaction.

Question 11

An energy profile for a reversible chemical reaction is represented below.



According to this profile, the activation energy of the reverse reaction is equal to

A. D.
B. C – B.
C. A – B.
D. C – D.

Question 12

When $Fe^{3+}(aq)$ reacts with $SCN^{-}(aq)$, the complex ion $Fe(NCS)^{2+}(aq)$ is produced as the equilibrium

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons Fe(NCS)^{2+}(aq),\ \Delta H < 0$$

is established.

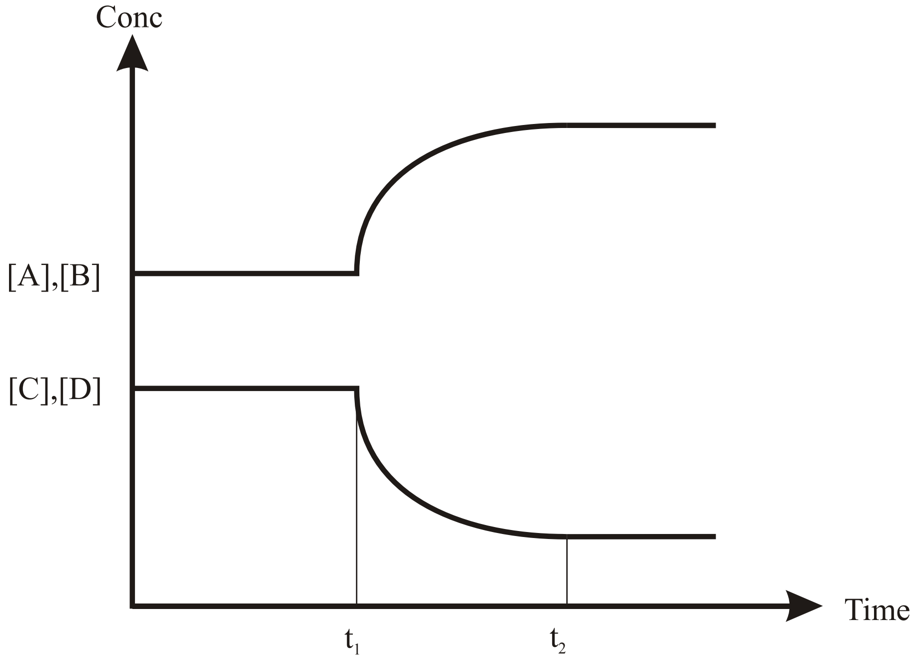
If 50 mL of this equilibrium mixture is added to 50 mL of water and allowed to return to equilibrium at the same temperature, then at the new equilibrium

A. the mass of Fe^{3+} present will have increased.
B. the concentration of SCN^{-} will have increased.
C. the equilibrium constant will be smaller.
D. the total number of particles will have decreased.

Question 13

Consider the general equilibrium system **$A + B \rightleftharpoons C + D,\ \Delta H < 0$**

The concentration-time graph below shows how this equilibrium responded to an imposed change at time **t_1** with equilibrium again being established at time **t_2**



What change was imposed at time t_1?

A. The volume of the equilibrium system was increased.
B. More of reactants A and B were added to the system.
C. The temperature of the system was increased.
D. A catalyst was added to the system.

Question 14

Chemical disinfectants such as chlorine are added to swimming pools to prevent the growth and spread of bacteria and viruses. Chlorine is moderately soluble in water, reacting with it according to the equation

$$Cl_2(g) + 3H_2O(l) \rightleftharpoons Cl^-(aq) + ClO^-(aq) + 2H_3O^+(aq)$$

In an experiment to test the effect of chlorine as a disinfectant, the same amount was added to a 50 L sample of tap water and a 50 L sample of sea-water respectively, both of which were at room temperature.

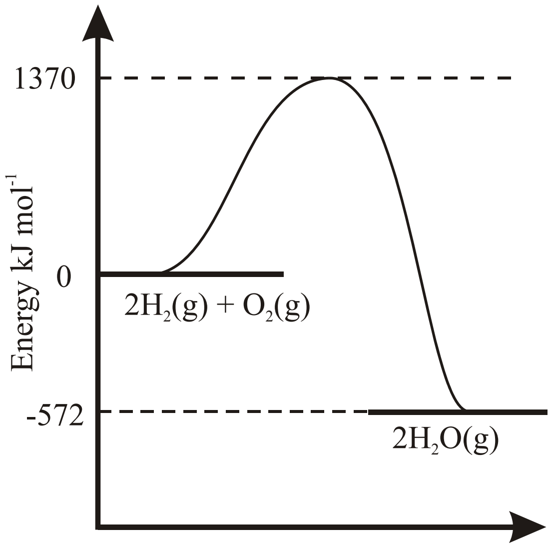
Which of the following outcomes would most likely be observed?

A. The [ClO^-] would be equal in both samples.
B. The sea-water sample would have the higher pH.
C. The [OH^-] would be lower in the sea-water sample.
D. The [Cl^-] would be lower in the sea-water sample.

Question 15

Shown below is the energy profile for the reaction described by the equation:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$



According to the information in this profile, the energy, in kJ, released during the formation of the O-H bonds in one mole of water vapour is

A. 286 kJ
B. 572 kJ
C. 971 kJ
D. 1942 kJ

Question 16

What is the concentration of hydroxide ions, $OH^-(aq)$, in a 0.050 M solution of hypochlorous acid at 25°C?

A. $2.0x10^{-13}$ M
B. $2.6x10^{-10}$M
C. $6.7x10^{-6}$ M
D. $3.8x10^{-5}$ M

Question 17

A current of 1.2 A is passed through 100 mL of 0.50 M $ZnSO_4(aq)$ for 5.0 minutes. The mass of substance produced at the cathode will be

A. 0.12 g
B. 0.24 g
C. 1.6 g
D. 3.3 g

Question 18

Vanadium is a metal that exhibits a wide variety of oxidation states.

Some of its oxidation state changes are represented in the standard half-cell potentials below

$$VO_3^-(aq) + 4H^+(aq) + e^- \rightarrow VO^{2+}(aq) + 2H_2O(l) \qquad E^o = 1.00\ V$$

$$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l) \qquad E^o = 0.32\ V$$

$$V^{3+}(aq) + e^- \rightarrow V^{2+}(aq) \qquad E^o = -0.26\ V$$

The colours associated with the oxidation states of vanadium are:

+2 – purple; +3 – green; +4 – blue; +5 – yellow

If each of the following substances were added, in excess, to separate aqueous solutions of ammonium vanadate, $NH_4VO_3(aq)$, which one would produce a green solution?

A. Zn(s)
B. KI(aq)
C. NaCl(aq)
D. Sn(s)

Question 19

Nickel-hydrogen (Ni-H_2) batteries possess very good electrical properties that have made them attractive for energy storage in satellites and space probes.

Ni-H_2 cells using 26% potassium hydroxide (KOH) as an electrolyte have shown a service life of 15 years or more.

The cathode is made from porous nickel plaque, which contains nickel hydroxide, NiO(OH).

