GENERAL COMMENTS

Results from the 2011 Chemistry 2 examination indicated that the exam was slightly more challenging than examination 1.

Students should be reminded that the examination is designed to assess their ability to apply their knowledge and understanding in familiar and new contexts, and in most questions this requires more than a recall of facts. While past and practice examinations are a valuable part of examination preparation, skill in question interpretation should be developed continually. Attempting questions from different sources throughout the semester to enhance the range of contexts to which students are exposed is recommended.

SPECIFIC INFORMATION

Performance on Section A of the examination was very strong, with only six questions (Questions 5, 8, 15, 16, 18 and 20) eliciting less than 60 per cent correct responses. In Question 5 the accurate measurement of the Fe^{3+}(aq) concentration required its concentration to be as low as possible at the point of analysis. Question 8 required a unit conversion, but this was missed by many. Questions 15 and 16 assessed students' ability to apply the supplied electrochemical data to the recharging of a secondary cell. Questions associated with the electrochemical series proved challenging throughout this exam. The fundamental principles of recharging – that electrons move from the (+) electrode to the (−) electrode, and that (+) is connected to (+) – were not well applied in Question 16. In Question 18 many students chose alternative C, which is consistent with a mole ratio error that occurs far less frequently when such a question appears in Section B. In Question 20, many students missed either the link between K_a and $[H_3O^+]$ or between pH and $[H_3O^+]$.

In reviewing performance on Section B, the designated key skill that requires students be able to 'apply chemical understandings: make connections between concepts, process information, apply understandings to familiar and new contexts' should be kept in mind. The most challenging component of this on examination 2 was 'new contexts'. Students should not expect that all questions on an examination will deal with contexts they have previously encountered in questions from textbooks and past or practice examinations. Practice reading questions carefully and identifying the underlying key knowledge are important components of examination preparation and everyday progress through VCE Chemistry. In Question 1b. many students struggled to produce a balanced chemical equation for the formation of a chemical produced from their 'selected' industrial chemical.

Question 2aii. required students to 'describe the main difference between a renewable and a non-renewable energy source'; however, it was evident from responses that many were confusing the terms 'renewable', 'recyclable' and 'reusable'. Question 2bii. clearly stated that 'oxygen is converted to oxide ions'. Most students were not awarded the mark for writing the associated half-equation and this can be attributed to incorrect identification of the charge on the oxide ion and/or inaccurate balancing. This issue continued in Questions 4b., 8a. and 8d.

Question 3 required students to assimilate data effectively. In Question 3a., most students did not allow for the fact that while the units of the calibration factor were kJ mol^{-1}, electrical energy (E) in the relationship $Q = E/V$ has the unit J. In Question 3bii. the data provided was either ignored or only partly used in many calculations and this reduced overall performance.

Question 4a. required students to integrate information from the electrochemical series in the Data Book with the information supplied in the question. Students were asked to explain why the given reaction would not be expected to occur. Those who considered the expected 'alternative reaction' option, predicted by effective use of the electrochemical series, generally identified that H_2O(l) at + 1.23 V was a stronger reductant, but did not pick up on the even stronger reductant, Br$^-$(aq) at + 1.09 V. Students should be encouraged to identify all locations of a particular species on the electrochemical series before drawing inferences about the likely extent of a reaction. This question required an explanation of 'why the given reaction is not predicted to occur', but many students simply gave a general explanation of 'why a reaction predicted to occur may not be seen to occur' and did not make effective use of either the supplied information or the extra relevant information available from the electrochemical series. Responses to Question 4b. suggested that students need more practice at deducing half-equations from unfamiliar overall redox equations.

It was evident that students found Question 5 challenging. Many students seemed confused by Question 5a. and struggled to explain the changes in the rates of the forward and reverse reactions as a reacting system moves to equilibrium. Many overlooked the most fundamental principle – that forward and reverse reaction rates are equal at

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equilibrium. The issue of effective question interpretation was evident in Question 5bi., where the link between high temperatures in combustion engines and the information supplied earlier in the question that the production of NO is endothermic was missed by many. In conjunction with overall performance on Question 3bii., this suggests that some students may limit the data or information they consider in framing a response to that which is seen in the part of the question being answered.

Responses to Question 6a. indicated the need to reinforce the requirement to include equilibrium arrows in 'known' equilibrium reactions, such as that between methanoic acid and water. Question 6bi. showed that the tendency to go with the familiar can sometimes be a little too automatic. Most students, well trained in the 'standard' assumptions associated with weak acid calculations, ignored one of the 'different' assumptions specified in the question and defaulted to well-learned techniques. In Question 6biii. many students struggled to link their explanation to information in, and responses given to, earlier parts of the question.

Responses to Question 7 demonstrated both strengths and deficiencies in depth of explanation. In Question 7a., students were asked to explain why the reaction moved in the forward direction when high pressures are used. Many answers simply used the phrase 'according to Le Chatelier's principle', rather than the more acceptable 'According to Le Chatelier's principle, the reaction moves to oppose the change/decrease the pressure.' However, in Question 7b. it was pleasing to see that the majority of students provided in-depth explanations beyond 'lowers the activation energy'.

