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VCE Chemistry ¾
AOS 1 Revision II [0.11]
Workshop Solutions

Error Logbook:

New Ideas/Concepts	Didn't Read Question
Pg / Q #: ____________ Notes:	Pg / Q #: ____________ Notes:
Algebraic/Arithmetic/ Calculator Input Mistake	**Working Out Not Detailed Enough**
Pg / Q #: ____________ Notes:	Pg / Q #: ____________ Notes:

Section A: Warm Up (10 Marks)

INSTRUCTION: 10 Marks. 7 Minutes Writing.

Question 1 (1 mark)

The oxidation state of Phosphorus in the pyrophosphate ion $P_2O_7^{4-}$ is:

A. +3.5

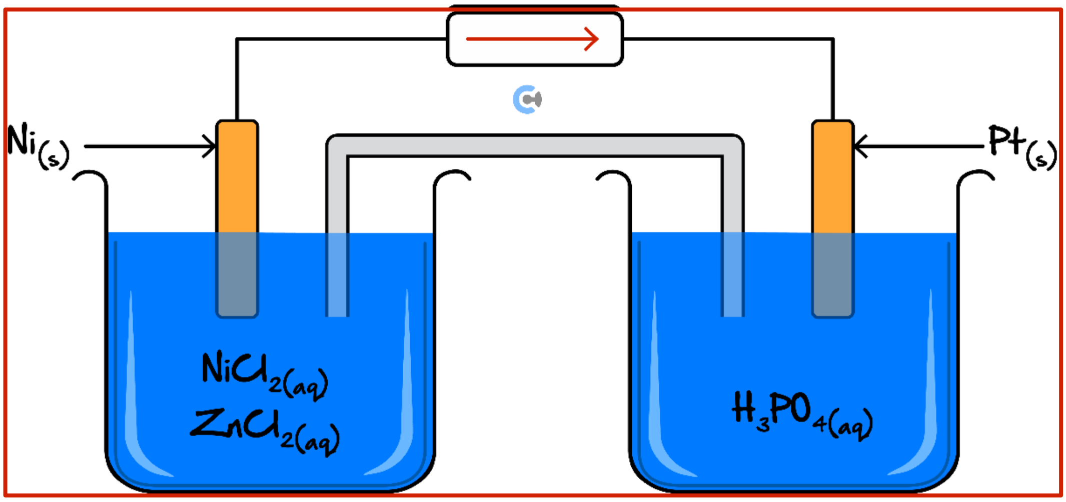
B. +5

C. +7

D. +10

Question 2 (9 marks)

A unique galvanic cell with multiple ions in the electrolyte solutions is shown below:



a.

i. State the species present which will most readily reduce. (1 mark)

H^+

ii. Write the balanced half-equation that occurs at the cathode. (1 mark)

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

iii. Thus, state the 2 things that would be observed at the cathode as this cell operates. (2 marks)

1. Bubbles due to the production of $H_2(g)$.

2. Less acidic $\rightarrow$ pH increase.

b. Label the direction of flow through the circuit by placing an arrow in the box provided. (1 mark)

c. Explain why the left half-cell does not use zinc as its electrode. Ensure to reference the energy transformations and any relevant half-equations. (2 marks)

If Zn(s) were used it would spontaneously react with the Ni^{2+} in the electrolyte: $Zn \rightarrow Zn^{2+} + 2e^-$ and $Ni^{2+} + 2e^- \rightarrow Ni$. This would produce thermal energy and disrupt the galvanic cell intending to produce electrical energy.

d. The cell was constructed to determine the EMF of the cell according to the electrochemical series.

i. State the theoretical EMF expected to be observed. (1 mark)

$0.40\ V$

ii. What would the concentration of Phosphoric acid (H_3PO_4) be in their set-up? (1 mark)