The anode includes a Teflon-bonded platinum black catalyst.

The reaction occurring at the (+) electrode when this battery is delivering energy is

$$NiOOH(s) + H_2O(l) + e^- \rightarrow Ni(OH)_2(s) + OH^-(aq)$$

The reaction at the (-) electrode is most likely to be

A. $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-$
B. $2H^+(aq) + 2e \rightarrow H_2(g)$
C. $H_2(g) \rightarrow 2H^+(aq) + 2e^-$
D. $H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$

Question 20

For the reaction KOH(s) + aq → KOH(aq), ΔH = -55.6 kJ mol^{-1}.
The temperature of water in a calorimeter is measured and then 1.00 g KOH is dissolved in the water. In order to return calorimeter to the temperature it was prior to the addition of the KOH,

A. 991 J of energy must be supplied from outside the calorimeter.
B. 991 kJ of energy must be removed from the calorimeter.
C. 55.6 kJ of energy must be supplied from outside the calorimeter.
D. 55.6 kJ of energy must be removed from the calorimeter.

End of Section A

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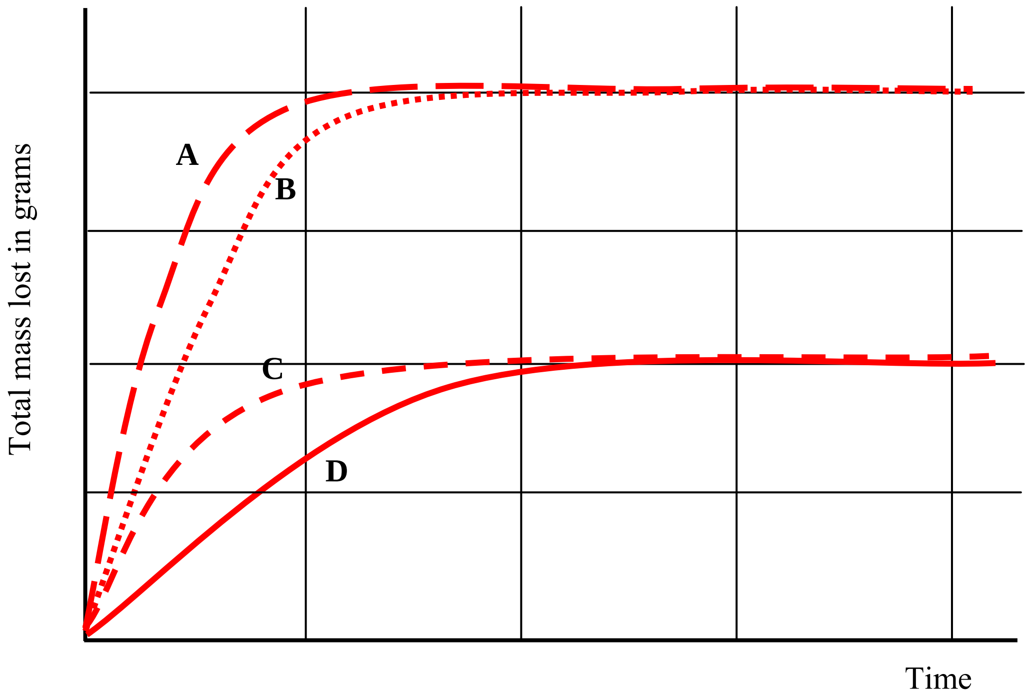
Short Answer Questions – Section B

Section B consists of 7 short answer questions.
You should answer all of these questions.
This section is worth approximately 75 per cent of the total marks available.
The marks allotted are shown at the end of each part of each question.
Questions should be answered in the spaces provided.

Question 1
Following a four part experimental investigation of the effect of concentration and particle size on the rate of the reaction between excess calcium carbonate, in the form of limestone, and hydrochloric acid, a student produces the set of graphs shown below.



a. Write a balanced equation for the reaction between calcium carbonate and hydrochloric acid.

(1 mark)

b. According to the Law of Conservation of Mass, there is 'no loss of mass' during a chemical reaction. How then could the student have used 'total mass loss' in this investigation?

(2 marks)

c. At what stage in the reaction do the graphs level off and become parallel to the horizontal (time) axis?

(1 mark)

d. In each part of the investigation, calcium carbonate was in excess and the student investigated the effect of one of the four possible combinations of limestone lumps, limestone powder, 1 M HCl(aq) and 2 M HCl(aq).
Indicate which of the graphs - A, B, C, D - corresponds to each of the four sets of reaction conditions listed below

1. limestone lumps and 200 mL of 2.0 M HCl(aq) :- _______
2. limestone lumps and 200 mL of 1.0 M HCl(aq) :- _______
3. powdered limestone and 200 mL of 2.0 M HCl(aq) :- _______
4. powdered limestone and 200 mL of 1.0 M HCl(aq) :- _______

4 x ½ = (2 marks)

e. Explain, referring to collision theory, why the rate of the reaction represented by graph A is so much greater than the rate of the reaction represented by Graph D.

(3 marks)

f. Why is the total mass loss shown in graphs A and B eventually twice the total mass loss shown in graphs C and D?

(1 mark)

g. Suggest an alternative method of monitoring the rate of the reaction between calcium carbonate and hydrochloric acid.

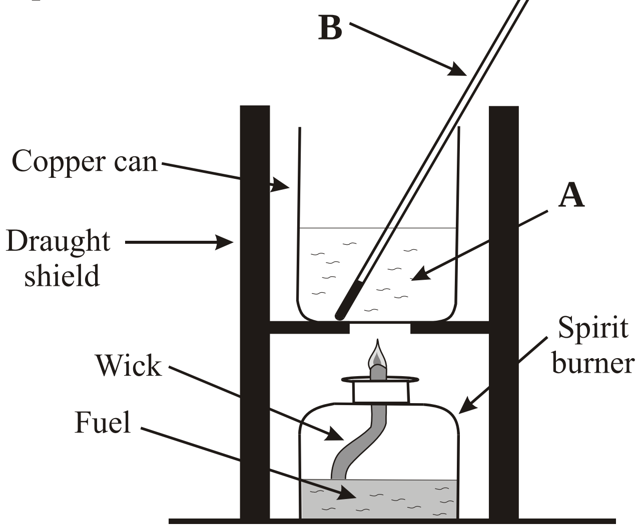
(1 mark)

Total Marks: 11 marks

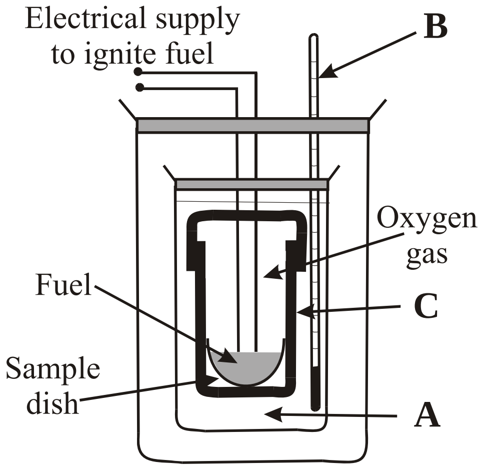
Question 2

Two sets of laboratory apparatus used to determine the molar enthalpy of combustion of a fuel are represented in the diagrams below.