Despite the specific mention of copper (I) in Question 8a., most students included Cu^{2+} in their half-equations. Responses to Questions 8b. and 8c. tended to overlook the 'equilibrium' information supplied in the question, despite a clear instruction in 8b. to refer to this information. In general, performance on Question 8d. was poor. Students did not seem to use the electrochemical series to identify a species present in the cell that could also be reduced. Question 8e. was well done and showed that students continue to handle calculation questions associated with Faraday's laws very competently.

Section A – Multiple-choice questions

The table below indicates the percentage of students who chose each option. The correct answer is indicated by shading.

Question	% A	% B	% C	% D	% No Answer	Comments
1	92	1	2	5	0	
2	71	4	19	6	0	The addition of a catalyst has no effect on the extent of reaction, hence the ratio [Z]/[X] is unaffected. If a system is at equilibrium, the addition of a catalyst will increase the rates of the forward and reverse reactions but will have no effect on the position of equilibrium.
3	9	10	5	75	0	Since there is the same number of mole of particles on both sides, the volume of the container is not a factor in the determination of the equilibrium constant or the concentration fraction (reaction quotient). $K = [HF]/\{[H_2]^{1/2}[F_2]^{1/2}$ $= (n(HF)/V)/\{(n(H_2)/V)^{1/2} \times \{(n(F_2)/V)^{1/2}\}$ $= (n(HF)/V)/\{(n(H_2)^{1/2}/V^{1/2}) \times \{(n(F_2)^{1/2}/V^{1/2})\}$ $= (n(HF)/V)/\{(n(H_2)^{1/2} \times (n(F_2)^{1/2}/V\} = n(HF)/\{n(H_2)^{1/2} \times (n(F_2)^{1/2}\}$
4	8	60	27	4	1	According to the information supplied with the equation for the reaction, the production of one mol HF releases 271 kJ of energy. So the energy released during the production of 0.220 mol HF is $0.220 \times 271 = 59.6$ kJ.
5	21	10	19	49	0	The most accurate determination of the concentration of Fe^{3+} ions happens when the equilibrium $Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{2+}(aq)$ is pushed as far to the right as possible to maximise the conversion of $Fe^{3+}(aq)$ to $FeSCN^{2+}(aq)$. A large excess of $SCN^-(aq)$ and a large equilibrium constant will favour increased conversion to $FeSCN^{2+}$. Since the forward reaction is exothermic, the analysis should also be carried out at low temperature.

Question	% A	% B	% C	% D	% No Answer	Comments
6	4	2	75	19	0	The energy profile below shows that the activation energy, E_a is greater than the enthalpy of the reaction, ΔH.
7	16	76	5	3	0	To get C(diamond) on the product side, combine the equations $C(graphite) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393$ kJ mol^{-1} and $CO_2(g) \rightarrow C(diamond) + O_2(g)$ $\Delta H = +395$ kJ mol^{-1} to give $C(graphite) \rightarrow C(diamond)$ $\Delta H = +2$ kJ mol^{-1} ΔH for $C(graphite) \rightarrow C(diamond)$ can also be deduced from the supplied energy profile.
8	30	14	11	44	0	Energy added to water $= 4.18$ J g^{-1} K$^{-1} \times m(H_2O) \times \Delta T$ $= 4.18 \times 100.0 \times 1.00$ $= 418$ J $\Delta H_c(C_4H_{10}) = -2874$ kJ mol^{-1} $n(C_4H_{10})$ required $=$ energy added/energy available from 1 mol C_4H_{10} $= 41810^{-3}$ kJ/2874 kJ mol^{-1} $= 1.45 \times 10^{-4}$ mol $m(C_4H_{10})$ required $= 1.45 \times 10^{-3}$ mol x 58.0 g mol^{-1} $= 8.44 \times 10^{-3}$ g Alternative A was consistent with non-conversion of J to kJ or vice versa when working out $n(C_4H_{10})$.