$c(H^+) = 1.0\ M \rightarrow c(H_3PO_4) = 0.33\ M$

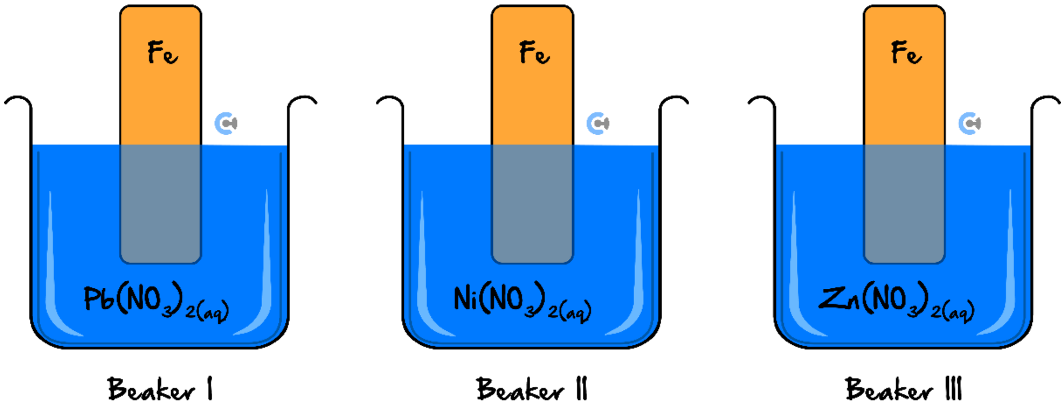
Space for Personal Notes

Section B: Ramping Up (8 Marks)

INSTRUCTION: 8 Marks. 6 Minutes Writing.

Question 3 (1 mark)

Three beakers, each containing an iron strip and a 1.0 *M* solution of a metal salt were set up as follows.



A reaction will occur in beaker(s):

A. I and II only.

B. I and III only.

C. II and III only.

D. III only.

Question 4 (1 mark)

Some strips of the metals iron, zinc and silver were placed in separate beakers, each containing 1.0 *M* nickel(II) sulfate, $NiSO_4$, solution in water at 25°C.

What is expected to occur over time?

A. Ni will be deposited in all of the beakers.

According to the electrochemical series, the reductants Fe and Zn are below the oxidant Ni^{2+}(aq) and will be oxidised to Fe^{2+}(aq) and Ni^{2+}(aq).

B. Ni will not be deposited in any of the beakers.

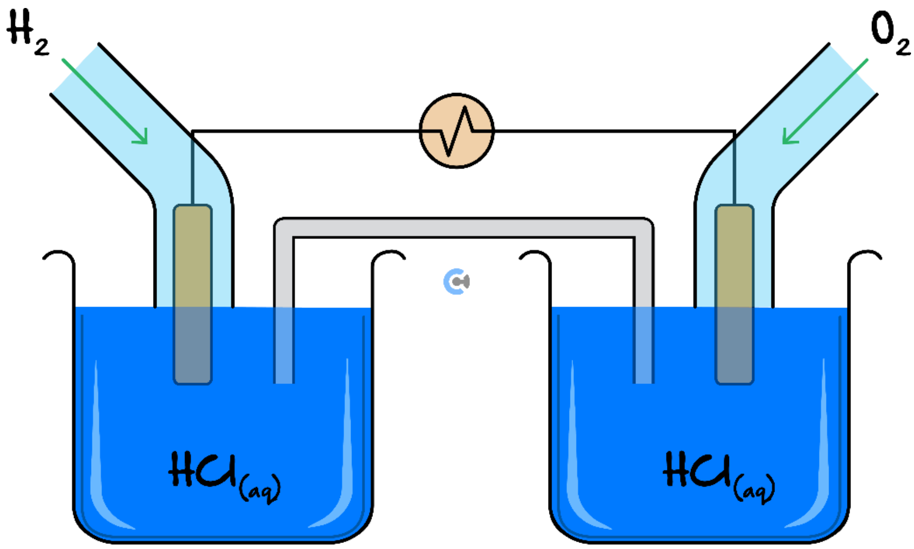
C. A reaction will occur only in the beaker containing Ag.