Aparatus 1



Aparatus 2



a. Identify the piece of equipment or chemical substance which is represented by

A: ____

B: ____

C: ____

(3 marks)

b. In an experimental investigation a 5.00×10^{-3} mol sample of a fuel was allowed to react completely in Apparatus 2. A temperature change of 37.5°C was recorded. When 1.52 kJ of electrical energy was added to the apparatus, a temperature change of 2.78°C was recorded.
Work out the molar enthalpy of combustion of fuel and use information from the Data Book to identify the fuel.

(3 marks)

c. When the same fuel was burnt in Apparatus 1 a temperature change of 15.35 °C was recorded in the 200 mL of liquid being heated. The mass of the spirit burner changed from 86.450 g to 86.025 g.
The density of the liquid being heated was 1.00 g mL^{-1}.

i. Calculate the energy absorbed by the 200 mL of liquid.

(1 mark)

ii. Calculate the number of mole of the fuel consumed.

(1 mark)

iii. Calculate the molar enthalpy of combustion of the fuel based on these data.

(1 mark)

d. Give two reasons for the difference in your answers to (b) and (c iii).

(2 marks)

Total 11 marks

Question 3

Methanoic acid is present in the stings of many insects, including bees and ants, which use it as a defence mechanism. During a bite an ant injects 4.0 mg methanoic acid.

a. Write a balanced equation for the ionisation of methanoic acid in aqueous solution.

(1 mark)

b. Write the equilibrium law for the ionisation of methanoic acid.

(1 mark)

c. The same amount of methanoic acid, as injected during an ant bite, was dissolved in water to produce 2.5 mL of solution. Calculate the pH of that solution.

(3 marks)

Total 5 marks

Question 4

Methanol can be produced by direct combination of carbon monoxide and hydrogen, according to the equilibrium.

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g), \quad \Delta H = -90 \text{ kJ mol}^{-1}$$

a. Explain how you can deduce the impact of an increase in temperature of the equilibrium mixture, on the yield of methanol.

(1 mark)

b. The industrial manufacture of methanol using this reaction is usually carried out at around 400 °C in the presence of a Cr_2O_3/ZnO catalyst. What would be the main reasons for using this combination of conditions?

(1 mark)

c. Explain how you can decide how the pressure on the equilibrium system should be altered in order to increase the yield of methanol.

(2 marks)

d. Write a balanced equation for the complete combustion of liquid methanol.

(1 mark)

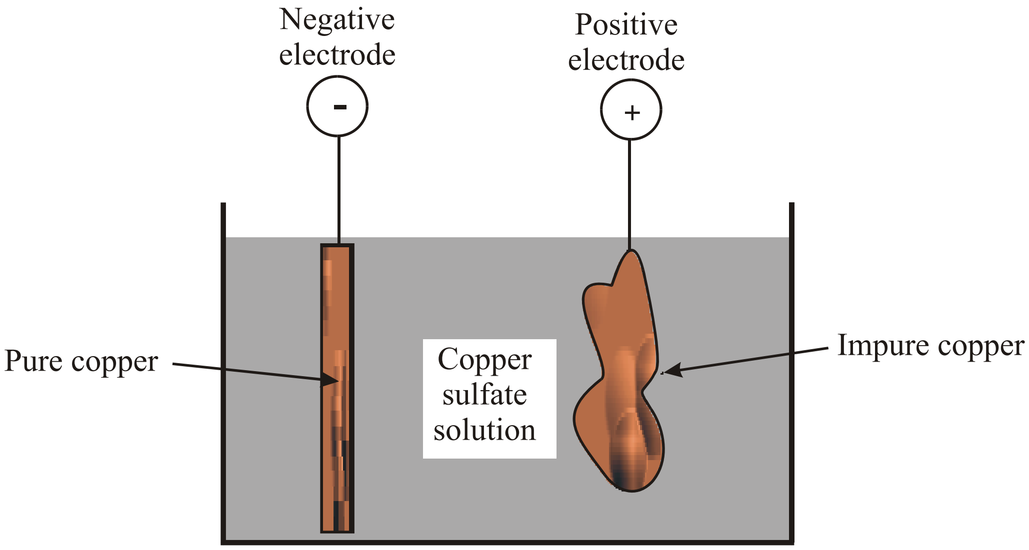
e. Liquid methanol is used in methanol-oxygen fuel cells, using an acid electrolyte. Write a balanced half-equation for the reaction occurring at the (-) electrode in such a methanol-oxygen fuel cell.

(1 mark)

Total 6 marks

Question 5

a. Impurities in a sample of impure copper may include gold and nickel. A sample of impure copper may be purified by controlled electrolysis. The diagram below shows a cell which may be used in such a purification. The impure copper acts as the positive electrode and a piece of pure copper as the negative electrode. An external power supply provides a controlled voltage which ensures transfer of only Cu atoms from the impure copper on to the pure copper sheet.



i. When the external power supply is connected to this cell which terminal of the power supply is connected to the impure copper electrode?

(1 mark)

ii. Write a balanced half-equation for the reaction occurring at the impure copper electrode.

(1 mark)

iii. On the basis of information available in the electrochemical series, what will happen to the gold and nickel impurities present in the impure copper?

(2 marks)

b. Fluorine gas is bubbled through a sample of purified water for a period of time. After some time, the pH of the water is tested and it is found to have decreased. There was no change in temperature of the water.
Explain, using an appropriate equation, why the pH of the water decreases as result of fluorine gas being bubbled through it.

(3 marks)

Total 7 marks

Question 6

The decomposition of hydrogen iodide occurs according to the equilibrium

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g), \qquad \Delta H = 52.1 \text{ kJ mol}^{-1}$$

a. During an investigation of this equilibrium the following data were recorded.

Experiment	Initial amount - mol			Equilibrium amount - mol		
	HI	I_2	H_2	HI	I_2	H_2
1	0.070	0.00	0.00	0.020		
2	0.00	0.025	0.040		0.005	

i. Calculate the value of the equilibrium constant, K, for Experiment 1.

(2 marks)

ii. Calculate the value of the equilibrium constant, K, for Experiment 2.

(2 marks)

iii. In which experiment, 1 or 2, was equilibrium established at the higher temperature? Explain your answer.

(1 mark)

b. During Unit 4 you studied one of the following chemicals in detail.

ammonia nitric acid ethane sulfuric acid

Circle the chemical you studied.

i. Write the chemical formula of the chemical you studied.

(1 mark)

ii. Write a balanced equation for the significant equilibrium reaction associated with the production of this chemical, and indicate whether the equilibrium constant for this equilibrium increases or decreases with temperature.

(2 marks)

iii. State one chemical property of this chemical and give an example of a use of the chemical which depends on this property.