Question	% A	% B	% C	% D	% No Answer	Comments
9	10	10	18	**62**	0	Energy added to water $= 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times m(\text{H}_2\text{O}) \times \Delta T$ $= 4.18 \times 100.0 \times 40.00$ $= 1.67 \times 10^4 \text{ J}$ Energy released from 1 mol ethanol $= 1.67 \times 10^4 \text{ J}/0.0200 \text{ mol}$ $= 8.36 \times 10^5 \text{ J mol}^{-1}$ $= 836 \text{ kJ mol}^{-1}$ $\Delta H_c(\text{ethanol}) = -836 \text{ kJ mol}^{-1}$
10	9	15	**67**	7	0	In galvanic cells, electrons flow from the (−) electrode to the (+) electrode. Since electrons always move from the site of oxidation to the site of reduction, the stronger oxidant – the species that is reduced – must be in the half-cell containing the (+) electrode. In the galvanic cell on the left, $\text{Ag(NH}_3)_2^+(\text{aq})$ is the strongest oxidant – stronger than $\text{Ag(CN)}_2^-(\text{aq})$. In the galvanic cell on the right, $\text{Ag}^+(\text{aq})$ is the stronger oxidant – stronger than $\text{Ag(NH}_3)_2^+(\text{aq})$. Hence, the correct order of the reduction half-equations in order of decreasing electrode potentials is $\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag(s)} \qquad E^\circ_2$ $\text{Ag(NH}_3)_2^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag(s)} + 2\text{NH}_3(\text{aq}) \qquad E^\circ_1$ $\text{Ag(CN)}_2^-(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag(s)} + 2\text{CN}^-(\text{aq}) \qquad E^\circ_3$ Students should be aware that each half-cell contains an oxidant and a reductant, and that electrons move from the half-cell containing the weaker oxidant to the half-cell containing the stronger oxidant. Also in a half-cell, the species with the element at the higher oxidation number is the oxidant.
11	8	4	**80**	8	0	As it was a galvanic cell, the reactants were the strongest oxidant, $\text{Ag}^+(\text{aq})$, and the strongest reductant, Cu(s) – identified from the electrochemical series, with the equation balanced for charge.
12	**74**	16	7	3	0	The predicted voltage, under standard conditions as $E = E^\circ(\text{half-cell containing the oxidant}) - E^\circ(\text{half-cell containing the reductant})$ $= 0.80 - 0.34$ $= 0.46 \text{ V}$
13	**67**	6	9	18	0	Discharging is a spontaneous reaction where the oxidant is in the half-equation with the higher E° value. When the half-equations are arranged in order of decreasing E° values, i.e. $\text{PbO}_2(\text{s}) + 4\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O(l)} \quad E^\circ = +1.69 \text{ V}$ $\text{PbSO}_4(\text{s}) + 2\text{e}^- \rightleftharpoons \text{Pb(s)} + \text{SO}_4^{2-}(\text{aq}) \qquad E^\circ = -0.36 \text{ V}$ it can be deduced that the reduction of strongest oxidant, $\text{PbO}_2(\text{s})$, is accompanied by a decrease in $[\text{H}^+]$ and so the pH increases.
14	10	3	**85**	2	0	During recharging, electrical energy is converted into chemical energy and the passage of electrical current through the external circuit also generates some heat energy.
15	17	26	**49**	8	0	During recharging, the direction of electron flow is reversed (compared to discharging) and electrons are 'forced' to flow from the (+) electrode to the (−) electrode. Since electrons always leave the anode (site of oxidation), then the (+) electrode is the anode during recharging. The half-reaction occurring at the (+) electrode during discharging was $\text{PbO}_2(\text{s}) + 4\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{e}^- \rightarrow \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O(l)}$. So the half-reaction occurring at the (+) electrode during recharging is $\text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O(l)} \rightarrow \text{PbO}_2(\text{s}) + 4\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{e}^-$.

Question	% A	% B	% C	% D	% No Answer	Comments
16	56	20	12	13	0	During recharging, electrons are 'forced' to flow from the (+) electrode to the (−) electrode. Since electrons always leave the site of oxidation, then the (+) electrode is the site of oxidation during recharging. When recharging a battery, the (+) terminal of the power supply is connected to the (+) electrode of the battery. Students should be aware of the fundamental difference between discharging and recharging in terms of the direction of electron flow and that electrons always move from the site of oxidation to the site of reduction. The (+) to (+) connection to the power supply requirement for recharging is a key principle of electrolysis.
17	6	3	14	77	0	Oxidation occurs at the anode and reduction occurs at the cathode in all electrochemical cells.
18	11	53	24	11	1	Reduction half-equations occurring at the (−) electrode were $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$ $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $Ag^+(aq) + e^- \rightarrow Ag(s)$ 0.030 Faraday implies $n(e^-)$ through each cell was 0/030 mol $n(Cr)$ deposited $= n(e^-)/3 = 0.030/3 = 0.010$ mol $n(Cu)$ deposited $= n(e^-)/2 = 0.030/2 = 0.015$ mol $n(Ag)$ deposited $= n(e^-) = 0.030$ mol
19	6	78	12	4	0	Reduction half-equation $M^{x+} + xe^- \rightarrow M$ The value of x is determined from the ratio $n(e^-)/n(M)$ $n(e^-) = Q/F = I \times t/F$ $= 1.50 \times 180/96500$ $= 2.8 \times 10^{-3}$ mol $n(e^-)/n(M) = 0.028/0.014$ $= 2$ Hence $x = 2$
20	27	11	10	52	0	With no other information supplied, the Data Book should have been an immediate reference for students. All the acids are weak acids. The acid with the highest pH will have the lowest $[H_3O^+]$. Since all acid solutions were 0.1 M, the acid with the lowest $[H_3O^+]$ is the one that ionises least in aqueous solution. This will be the acid with the lowest K_a, i.e. hypobromous acid.

Section B – Short answer questions

Question 1
1ai.

Marks	0	1	Average
%	6	94	**1**

There was a range of possible answers for each chemical. Correct responses included:
- NH_3 – irritating to the respiratory system, irritating to eyes
- C_2H_4 – flammable, toxic by inhalation
- HNO_3 – causes burns, risk of serious damage to the eyes
- H_2SO_4 – toxic if swallowed, irritating to skin.

Most students were able to identify a risk phrase applicable to their selected industrial chemical. Some students either did not identify the chemical as instructed or circled more than one, including a phrase that was not appropriate; for example, 'flammable' for sulfuric acid.

1aii.

Marks	0	1	Average
%	22	78	**0.8**

Appropriate safety phrases for the risk phrases given included:

- flammable – keep away from flames
- toxic by inhalation – wear breathing apparatus
- toxic in contact with skin – wear protective clothing and gloves
- toxic if swallowed – do not ingest
- causes burns – wear protective clothing and gloves, wear face and eye protection
- irritating to respiratory system – use in a well-ventilated area
- irritating to skin – wear protective clothing and gloves
- risk of serious damage to eyes – wear safety goggles
- vapours may cause drowsiness and dizziness – use in a well-ventilated area.