D. A reaction will occur only in the beakers containing Fe and Zn.

Question 5 (6 marks)

Hydrogen and oxygen gas cells are set up and compared.

a. The following hydrogen and oxygen gas cell, Cell *A* was set up in a school laboratory.



i. State a material that can be used as a material for the electrode. Justify your answer. (1 mark)

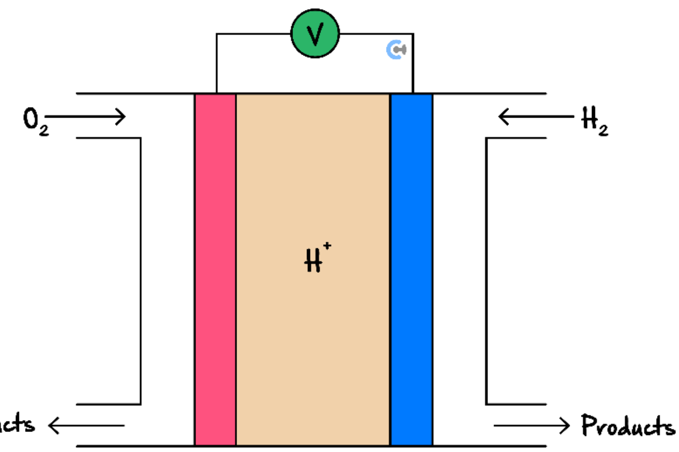
Graphite → conductive and inert.

ii. Write the half-equation which occurs at the: (2 marks)

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

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

Suppose a pH meter was placed into the electrolyte of the following cell, Cell *B*:



b. Explain what would be observed when measuring the pH in Cell B compared to if pH meters were placed into each of the electrolytes of the half-cells in Cell B. (3 marks)

In a fuel cell, H^+ is produced as well as consumed through the **shared** electrolyte, so a pH change would NOT be observed. (no H^+ in overall equation) (1)
However, in the hydrogen/oxygen cell the pH metrs would show changes in pH: at the anode, H^+ is produced, so pH would decrease (2), whereas at cathode, H^+ is consumed, so pH would increase. (3)

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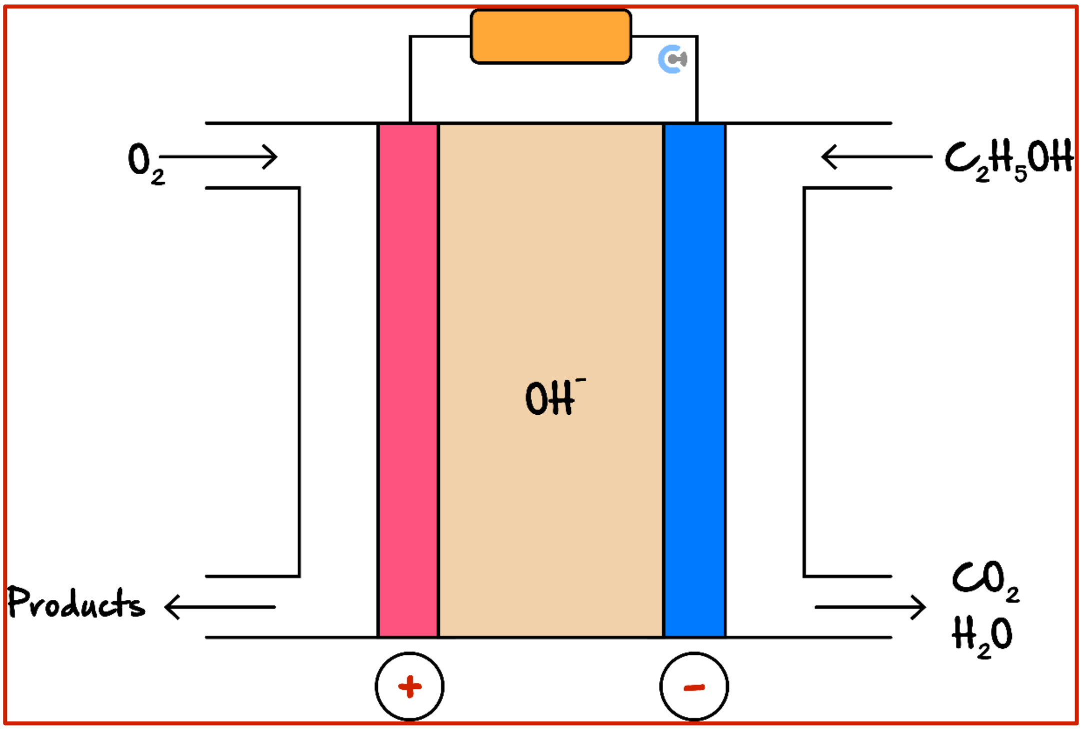
Section C: Getting Trickier I (11 Marks)