(2 marks)

Total 10 marks

Question 7

An unknown metal M forms a soluble compound, $M(NO_3)_2$.
An aqueous solution of $M(NO_3)_2$ is electrolysed.
When a constant current of 2.50 amperes is applied for 35.0 minutes, 3.06 grams of the metal M is deposited.

a. Calculate the molar mass of M and identify the metal.

(3 marks)

b. On an extended electrochemical series, the metal appears in a half-equation with E^0 = - 40 V.
A galvanic cell is constructed in which this metal is the reductant and the other cell contains Cu^{2+}(aq) ions.

i. State the theoretical voltage of this cell and explain why the voltage generated could be significantly less.

(2 mark)

ii. Give the chemical formula of a compound which could be used in the salt-bridge solution and explain how this compound interacts with half-cell containing the metal, M.

(1 marks)

Total 6 marks

End of Section B

End of Trial Exam

Suggested Answers

VCE Chemistry 2010 Year 12 Trial Exam Unit 4

Multiple Choice Answers – Section A

Q1 C A fundamental principle of equilibrium is that the rates of the forward and reverse reactions are equal.
The forward reaction $CH_3CH_2OH(l) + CH_3COOH(l) \rightarrow CH_3COOCH_2CH_3(l) + H_2O(l)$ – a condensation reaction, proceeds at the same rate as the reverse reaction $CH_3COOCH_2CH_3(l) + H_2O(l) \rightarrow CH_3CH_2OH(l) + CH_3COOH(l)$, an hydrolysis reaction.
There is insufficient data to determine the value of the concentration fraction at equilibrium (or *K*) or the relative concentrations of ethanol and ethyl ethanoate at equilibrium.
The pH of the mixture is most likely to be less than 7 because of the acid catalyst.

Q2 D The electrochemical cell is an electrolysis cell, since it has an external power supply.
The only species present are $Zn^{2+}(l)$ and $Cl^-(l)$
At the anode (+), i.e. electrode Q, the reductant, $Cl^-(l)$ is oxidised according to $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$,
i.e. chloride ions lose electrons at electrode Q.
At the cathode (-), i.e. electrode P, the oxidant, $Zn^{2+}(l)$ is reduced according to $Zn^{2+}(l) + 2e^- \rightarrow Zn(s)$

Q3 A Since *K* is 92 at 250°C and 2.3×10^4 at 30°C, it increases as the temperature decreases. Since *K* is effectively the ratio [products] / [reactants], there are a greater proportion of products at the lower temperature indicating that the forward reaction is exothermic.
Hence $\Delta H < 0$ and the yield of Z increases as the temperature decreases, and decreases as the temperature increases.

Q4 D The concentration-time graph shows that the $[OH^-]$ decreases instantaneously as a result of the imposed change and then increases, but not back to original equilibrium concentration as the system returns to equilibrium.

A. the addition of NaOH will cause an immediate increase in the $[OH^-]$, which will then decrease as the system returns to equilibrium.
B. Since the self-ionisation of water $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ is endothermic, cooling pure water will favour the reverse reaction and so the $[OH^-]$ will decrease but NOT suddenly. It will be a gradual decrease as the temperature drops to the new temperature.
C. Since the self-ionisation of water $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ is endothermic, heating pure water will favour the forward reaction and so the $[OH^-]$ will increase.

D. $NH_3(aq)$ is in equilibrium according to

$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

Doubling the volume by adding an equal volume of water will cause all concentrations to instantaneously halve. So the $[OH^-]$ instantaneously halves. Because H_2O is also a reactant, the reaction moves to the right as the system returns to equilibrium. Hence the $[OH^-]$ increases but is still lower than its concentration before the addition of water when equilibrium is again established.

Q5 C $n(NaOH)$ reacting = $1.0 \times 25\times10^{-3} = 0.025$ mol

$n(HCl)$ reacting = $1.0 \times 25\times10^{-3} = 0.025$ mol

So the energy released in to 50 mL solution from the reaction of 0.025 mol of HCl and 0.025 mol NaOH will cause the temperature to increase by 6°C.

A. 0.050 mol HCl and 0.050 mol NaOH would release twice as much energy, so the temperature of 50 mL of solution would increase by 12 °C.
B. 0.10 mol HCl and 0.10 mol NaOH, i.e. four times as much as the original reaction, will release four times as much energy, so the total solution volume of 100 mL would increase by 12 °C.
C. 0.050 mol HCl and 0.050 mol NaOH would release twice as much energy, so the total solution volume of 100 mL, (i.e. twice the original total volume of 50 mL) would increase by 6 °C.
D. 0.20 mol HCl and 0.20 mol NaOH, i.e. eight times as much as the original reaction, will release eight times as much energy, so the total solution volume of 200 (i.e. four times the original total volume of 50 mL) would increase by 12 °C.

Q6 B The self-ionisation constant of water = 1.0×10^{-14} at 25 °C.

In pure water, $[H_3O^+]$ is always equal to $[OH^-]$ but they are only both 10^{-7} M at 25°C.

The higher K_w at 45°C indicates that the equilibrium

$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

lies further to the right at 45 °C than at 25°C,

i.e. $[H_3O^+] > 10^{-7}$ M and $[OH^-] > 10^{-7}$ M

Consequently at 15°C the equilibrium lies further to the left than at 25°C and so the $[H_3O^+] < 10^{-7}$ M and $[OH^-] < 10^{-7}$ M

Since $[H_3O^+] < 10^{-7}$ M at 25°C, the pH of pure water is > 7.

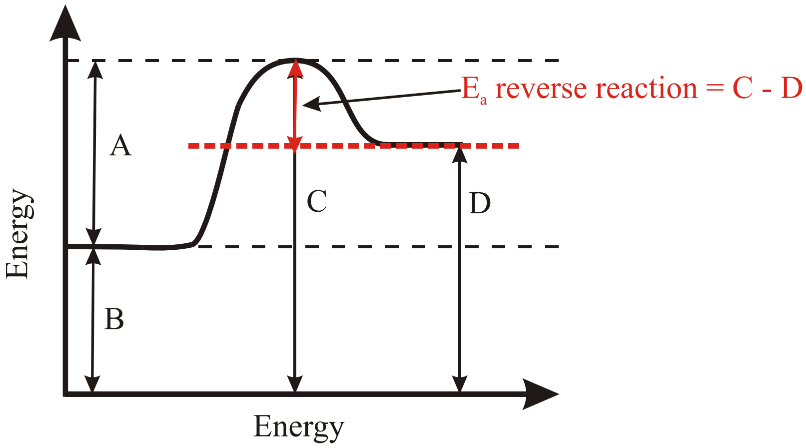
Q7 B According to the electrochemical series, in 1 M $H_2SO_4(aq)$ the strongest oxidant present is $H^+(aq)$ and this is reduced to $H_2(g)$ at the cathode (-) during electrolysis;
$2H^+(aq) + 2e^- \rightarrow H_2(g)$.
The strongest reductant present is $H_2O(l)$, which is oxidised at the anode (+) according to
$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$
$n(H_2)$ produced $= 2.0\text{ g} / 2.0\text{ g mol}^{-1}$
$= 1.0\text{ mol}$
Since according to the half-equations
$n(H_2) = \frac{1}{2} \text{ x } n(e^-)$, and $n(O_2) = {}^1/_4 \text{ x } n(e^-)$, then
$n(O_2) = \frac{1}{2} \text{ x } n(H_2)$
This is also consistent with the overall redox equation $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$
$n(O_2) = \frac{1}{2} \text{ x } 1.0 = 0.50\text{ mol}$
$m(O_2) = 0.50\text{ mol x } 32.0\text{ g mol}^{-1}$
$= 16.0\text{ g}$