While most students handled this question well, a significant proportion of responses were either superficial or inappropriate. Some students seemed to consider that 'wear a lab coat' was appropriate for any risk. Such a limited response was insufficient for risks such as 'toxic in contact with skin' or 'causes burns' and an illogical response for risks such as 'toxic by inhalation' or 'irritating to eyes'. Also of concern was the frequency of 'wear a mask' as a response for 'toxic by inhalation'. A small number of students simply named a piece of safety equipment, for example gloves, but did not indicate how it should be used.

1bi.

Marks	0	1	Average
%	27	73	**0.8**

Acceptable responses included:

- NH_3 – ammonium nitrate or NH_4NO_3; ammonium sulfate or $(NH_4)_2SO_4$; urea or $(NH_2)_2CO$
- C_2H_4 – ethanol or C_2H_6O or CH_3CH_2OH; chloroethane or C_2H_5Cl or CH_3CH_2Cl; polyethylene
- HNO_3 – ammonium nitrate or NH_4NO_3; potassium nitrate or KNO_3
- H_2SO_4 – ammonium sulfate or $(NH_4)_2SO_4$; oleum, $H_2S_2O_7$; superphosphate.

Some students seemed to misinterpret this question and focused on their selected chemical as a product. A large number of students who selected sulfuric acid, ammonia or nitric acide simply used the word 'fertiliser'. Since the question asked for 'one useful industrial chemical', the name or formula of the chemical was required for full marks. However, if the fertiliser was clearly identified in part bii. then the student was awarded the mark for part bi.

1bii.

Marks	0	1	2	Average
%	45	10	46	**1**

The equation had to be balanced and include the 'selected' chemical as a reactant. Acceptable equations included:

- $NH_3(g) + HNO_3(aq) \rightarrow NH_4NO_3(aq)$: for the production of ammonium nitrate from ammonia or nitric acid
- $C_2H_4(g) + H_2O(g) \rightarrow C_2H_6O(g)$: for the production of ethanol from ethene
- $NH_3(g) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$: for the production of ammonium sulfate from sulfuric acid
- $SO_3(g) + H_2SO_4(l) \rightarrow H_2S_2O_7(l)$: for the production of oleum from sulfuric acid
- $nC_2H_4 \rightarrow -(-C_2H_4-)_n-$: for the production of polyethylene from ethene.

Many students were unable to produce a realistic equation involving their selected chemical as a reactant. In many cases the student's intention was evident, but the use of an incorrect formula such as NH_4SO_4 made their equation difficult to balance or led to an unrealistic equation.

Given that students would have been expected to have studied the 'uses' of their selected chemical it was surprising that more students did not provide an appropriate balanced equation.

A small number of students who selected sulfuric acid provided or attempted to provide an equation for the production of the fertiliser superphosphate. While it is admirable that students can successfully remember the rather complex equation, there is a far simpler option as far as providing a balanced equation is concerned, in the production of the fertiliser ammonium sulfate.

Question 2
2ai–ii.

Marks	0	1	2	Average
%	19	35	46	**1.3**

2ai.
Reaction 1 – coal
Reaction 2 – crude oil or petroleum
Reaction 3 – yellowcake or uranium ore
Reaction 4 – biomass, plants, animal fat or vegetable oil

This question proved challenging for some students. The four reactions represented the combustion of carbon, the combustion of octane, nuclear fission of uranium-235 and the combustion of the methyl ester of a fatty acid (biodiesel) respectively. However, some students either could not describe an energy source for one of the fuels or identified an energy source not appropriate for the fuel.

2aii.
Renewable energy resources are able to be replenished in a short time or can be replaced by natural processes.
Non-renewable energy resources are used at a faster rate than they can be replenished or are finite.
Part ii. revealed a serious misunderstanding of what constitutes a renewable energy resource, with many students simply referring to recycling or reusing energy sources.

2aiii.

Marks	0	1	2	Average
%	16	15	69	**1.6**

Acceptable responses included:
- electrical energy used to power electrical equipment, provide lighting, etc. – this could be applied to all the fuel sources, including octane from crude oil, which can be used in generators
- mechanical energy used to power motor vehicles – for crude oil and biomass fuel sources
- mechanical energy which is converted to electrical energy – for coal and uranium fuel sources.

2bi.

Marks	0	1	Average
%	32	68	**0.7**

Acceptable responses included:
- fuel cells require a continuous supply of reactants, whereas in a galvanic cell they are stored in the cell
- some products are removed from fuel cells
- fuel cell electrodes are porous.

While generally strong, performance on this question was perhaps a little below expectation. Differences in the principles of operation of fuel cells compared to other galvanic cells are distinct. Students need to be careful in the use of the term 'rechargeable' when dealing with electrochemical cells. It was only acceptable in this context if it was clear that it was being used in reference to maintaining the supply of reactants for the fuel cell.

Question 2bii.

| Marks	0	1	Average
%	63	37	**0.4**

Either of:

- $O_2(g) + 4e^- \rightarrow 2O^{2-}(g)$
- $\frac{1}{2}O_2(g) + 2e^- \rightarrow O^{2-}(g)$.