INSTRUCTION: 11 Marks. 9 Minutes Writing.

Question 6 (11 marks)

The following cell was prototyped at standard conditions by NASA for its potential widespread use in society in the near future.



a. State if the cell above is classified as a galvanic cell or not. Justify your answer. (1 mark)

Yes, it is. A galvanic cell is merely any electrochemical cell which converts chemical energy into electrical energy; it has nothing to do with it being a fuel cell or primary or secondary cell.

b. Label the polarity of the electrodes by placing a + and – in the circles provided. (1 mark)

c. Write the balanced equation and overall reaction that takes place. (1 mark)

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

d. Write the balanced half-equation for the:

i. Oxidation half-cell. (1 mark)

$C_2H_5OH(l) + 12OH^-(aq) \rightarrow 2CO_2(g) + 9H_2O(l) + 12e^-$

ii. Reduction half-cell. (1 mark)

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

e. It is found that this cell generates 1.174 V at SLC.

i. Find the standard electrode potential of the oxidation half-equation. (1 mark)

$$1.174 = 0.40 - E^0$$
$$E^0 = -0.774\ V$$

ii. Calculate the volume of oxygen needed for this cell to be able to generate 3.00 MJ of energy. (3 marks)

$$n(C_2H_5OH) = \frac{q}{\Delta H} = \frac{3 \times 10^3}{1370} = 2.19\ mol$$
$$n(O_2) = 3n(C_2H_5OH) = 6.57\ mol$$
$$V(O_2) = n \times V_m = 6.57 \times 24.8 = 163\ L$$

f. These cells are yet to be widely implemented in society.

State **two** features of these cells that make them difficult to use in day-to-day devices such as motor vehicles. Explain your answer with reference to at least **one** United Nations Sustainable Development Goals. Use item **26. i.** of the Data Book. (2 marks)

Constant supply of reactants → impractical for portable use.
Expensive due to PICCY electrodes → not feasible for day-to-day use in society.
Goal 7 – "Affordable and clean energy" the cell is clean but not affordable
Goal 11 – "Sustainable cities and communities" – the cell is not integrated into cities and communities.

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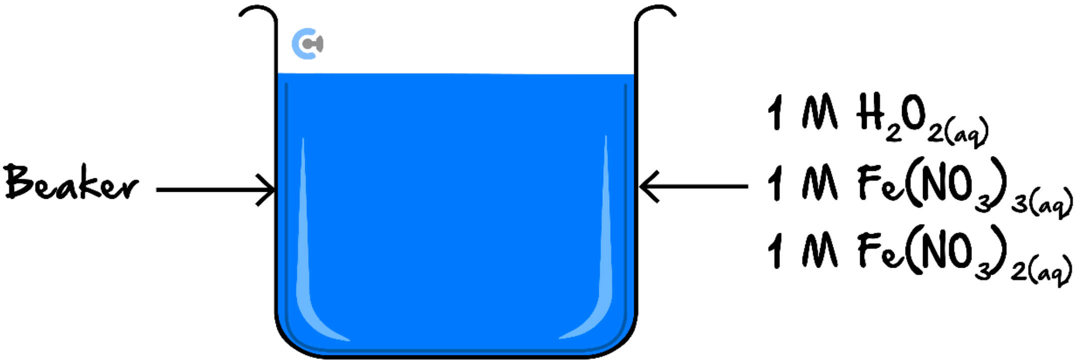
Section D: Getting Trickier II (11 Marks)