Q8 B $E(\text{cell}) = E^0(\text{oxidant half-cell}) - E^0(\text{reductant half-cell})$
So if $X^{2+}(aq) / X(s)$ contains the oxidant
$0.94 = E^0(X^{2+}/X) - E^0(Ag^+/Ag)$
$= E^0(X^{2+}/X) - 0.80$
$E^0(X^{2+}/X) = 1.74\text{ V}$
So if $X^{2+}(aq) / X(s)$ contains the reductant
$0.94 = E^0(Ag^+/Ag) - E^0(X^{2+}/X)$
$= 0.80 - E^0\ (X^{2+}/X)$
$E^0(X^{2+}/X) = -0.14\text{ V}$
There is no half-equation on the electrochemical series with $E^0 = 1.74$, but at $E^0 = -0.14$ is the redox half-equation $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$.
Hence X is most likely to be tin.

Q9 A According to the equilibrium $2X_2(g) + 3Y_2(g) \rightleftharpoons 2X_2Y_3$,
2 mol X_2 reacts with 3 mol Y_2 to produce 2 mol X_2Y_3, so
1 mol X_2 could react with 1.5 mol Y_2 to produce 1 mol X_2Y_3.
However if all 1.0 mol X_2 initially present reacted, the system would not be at equilibrium – all reactants and products must be present if the system is at equilibrium.
So less than 1.0 mol X_2 will have reacted when equilibrium has been achieved and less than 1.0 mol X_2Y_3 will be present.

Q10 C Calculate the reaction quotient (concentration fraction), $Q = [NO_2]^2 / ([NO]^2[O^2])$
$= 0.10^2 / (0.10^2 \text{ x } 0.010)$
$= 100$
Since Q (100) < K_c ($3\text{x}10^6$), the system is not at equilibrium. Since Q must increase to get to the same value as K_c at equilibrium, the forward reaction must dominate as the system moves towards equilibrium. So the rate of the forward reaction will be faster than the rate of the reverse reaction and, since there are fewer particles on the product side, the pressure will decrease.

Q11 D



Q12 A When 50 mL of water is added to 50 mL of the equilibrium mixture, all concentrations instantaneously halve.
The concentration fraction $[Fe(NCS)^{2+}] / ([Fe^{3+}][SCN^-])$ doubles, i.e. becomes greater than the equilibrium constant.
Since the concentration fraction must decrease for the system to get back to equilibrium, the reverse reaction will be favoured. This will

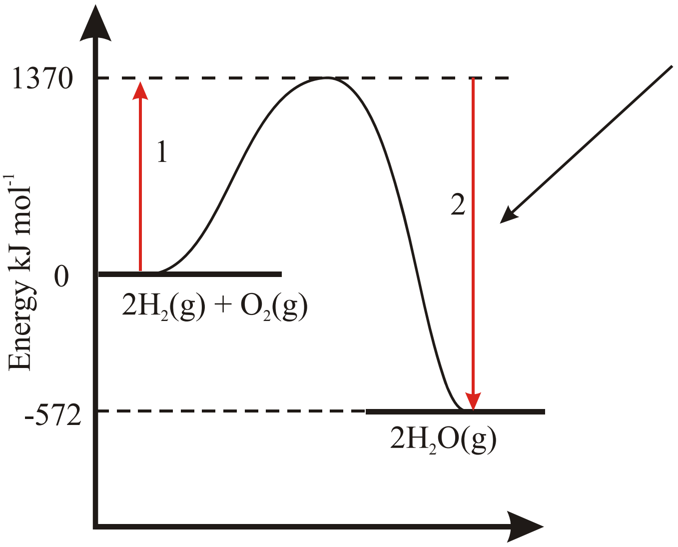
A. increase the $m(Fe^{2+})$
B. increase the $[SCN^-]$ but the concentration at the new equilibrium will be smaller than at the original equilibrium.
C. have no effect on the equilibrium constant since the temperature is the same at both the initial and final equilibria.
D. increase the total number of particles, because there are more particles on the reactant side of the equation.

Q13 C Since all the changes in concentration as a response to the change in conditions are gradual, rather than any of them being instantaneous, the change must have been a change in temperature.
Since the concentrations of the reactants increase as a result of the temperature change, the reverse reaction is favoured. According to the ΔH information, the forward reaction is exothermic, hence the reverse reaction is endothermic and favoured by an increase in temperature.

Q14 B The fundamental difference between sea-water and tap water as far as this equilibrium is concerned was the presence of $Cl^-(aq)$ ions in sea-water.
Thus the sea-water has higher $c(Cl^-)$. This has the effect of pushing the equilibrium in sea-water further to the left, relative to the equilibrium in tap water.
So

A. the $[ClO^-]$ would be lower in sea-water
B. the sea-water would have a lower $[H_3O^+]$ hence a higher pH
C. since sea-water has a lower $[H_3O^+]$ and $[H_3O^+][OH^-]$ is constant, then the $[OH^-]$ will be higher in sea-water.
D. it is as if $Cl^-(aq)$ was added to the tap water equilibrium; the $[Cl^-]$ increases instantaneously and then decreases as the system moves back to equilibrium. However it remains higher at the new equilibrium than at the original equilibrium. So the $[Cl^-]$ will be higher in the sea-water equilibrium.

Q15 C The energy profile shows the energy needed to break the bonds (**1.**) in 2 mol H_2 and 1 mol O_2, i.e. the activation energy of the reaction, and the energy released (**2.**) when bonds are formed in 2 mol H_2O.