Common errors with this half-equation included the incorrect charge on the oxide ion, different numbers of oxygen atoms on each side and charge imbalance by having the incorrect number of electrons. $O_2 + 2e^- \rightarrow O^{2-}$ was a common incorrect half-equation, as was $O_2 + 2e^- \rightarrow 2O^-$.

2biii.

Marks	0	1	Average
%	68	32	**0.3**

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
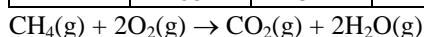
Students who provided an incorrect half-equation in part ii. were generally unable to provide a correct equation for the the overall reaction. To obtain the mark for a correctly balanced equation all states had to be (g).

Question 3
3a.

Marks	0	1	2	Average
%	21	42	37	**1.2**

E added to calorimeter
$$\begin{aligned} &= CF \times \Delta H \\ &= 0.354 \times 9.50 \\ &= 3.36 \text{ kJ} \end{aligned}$$
Electrical energy $E = VIt$, and electric charge $Q = It$. Therefore, $E = QV$ so
$$\begin{aligned} Q &= E/V \\ &= 3.36 \times 10^3 \text{ (J)}/5.10 \\ &= 659 \text{ (C)} \end{aligned}$$

The most common answer to this question was 0.659. This indicated that most students did not recall that in the relationship $Q = E/V$, electrical energy is in joules and did not convert kJ to J. A significant number of students divided the temperature change by the calibration factor or vice versa.

3bi.

Marks	0	1	2	Average
%	20	10	70	**1.5**

Exothermic – temperature of the water increased, indicating that energy was released

Overall performance on this question was quite strong. Students who received only one mark generally did not refer to the increase in temperature. Those who did not receive any marks generally considered, incorrectly, that the data was consistent with an endothermic reaction.

3bii.

Marks	0	1	2	3	4	Average
%	32	13	15	20	20	**1.9**

$\Delta T = 33.0 - 22.0 = 11.0 \text{ °C}$
Energy released
$$\begin{aligned} &= 0.354 \times 11.0 \\ &= 3.89 \text{ kJ} \end{aligned}$$
$$\begin{aligned} n(CaCl_2) &= 6.038/111.1 \\ &= 0.05435 \text{ mol} \end{aligned}$$
Energy released per mol
$$\begin{aligned} &= 3.89/0.05435 \\ &= 71.7 \text{ kJ mol}^{-1} \end{aligned}$$
$\Delta H = -71.7 \text{ kJ (mol}^{-1})$

Common errors included:
- using the energy value calculated in part a. or multiplying the calibration factor by 9.50, the temperature rather than the temperature change deduced from the graph provided
- not calculating the $n(CaCl_2)$ using the supplied data (6.038 g)
- not providing the 'sign' for ΔH
- not giving the final answer to three significant figures and leaving out or providing incorrect units.
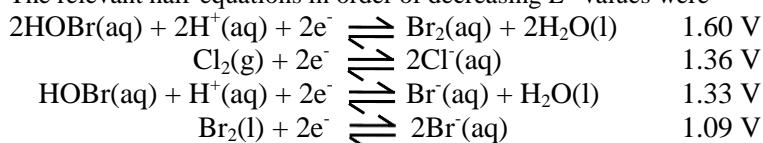
There was evidence that some students struggled to assimilate the data provided in part bi.

Question 4
4a.

Marks	0	1	2	Average
%	76	22	2	0.3

The relevant half-equations in order of decreasing $E°$ values were

$$2HOBr(aq) + 2H^+(aq) + 2e^- \rightleftharpoons Br_2(aq) + 2H_2O(l) \qquad 1.60 \text{ V}$$
$$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq) \qquad 1.36 \text{ V}$$
$$HOBr(aq) + H^+(aq) + 2e^- \rightleftharpoons Br^-(aq) + H_2O(l) \qquad 1.33 \text{ V}$$
$$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq) \qquad 1.09 \text{ V}$$

Acceptable responses included:
- reaction between $Cl_2(g)$ and $Br^-(aq)$ would be predicted to occur but to produce $Br_2(l)$ rather than $HOBr(aq)$
- the reaction predicted to occur would be $Cl_2(g) + 2Br^-(aq) \rightleftharpoons 2Cl^-(aq) + Br_2(l)$
- $Br^-(aq)$ at the lower $E°$ value of 1.09 V would be preferentially oxidised to $Br_2(l)$
- unreacted $Br^-(aq)$ to produce Br_2 thus limiting the extent of the reaction between $Cl_2(g)$ and $Br^-(aq)$.

To gain marks for this question, students needed to indicate that the product of the reaction would be $Br_2(l)$ rather than HOBr and give a reason, based on the information provided in the $E°$ values, why this was the case or why HOBr would not be expected to be produced to any significant extent.

Very few students came to grips with the range of this question, mainly because $Br^-(aq)$ as a reductant at 1.09 V on the electrochemical series was not picked up on. Many focused on $H_2O(l)$ as a reductant at 1.23 V and argued, incorrectly, that $Cl_2(g)$ would react with water. Students should be encouraged to identify all locations of a particular species on the electrochemical series before making inferences about potential reactions or lack thereof.