INSTRUCTION: 11 Marks. 10 Minutes Writing.

Question 7 (1 mark)

At standard conditions, solutions of Hydrogen peroxide, H_2O_2 , iron(III) nitrate, $Fe(NO_3)_3$ and iron(II) nitrate, $Fe(NO_3)_2$, were added to a beaker. The initial concentrations of H_2O_2, $Fe(NO_3)_3$ and $Fe(NO_3)_2$ in the beaker were all 1 M.



Which one of the following statements is correct?

A. Iron, Fe, is deposited at the

B. The two half-reactions in the

C. The concentration of H_2O_2(

D. The temperature of the conte
H_2O_2(aq).

30.	C	26	11	**48**	15	The relative locations of the species H_2O_2(aq), Fe^{3+}(aq) and Fe^{2+}(aq) on the electrochemical series: $H_2O_2(aq) + 2H^+(aq) + 2e^- \leftrightharpoons 2H_2O(l)$ +1.77 V $Fe^{3+}(aq) + e^- \leftrightharpoons Fe^{2+}(aq)$ +0.77 V $O_2(s) + 2H^+(aq) + 2e^- \leftrightharpoons H_2O_2(aq)$ +0.68 V $Fe^{2+}(aq) + 2e^- \leftrightharpoons Fe(s)$ −0.44 V H_2O_2(aq) is both the strongest oxidising agent and the strongest reducing agent present. The equation for the redox reaction occurring will be $2H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(l)$. So, the concentration of H_2O_2(aq) decreases. The reaction cannot produce a voltage because it is not in an electrochemical cell. The reaction is exothermic, so the temperature of the contents increases.

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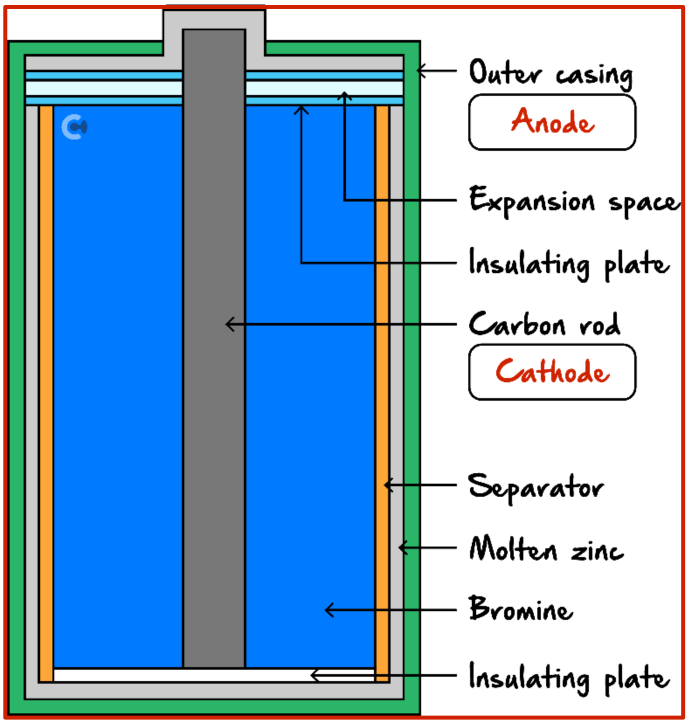
Question 8 (10 marks)

The following primary cell is being studied by chemists to determine whether it is viable for day-to-day use in society.