Energy released during the formation of O-H bonds in 2 mol H_2O

$E = 370 - (-572)$
$= 1942$ kJ

Hence for the formation of O-H bonds in 1 mol H_2O
$= 1942/2$
$= 971$ kJ

Q16 B Access information for hypochlorous acid form the Data Book
$HOCl(aq) + H_2O(l) \rightleftharpoons ClO^-(aq) + H_3O^+(aq)$; $K_a = 2.9\text{x}10^{-8}$
$[ClO^-][H_3O^+] / [HOCl] = 2.9\text{x}10^{-5}$
Weak acid assumptions:
$[HOCl]_{equilibrium} =$ initial $[HOCl] = 0.050$ M
$[ClO^-] = [H_3O^+]$
So $[H_3O^+]^2 / 0.050 = 2.9\text{x}10^{-8}$
$[H_3O^+]^2 = 0.050 \text{ x } 2.9\text{x}10^{-8}$
$[H_3O^+] = \sqrt{}(0.050 \text{ x } 2.9\text{x}10^{-8})$
$= 3.8\text{x}10^{-5}$ M
At 25°C, $[OH^-] = 10^{-14} / (3.8\text{x}10^{-5})$
$= 2.6\text{x}10^{-10}$ M

Q17 A At the cathode (-) the strongest oxidant present, i.e. $Zn^{2+}(aq)$, will be reduced according to
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
Charge passed $Q = It$
$= 1.2 \text{ x } 5.0 \text{ x } 60$
$= 360$ C
$n(e^-) = Q / F$
$= 360 / 96500$
$= 3.73\text{x}10^{-3}$ mol
$n(Zn) = ½ \text{ x } n(e^-)$
$= ½ \text{ x } 3.73\text{x}10^{-3}$
$= 1.87\text{x}10^{-3}$ mol
$m(Zn) = 1.87\text{x}10^{-3} \text{ x } 65.4$
$= 0.12$ g

Q18 D Ammonium vanadate dissociates in aqueous solution $NH_4VO_3(aq) \rightarrow NH_4^+(aq) + VO_3^-(aq)$

To get the green solution, the vanadate ion, $VO_3^-(aq)$, has to be reduced through to $V^{3+}(aq)$. So the reductant used must be capable of reducing $VO_3^-(aq)$ to $VO^{2+}(aq)$ and then to $V^{3+}(aq)$. But it must not be strong enough the reduce $V^{3+}(aq)$ to $V^{2+}(aq)$.

Including all the species added to the separate solutions of ammonium vanadate, i.e. Zn(s), $K^+(aq)$, $I^-(aq)$, Sn(s), $Na^+(aq)$, $Cl^-(aq)$ in the electrochemical series data.

$Cl_2(g) + 2e^- \rightarrow 2\mathbf{Cl}^-(aq)$ $E^o = 1.36$ V
$VO_3^-(aq) + 4H^+(aq) + e^- \rightarrow VO^{2+}(aq) + 2H_2O(l)$ $E^o = 1.00$ V
$I_2(s) + 2e^- \rightarrow 2\mathbf{I}^-(aq)$ $E^o = 0.54$ V
$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l)$ $E^o = 0.32$ V
$Sn^{2+}(aq) + 2e^- \rightarrow \mathbf{Sn}(s)$ $E^o = -0.14$ V
$V^{3+}(aq) + e^- \rightarrow V^{2+}(aq)$ $E^o = -0.26$ V
$Zn^{2+}(aq) + 2e^- \rightarrow \mathbf{Zn}(s)$ $E^o = -0.76$ V
$\mathbf{Na}^+(aq) + e^- \rightarrow Na(s)$ $E^o = -2.71$ V
$\mathbf{K}^+(aq) + e^- \rightarrow K(s)$ $E^o = -2.93$ V

This shows that of the four reductants added, $Cl^-(aq)$, $I^-(aq)$, Sn(s) and Zn(s), the only one capable of reducing $VO_3^-(aq)$ to $VO^{2+}(aq)$ and then $VO^{2+}(aq)$ to $V^{3+}(aq)$ without reducing $V^{3+}(aq)$ to $V^{2+}(aq)$ is **tin, Sn(s)**.

$VO_3^-(aq) + 4H^+(aq) + e^- \rightarrow VO^{2+}(aq) + 2H_2O(l)$ $E^o = 1.00$ V
$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l)$ $E^o = 0.32$ V
$Sn^{2+}(aq) + 2e^- \rightarrow \mathbf{Sn}(s)$ $E^o = -0.14$ V

Q19 D Since NiOOH is reduced, it must be the oxidant. This suggests that H_2 must be the reductant.

Since H_2 is the reductant, it causes the reduction of NiOOH to $Ni(OH)_2$ and is itself oxidised. Since the electrolyte is alkaline, i.e. contains $OH^-(aq)$, the oxidation half-equation must contain $OH^-(aq)$ ions and may be obtained from the electrochemical series in the Data Book, i.e.

$H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$.

Q20 B *n*(KOH) = 1.00 / 56.1
= 0.0178 mol

Since the reaction is exothermic, energy is released when KOH dissolves.

Energy released = *n*(KOH) x Energy per mole
= 0.0178 x 55.6
= 0.991 kJ

Because 991 kJ of energy is released in the reaction, the temperature inside the calorimeter rises. So to return the calorimeter to its temperature prior to the reaction, 0.991 kJ, i.e. 991 J of energy must be removed from the calorimeter.

Short Answer (Answers) – Section B

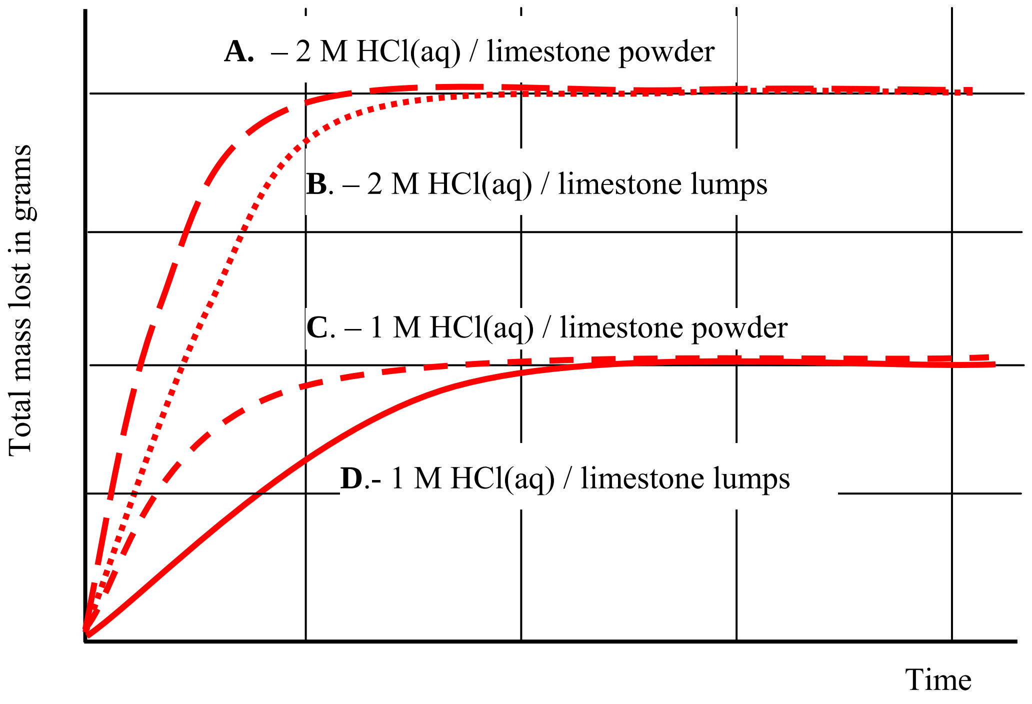
Question 1

a. $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$ ❶ *or*
$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$

b. In the reaction $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$, the law of conservation of mass holds because the total of the masses of $CaCl_2$, H_2O and CO_2 produced, i.e. the products is equal to the total of the masses of $CaCO_3$ and HCl reacting, i.e. the reactants.
However because CO_2 is produced as a gas, if the reaction is carried out in an open flask the CO_2 will escape, i.e. be 'lost' to the atmosphere.❶ Hence monitoring the total decrease in mass of the reaction flask as the reaction proceeds will be equal to the total mass of CO_2 produced.❶

c. Since the $CaCO_3$ is in excess, the total mass loss will become constant when all of the hydrochloric acid has reacted. ❶

d.