Many students applied the terms 'oxidant' and 'reductant' incorrectly when referring to Cl_2, Cl^- and Br^-. Students should be aware that reductants are on the right-hand side of the electrochemical series. It is perhaps useful if that knowledge extends to the understanding that in a redox pair, for example $Cl_2(g)/Cl^-(aq)$, the oxidant is the species with the higher oxidation number.

Many students gave a response similar to 'the rate of reaction may be too slow for reaction to be observed/ electrochemical series tells us nothing about reaction rate'. This was not relevant because the question stated that 'the reaction between chlorine gas and bromide ions (to produce HOBr) is **not predicted to occur to any extent** under standard conditions'.

4bii.

Marks	0	1	Average
%	55	45	0.5

$$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$$
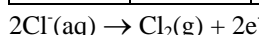
This half-equation proved to be challenging for some students.

A check of oxidation numbers shows that Cl is -1, in HCl, on the reactant side and -1, in $MnCl_2$, and 0, in Cl_2, on the product side of the equation given. However, many students seemed to overlook that HCl(aq) would be expected to be fully ionised and ran into difficulty using HCl(aq) rather than $Cl^-(aq)$ in their half-equation.

Many attempts were made to produce half-equations such as $4HCl(aq) \rightarrow Cl_2(g) + 4H^+(aq) + 2Cl^-(aq) + 2e^-$. Students should be made aware that this is not an acceptable alternative to the correct equation because it implies that the ionisation of HCl(aq) is a consequence of the oxidation.

A significant number of attempts to include MnO_2 and $MnCl_2$ in oxidation half-equations were made, suggesting that students did not identify that the oxidation number of Mn decreases from +4 to +2.

Question 5
5ai–ii.

Marks	0	1	2	Average
%	43	34	24	**0.8**

In part i. most students struggled to coherently explain the effect on the reaction rates as the system moved to equilibrium. Many did not pick up that only NO was initially present in the reaction vessel. Others tried to base their explanations on the endothermic and exothermic nature of the forward and reverse reactions respectively.

The changes in rates as the reaction approaches equilibrium can be argued via simulated rate-time graphs.
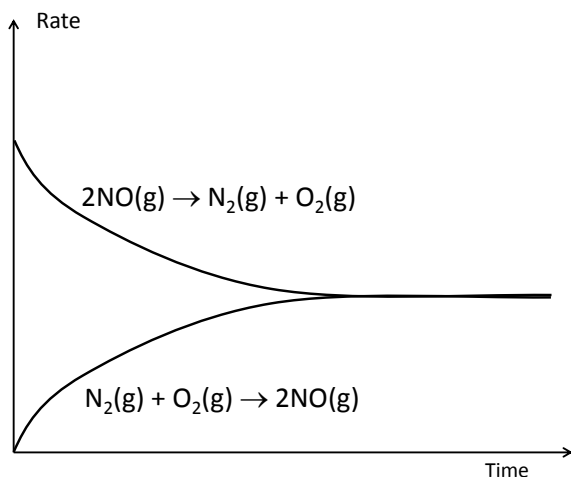
Students should be aware that the condition of equilibrium is that the rates of the forward and reverse reactions are equal.

There is scope for improved student understanding of the role of reaction rate in equilibrium and movement towards or out of equilibrium. When using Le Chatelier's principle to explain the effect of changes on a system at equilibrium, students should be aware that the response happens because of differences in the rates of forward and reverse reactions.

5ai.
Either of:
- the rate of the reverse reaction $[2NO(g) \rightarrow N_2(g) + O_2(g)]$ decreases and the rate of the forward reaction $[N_2(g) + O_2(g) \rightarrow 2NO(g)]$ increases as the system moves to equilibrium, so the reverse rate is always greater than the forward rate
- because the reverse reaction $[2NO(g) \rightarrow N_2(g) + O_2(g)]$ is favoured, as the system heads to equilibrium it must remain faster than the forward reaction $[2NO(g) \rightarrow N_2(g) + O_2(g)]$.

5aii.
Either of:
- the rates of the forward and reverse reactions are equal at equilibrium
- if the rate of either reaction was greater than the other the system would not be at equilibrium.

5bi–ii.

| Marks	0	1	2	3	4	5	Average
%	18	11	19	22	13	17	**2.5**

In part bi. most students did not draw the link between the 'combustion engine' and the endothermic nature of the production of NO. Many focused on the catalysts, overlooking the fact that they were present to speed up the decomposition of NO, not aid its production.

The enthalpy profile in part bii. was generally well handled, with a significant proportion of the students being awarded full marks. Sketching the profile for the formation of NO, an endothermic reaction, was a common error.

5bi.