Given the overall reaction occurring is:

$$Zn(l) + Br_2(l) \rightarrow ZnBr_2(l)$$

a.

i. Explain, using oxidation numbers, which species is the reducing agent. (1 mark)

Zn(l) as oxidation number increases from 0 to +2, so it undergoes oxidation and acts as the reducing agent.

ii. Identify which physical component is acting as the anode in the cell - the outer casing or the carbon rod - by labelling one of the boxes on the diagram above. (1 mark)

iii. Label the cathode above and explain the chemistry occurring in the region surrounding the cathode. (2 marks)

The $Br_2(l)$ will migrate towards the inert carbon rod, receive electrons through it (as it a conducting electrode) and then reduce into bromide ions (Br^-).

b. Explain the purpose of the separator. (1 mark)

To ensure the reactants do not directly come into contact, preventing a direct redox reaction.

c. The insulating plates within the cell ensure humans do not get burned when handling the cell. List and explain **two** reasons why the cell gets very hot when operating. (2 marks)

1. The zinc is molten which means the cell must already be operating at very high temperature.
2. The reactions are exothermic which release heat and make the cell even hotter.

d. The cell is left running for an hour and produces a stable current of 5.0 A and a total of 300 kJ of electrical energy.

i. Calculate the electromotive force produced by the cell. (2 marks)

$$Q = It = 5.0 \times 60 \times 60 = 18\,000\ C$$

$$V = \frac{E}{Q} = \frac{3 \times 10^5}{18\,000} = 16.7\ V$$

ii. If the main goal was to produce the maximal possible EMF, propose an alternative molten metal to use in this cell instead of zinc. (1 mark)

Anything lower than Zn on ECS → Li, for example.

Let's take a BREAK!

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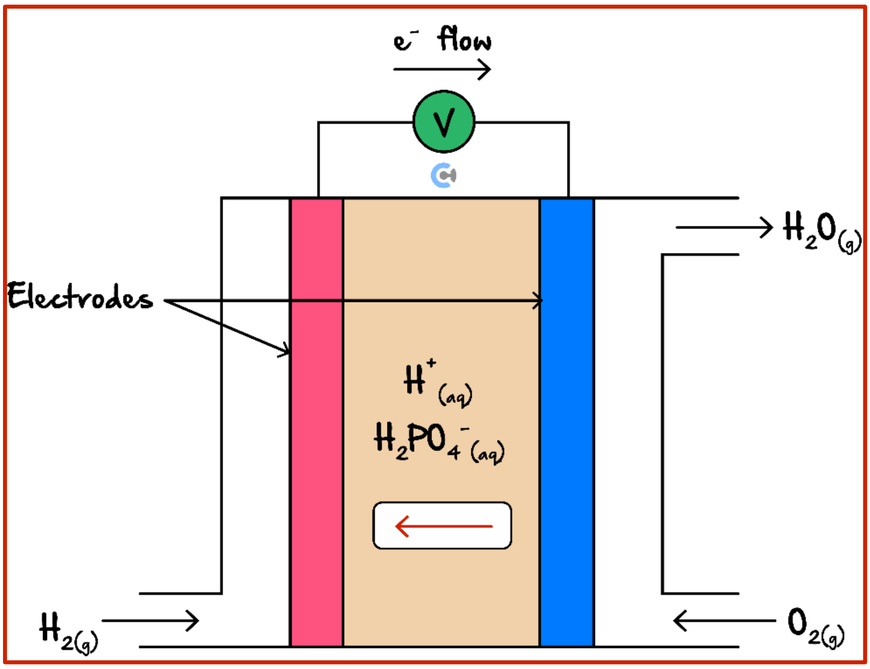
Section E: VCAA-Level Questions I (8 Marks)

INSTRUCTION: 8 Marks. 0.5 Minutes Reading. 7 Minutes Writing.