1. limestone lumps and 200 mL of 2.0 M HCl(aq) :- ***B***
2. limestone lumps and 200 mL of 1.0 M HCl(aq) :- ***D***
3. powdered limestone and 200 mL of 2.0 M HCl(aq) :- ***A***
4. powdered limestone and 200 mL of 1.0 M HCl(aq) :- ***C*** **❶❶ - 4 x ½**

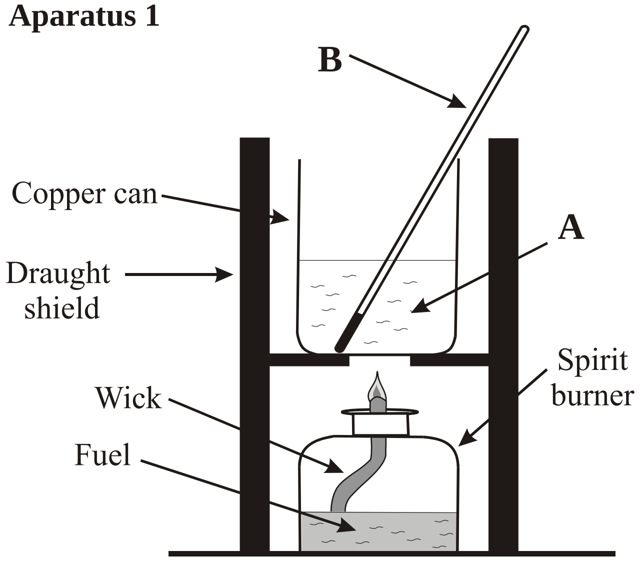
e. Powdered limestone provides a greater surface area for reaction than do the limestone lumps, and the 2 M HCl(aq) provides more particles (H^+ ions) to collide with the $CaCO_3(s)$ than does 1 M HCl(aq).
So in reaction A there is a both a greater surface area of $CaCO_3$ ❶ with which H^+ ions can collide and more H^+ ions ❶ available to collide with the $CaCO_3$. This allows more collisions per second between H^+ ions and $CaCO_3$ collisions, more successful collisions (collisions with energy greater than the activation energy) per second ❶and hence a faster reaction rate.

f. Since the $CaCO_3$ is in excess the $m(CO_2)$ produced depends on the $n(HCl)$. 200 mL of 2.0 M HCl(aq) has twice the $n(HCl)$ as 200 mL of 1.0 M HCl(aq) and so should produce approximately twice as much CO_2.❶

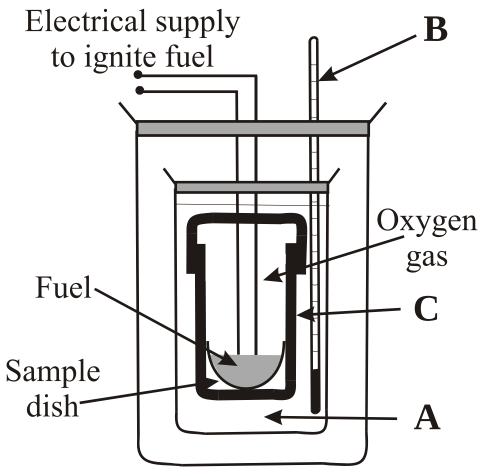
g. Either **measure the volume of CO_2 produced** by collecting it in a gas syringe *or* use a pH meter to **monitor the pH** of the solution (since the $[H^+]$ decreases the pH should increase as the reaction proceeds). ❶

Question 2

Aparatus 1



Aparatus 2



a. A:water / H_2O ❶; B: thermometer ❶; C: reaction bomb. ❶

b. Calibration factor $= 1.52$ kJ / 2.78°C
$= 0.547$ kJ $°C^{-1}$ ❶
Energy released by 5.00×10^{-3} mol fuel $= 0.547$ kJ $°C^{-1}$ x 37.5°C
$= 20.5$ kJ
Energy released by one mole of fuel $= 20.5 / 5.00 \times 10^{-3}$
$= 4.10 \times 10^3$ kJ ❶
$\Delta H_c(\text{fuel}) = -4.10 \times 10^3$ kJ mol^{-1}
According to the Data Book, the fuel is most likely to be hexane, C_6H_{14} ❶

c. i. Energy absorbed by water = SHC (H_2O) x $m(H_2O)$ x ΔT
Since the density is 1.00 g mL^{-1}, $m(H_2O) = 200$ g, $\Delta T = 12.35$ °C = 12.35 K
Energy absorbed by water $= 4.18$ J g^{-1} K^{-1} x 200 g x 15.35 K
$= 1.28 \times 10^4$ J
$= 12.8$ kJ ❶

ii. $m(C_6H_{14})$ reacting $= 86.450 - 86.025$
$= 0.425$ g
$n(C_6H_{14})$ consumed $= 0.425$ g / 86 g mol^{-1}
$= 4.94 \times 10^{-3}$ mol ❶

iii. Energy released by 1 mol $C_6H_{14} = 12.8$ kJ / 4.94×10^{-3} mol
$= 2.59 \times 10^3$ kJ mol^{-1}
$\Delta H_c(C_6H_{14}) = -2.59 \times 10^3$ kJ mol^{-1} ❶

d. Less energy per mole of fuel was calculated for Apparatus 1. Possible reasons include

1. loss of heat released by the burning fuel directly to the atmosphere; a draught shield is not as effective as the insulation in Apparatus 2.
2. loss of heat transferred from the water to the atmosphere through the open top of the copper can.
3. heat released from the burning fuel transferred to components of the system, i.e. stand, copper can, thermometer
4. evaporation of some fuel through the wick of the spirit burner, i.e. not all of the measured amount of fuel may have actually undergone combustion.
5. incomplete combustion of the fuel

❶❶ - two marks for any two valid reasons.