The production of NO from N_2 and O_2 is endothermic, so this reaction is favoured by the high temperature in the engine.
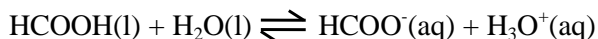
5bii.
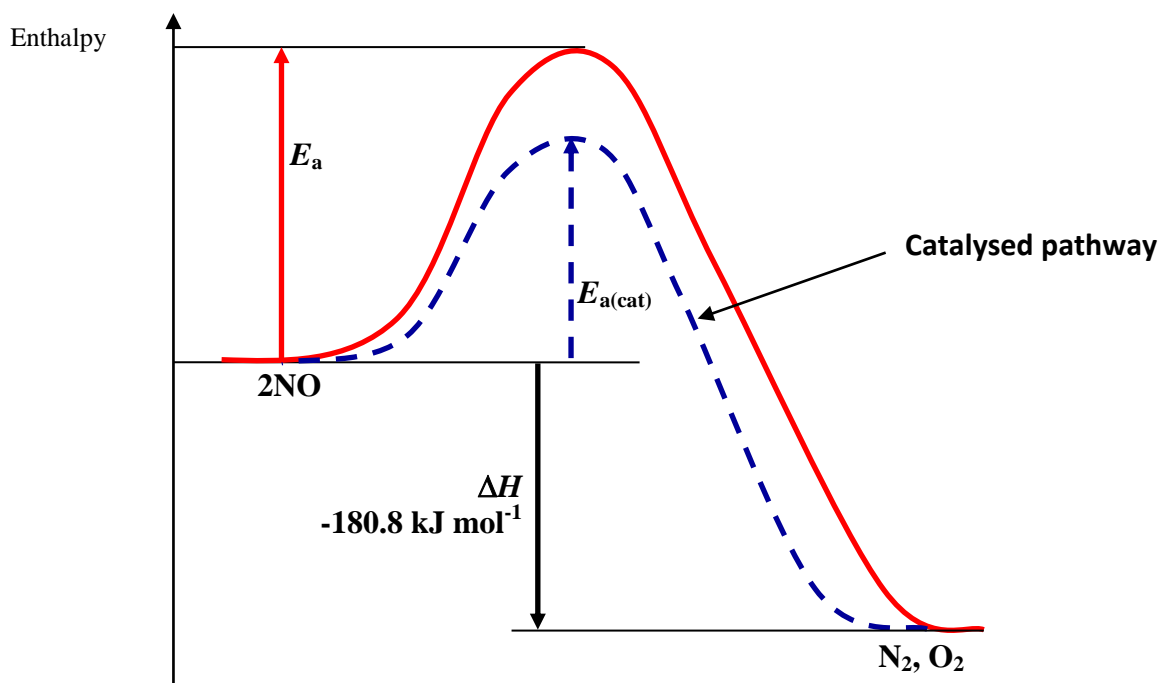
Question 6

6ai.

Marks	0	1	Average
%	56	44	**0.5**

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

Since methanoic acid is weak acid, and identified in the Data Book, equilibrium arrows were required in the equation; however, they were frequently absent.

6aii.

Marks	0	1	Average
%	26	74	**0.8**

$$K_a = [HCOO^-][H_3O^+]/[HCOOH]$$

The inclusion of $[H_2O]$ in the equilibrium law was the most common error.

6bi.

Marks	0	1	2	3	Average
%	50	27	2	20	1

$[HCOOH] = 0.500/2.00 = 0.250$ M
$[HCOO^-] = 0.100/2.00 = 0.0500$ M
$1.8 \times 10^{-4} = 0.0500 \times [H_3O^+]/0.250$
$[H_3O^+] = 1.8 \times 10^{-4} \times 0.250/0.0500$
$\quad\quad = 9.0 \times 10^{-4}$ M

This question required students to apply their knowledge of weak acids in a context that may have been different to that they had experienced previously. The equilibrium applied to an aqueous solution made from a mixture of the weak acid and its conjugate base, rather than the more familiar context which considers a solution produced by adding only the weak acid to water. Students were instructed to apply, in their calculations, the assumption that $[HCOO^-]$ at equilibrium was equal to the initial $[HCOO^-]$.

Students who did not realise the significance of adding $HCOO^-$ ions as well as HCOOH approached this question as if only methanoic acid had been added to water. These students generally applied the standard assumption for the hydrolysis of a weak acid – the equilibrium $[H_3O^+]$ = the equilibrium $[HCOO^-]$ – to set up an equilibrium expression that included $[H_3O^+]^2$, which was not appropriate in this context.

A relatively common error was calculating the initial $[HCOOH]$ as 0.300 M, as if all the 0.100 mol $HCOO^-$ added was converted to HCOOH. This may have been based on an assumption that since the conjugate base of a strong acid is a weak base then the conjugate base of weak acid is a strong acid. However, most compounds commonly referred to as weak acids, for example those in Table 12 of the Data Book, have weak conjugate bases.

Because HCOOH(aq) and $HCOO^-$(aq) are a weak acid and a weak base respectively, the designated assumptions were reasonable because the ionisation of HCOOH had minimal impact on the $[HCOO^-]$ and the reaction of $HCOO^-$ had minimal impact on the $[HCOOH]$.

6bii.

Marks	0	1	Average
%	30	70	0.7

pH $\quad = -\log_{10}(9.0 \times 10^{-4})$
$\quad\quad = 3.05$ (3.0)

The conversion of $[H_3O^+]$ to pH was generally well handled by those students who had calculated a $[H_3O^+]$ in part 6bi. While the answer 3.0, or similar depending on $[H_3O^+]$ calculated in 6bi., was accepted, it is appropriate to make the point that the pH expressed to two significant figures was 3.05. When a logarithm is found, the integer before the decimal point locates the decimal point of the number $[H_3O^+]$ and is not considered when deciding the number of significant figures.

In mathematical terms, the number to the left of the decimal point in a logarithm is called the characteristic, and the number to the right of the decimal point is called the mantissa. The mantissa has as many significant figures as the number from which the logarithm was determined. Hence $[H_3O^+] = 9.0 \times 10^{-4}$, pH = 3.05.