Question 3

a. $HCOOH(aq) + H_2O(l) \rightleftharpoons HCOO^-(aq) + H_3O^+(aq)$ ❶

b. $K = [HCOO^-][H_3O^+][HCOOH]$ ❶ *or*
$[HCOO^-][H_3O^+][HCOOH] = 1.8\text{x}10^{-4}$

c. $n(HCOOH) = 4.0\text{x}10^{-3}\ \text{g}\ / 46.0\ \text{g mol}^{-1}$
$= 8.7\text{x}10^{-5}$ mol
$c(HCOOH) = 8.7\text{x}10^{-5}\ \text{mol} / 2.5\text{x}10^{-3}\ \text{L}$
$= 0.0348$ M ❶

Weak acid assumptions:
$[HCOOH]_{\text{equilibrium}} = [HCOOH]$ initially $= 0.0348$ M
$[HCOO^-] = [H_3O^+]$
$[H_3O^+]^2 / [HCOOH] = 1.8\text{x}10^{-4}$
$[H_3O^+]^2 / 0.0348 = 1.8\text{x}10^{-4}$
$[H_3O^+] = \sqrt{(0.0348 \text{ x } 1.8\text{x}10^{-4})}$
$= 2.50\text{x}10^{-3}$ M ❶
$pH = -\log_{10}(2.50\text{x}10^{-3})$
$= 2.60$ ❶

Question 4

a. Since ΔH (-90 kJ mol^{-1}) is < 0, the forward reaction is exothermic and would be favoured at lower temperatures. Hence for this equilibrium, the yield would be greater at lower temperatures and smaller at higher temperatures. ❶

b. Because the forward reaction is exothermic, the yield of methanol is better at lower temperatures. However lower temperature means a slower rate of reaction. So the use of moderately low temperature (400°C) and a catalyst provides a compromise to this conflict between rate and yield and allows for both a good yield and a high rate of reaction. ❶

c. The pressure in a gaseous equilibrium system depends on the total number of particles present. According to the equation $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$, the production of methanol reduces the total number of particles and so reduces the pressure. ❶ According to Le Chatelier's principle, the system will move to decrease the pressure, and increase the yield of CH_3OH, if it is subjected to a pressure increase. ❶

d. $2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$ ❶ *or*
$CH_3OH(l) + 1.5O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

e. $CH_3OH(l)$ is oxidised to $CO_2(g)$ at the (-) electrode (anode) in a fuel cell
$CH_3OH(l) + H_2O(l) \rightarrow CO_2(g) + 6H^+(aq) + 6e^-$ ❶

Question 5

a. i. The positive (electrode) ❶. In electrolysis or cell recharging the (+) terminal of the power supply is always connected to the (+) electrode.

ii. $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ ❶

iii. The locations of Cu, Ni and Au on the electrochemical series provide the clues. Since Cu must be oxidised at the anode, (+) electrode, Ni(s) will also be oxidised, to $Ni^{2+}(aq)$ because it is a stronger reductant than Cu.
However because $Cu^{2+}(aq)$ is a stronger oxidant, $Ni^{2+}(aq)$ will not be reduced at the cathode, (-) electrode, and will remain in the electrolyte solution. ❶
Since Cu is a much stronger reductant than Au, the voltage applied to ensure oxidation of Cu in the impure copper will not be enough to oxidise the Au. Hence over time the Au atoms will fall to the bottom of the cell and collect under the impure copper electrode. ❶

b. According to the electrochemical series, $F_2(g)$ will react with $H_2O(l)$ since the oxidant, F_2, is higher than the reductant, H_2O. ❶
The overall equation for the reaction can be established from the half-equations
Reduction: $F_2(g) + 2e^- \rightarrow 2F^-(aq)$
Oxidation: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$
Overall equation: $2F_2(g) + 2H_2O(l) \rightarrow 4F^-(aq) + 4H^+(aq) + O_2(g)$ ❶
The pH of the water decreases because of the increase in concentration of H^+ during the reaction. ❶

Question 6

a. i. Since the amount of HI present is lower at equilibrium, the forward reaction dominates as the reaction moves to equilibrium.

	$2HI(g)$	$\rightleftharpoons$	$H_2(g)$	$+ I_2(g)$		
Initially	0.070		-	-	mol	
Reacting	*0.050*	→	*0.025*	*0.025*	*mol*	
Equilibrium	0.020		**0.025**	**0.025**	mol	❶

Since there are equal numbers of particles on both sides of the equation, the volume of the container is not needed to work out the value of *K*. You can assume a volume of one litre

$K = [H_2][I_2] / [HI]^2$
$= 0.025 \times 0.025 / 0.020^2$
$= 1.6$ ❶

ii. Since the amount of I_2 present is lower at equilibrium, the reverse reaction dominates as the reaction moves to equilibrium.

	$2HI(g)$	$\rightleftharpoons$	$H_2(g)$	$+ I_2(g)$		
Initially	-		0.040	0.025	mol	
Reacting	*0.040*	←	*0.020*	*0.020*	*mol*	
Equilibrium	**0.040**		**0.020**	0.005	mol	❶

Since there are equal numbers of particles on both sides of the equation, the volume of the container is not needed to work out the value of *K*. You can assume a volume of one litre

$K = [H_2][I_2] / [HI]^2$
$= 0.020 \times 0.005 / 0.040^2$
$= 0.06$ ❶

iii. The forward reaction is endothermic, and so is favoured by higher temperature. So the yield of products, and the value of *K* will increase as the temperature increases. Hence Experiment 1, with the higher *K* value, was at the higher temperature at equilibrium. ❶

b.

i. ❶	**NH_3**	**HNO_3**	**C_2H_4**	**H_2SO_4**
ii. ❶ ❶	**$N_2 + 3H_2 \rightleftharpoons 2NH_3$** Exothermic ***K* decreases**	**$2NO + O_2 \rightleftharpoons 2NO_2$** Exothermic ***K* decreases**	**$C_3H_8 \rightleftharpoons C_2H_4 + CH_4$** Endothermic ***K* increases**	**$2SO_2 + O_2 \rightleftharpoons 2SO_3$** Exothermic ***K* decreases**
iii. ❶ ❶	e.g. Reductant Nitric acid production e.g. Base Fertiliser production	e.g. Acid Fertiliser production e.g. Oxidant Production of NO_2	e.g. Unsaturated hydrocarbon Addition polymer production.	e.g. Dehydrating agent Drying gases. e.g. Acid Superphosphate production e.g. Oxidant Reactions with metals

Question 7

a. $Q = It$
$= 2.50 \text{ x } 35.0\text{x}60$
$= 5.25\text{x}10^3$ C
$n(e^-) = Q / F$
$= 5.25\text{x}10^3 / 96500$
$= 0.0544$ mol ❶
Reduction half-equation: $M^{2+}(aq) + 2e^- \rightarrow M(s)$
$n(M) = ½ \text{ x } n(e^-)$
$= ½ \text{ x } 0.0544$
$= 0.0272$ mol ❶
Molar mass M $= m(M) / n(M)$
$= 3.06 / 0.0272$
$= 112.5$ g mol^{-1}
Element is cadmium, Cd ❶

b. i. $E(\text{cell}) = E^0(Cu^{2+}/Cu) – E^0(Cd^{2+}/Cd)$
$= 0.34 – (-0.40)$
$= 0.74$ V ❶
If the solutions in the half-cells are less than 1 M, e.g. 0.10 M $Cu^{2+}(aq)$ and 0.10 M $Cd^{2+}(aq)$, the generated voltage would be less than 0.74 V. ❶

ii. KNO_3
$NO_3^-(aq)$ ions flow from the salt bridge to counteract the build up of positive charge in the solution due to the oxidation half-equation ($Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$), and so maintain the half-cell solution charge neutrality essential for effective operation of the galvanic cell. ❶

End of Suggested Answers