6biii.

Marks	0	1	2	Average
%	45	35	20	0.8

Solution B
The presence of $HCOO^-$(aq) pushes the equilibrium {HCOOH(aq) + H_2O(l) $\rightleftharpoons$ $HCOO^-$(aq) + H_3O^+(aq)} to the left, decreasing the $[H_3O^+]$ and so increasing the pH.

While the majority of students were able to identify that solution B, the solution containing $HCOO^-$(aq), had the higher pH, many struggled to give a valid explanation of why this was the case. A small number of students accurately compared the pH of the solutions, A 2.02, B 3.05.

Question 7

7ai.

Marks	0	1	2	Average
%	14	14	73	**1.6**

Exothermic – concentration (yield) of methanol increases as the temperature decreases

This question was well handled. There was a significant number of instances where students considered the temperatures 400 K, 500 K, etc. to be equilibrium constant values.

7aii.

Marks	0	1	Average
%	22	78	**0.8**

Acceptable responses included:
- to increase the rate of reaction
- at a lower temperature the rate of reaction will be too slow
- at low temperature there is a conflict between rate and yield
- 520 K provides a compromise between rate and yield factors.

7aiii.

Marks	0	1	2	Average
%	28	39	33	**1.1**

According to Le Chatelier's principle, the system will move to accommodate high pressures by moving to decrease the pressure/to (partially) oppose the change by favouring the forward reaction which produces fewer molecules.

Most responses to this question referred to the shift to the side of the equilibrium with fewer particles, but explanations of why this occurs were either superficial or not included. Many simply referred to Le Chatelier's principle in general, without mentioning that the shift to the product side is in order to oppose the applied change and decrease the pressure. Students should be made aware that the statement 'due to Le Chatelier's principle' in isolation is not a valid explanation of why a reaction shifts direction.

7b.

Marks	0	1	Average
%	39	61	**0.6**

Acceptable responses included:
- by lowering activation energy via alternative reaction pathway
- by lowering activation energy to increase the proportion of successful (fruitful) collisions
- by providing an reaction alternative pathway, which increases the proportion of successful collisions.

Most students were aware that a catalyst lowers the activation energy. The more successful students showed the required awareness of how this lower activation energy is achieved and/or how it leads to a faster reaction rate.

Question 8

8a.

Marks	0	1	Average
%	75	25	**0.3**

Either of:
- $Cu^+(aq) + e^- \rightarrow Cu(s)$
- $CuCN(aq) + e^- \rightarrow Cu(s) + CN^-(aq)$.

Most students provided a half-equation for the reduction of $Cu^{2+}(aq)$ which, while allowing for a degree of automatic default because of familiarity with the species, was surprising given that copper(I) was identified in the question.

8b.

Marks	0	1	Average
%	60	40	**0.4**

Excess CN^-(aq) [from KCN(aq)] pushes the position of equilibrium to the right, thus decreasing the $c(Cu^+)$.

The instruction 'referring to this information' in the question was generally overlooked, with most students not linking the low $[Cu2^+]$ to the fact that the reaction was an equilibrium reaction. Statements such as 'the CN^- ions react with Cu^+ ions', which effectively repeated a statement in the question, were common. Some good responses related the high value of the equilibrium constant to maintaining a low $[Cu^+]$.

8c.

Marks	0	1	Average
%	58	42	**0.4**

Acceptable responses included:
- increased concentration of OH^-(aq) from KOH(aq) pushes the position of equilibrium to the left and so decreases the concentration of HCN(aq)
- potassium hydroxide, a strong base, will neutralise the HCN.

Responses such as 'more CN^- is produced' were relatively common. Students should be made aware that while this will be a consequence of the addition of KOH, it does not adequately explain how that addition lowers the [HCN].

8d.

Marks	0	1	Average
%	71	29	**0.3**

$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$
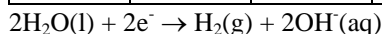
There was a variety of incorrect responses to this question, the most common involving the reduction of H^+(aq) or K^+(aq). This was unexpected, given that the reduction (or oxidation) of water is always an option in the electrolysis of aqueous solutions. The most likely option for the production of gas at the cathode, on application of higher voltage, would be the reduction of the strongest oxidant, other than Cu^+(aq) which is present in the cell, i.e. reduction of H_2O. This question should provide good impetus for discussion about effective application of the electrochemical series in unfamiliar contexts.

8e.

Marks	0	1	2	3	4	Average
%	19	7	7	19	48	**2.7**

m(Cu) formed $= 36.4 – 25.2$
$= 11.2$ g
$n(e^-) = n$(Cu)
$= 11.2/63.6$
$= 0.176$ mol
$Q = 0.176 \times 96500$
$= 1.70 \times 10^4$ C
$t = Q/I$
$= 1.70 \times 10^4/0.900$
$= 1.89 \times 10^4$ s
$= 315$ (minutes)

This question was well handled. Students who had a 'correctly balanced' half-equation for $Cu2^+$(aq) in part a. were able to receive full marks in this question for the answer 629 minutes; however, this consequential allowance did not apply for non-Cu half-equations in part a.

The most common errors on this question were associated with the calculation of the 'time', both from the charge and in the conversion to minutes.