Question 9 (8 marks)

A fuel cell that can provide power for buses is the phosphoric acid fuel cell, PAFC. The electrolyte is concentrated phosphoric acid and the reactants are hydrogen and oxygen gases.

A simplified sketch of a phosphoric acid fuel cell is given below.



a. Give the equation for the half-reaction that takes place at the:

i. Anode of this cell. (1 mark)

Question 8ai-ii.

Marks	0	1	2	Average
%	37	20	43	1.1

8ai.
$H_2(g) \rightarrow 2H^+(aq) + 2e^-$*

ii. Cathode of this cell. (1 mark)

8aii.
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$*

These half equations could be obtained directly from the electrochemical series.

b. **On the diagram** of the fuel cell, draw an arrow to show the direction in which the $H_2PO_4^-$ ion moves as the cell delivers an electrical current. (1 mark)

Marks	0	1	Average
%	48	52	0.6

c.

i. A particular cell operates at 0.92 V. How much energy, in kJ, is delivered per mole of hydrogen in this fuel cell? (2 marks)

8ci.
$n(e^-) = 2$ mol
$Q = 2 \times 96500$* C
$E = QV$
$= 2 \times 96500 \times 0.92$
$= 1.78 \times 10^5$ J
$= 1.8 \times 10^2$* (kJ)

Question 8ci. was poorly done with most students unable to make the links between E = VIt and Q = It to get the required relationship E = VQ. In 8cii. the heat of combustion of H_2 was available from the data book.

ii. By comparing the energy delivered per mole of hydrogen in the fuel cell and the heat of combustion of hydrogen, calculate the energy efficiency of this fuel cell. (1 mark)

8cii.
Efficiency = (178 / 286) x 100
= **62.0* % 6**

d. Describe one advantage and one disadvantage of such a fuel cell compared with a petrol-driven car engine. (2 marks)

Advantage:

Question 8d.

Marks	**0**	**1**	**2**	**Average**
%	21	35	44	**1.3**

- advantage: Less energy is lost as heat/less noisy/more efficient energy conversion/no CO_2 produced/only product is H_2O*
- disadvantage: Cost/hydrogen is difficult to produce/store/distribute*

Disadvantage:

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Section F: Multiple Choice Questions (9 Marks)

INSTRUCTION: 9 Marks. 9 Minutes Writing.

Question 10 (1 mark)

Inspired from VCAA Chemistry Exam 2002
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/chem22002.pdf#page=4

A VCE chemistry student sets up a galvanic cell using two standard half-cells with half-reactions:

Half-cell 1: $Cr^{3+}(aq) + e^- \rightarrow Cr^{2+}(aq)$

Half-cell 2: $Cr(s) \rightarrow + Cr^{2+}(aq) + 2e^-$

Suitable materials for the electrodes of the two half-cells are:

	Half-cell 1	**Half-cell 2**
A.	Platinum	Platinum
B.	Platinum	Chromium
C.	Chromium	Chromium
D.	Chromium	Platinum

Question 11 (1 mark)

Which statement about the electrochemical series is correct?

A. Strong oxidising agents have strong conjugate-reducing agents.

B. The stronger the reductant, the more positive the E^0 value.

C. The stronger the reductant, the closer to zero the E^0 value.

D. Strong reducing agents donate electrons more readily than weak ones.

Strong oxidising agents have *weak* conjugated reducing agents (A incorrect) and the stronger the reductant the more *negative* the E^0 value (B, C incorrect).

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Question 12 (1 mark)

Perchloric acid, $HClO_4$, is an extremely strong acid and a very powerful oxidising agent.

Which of the following formulae represents the substance **least likely** to be produced when perchloric acid acts as the oxidant in a redox reaction?

A. Cl_2

B. Cl_2O

C. Cl_2O_2

D. Cl_2O_7

When $HClO_4$ acts as an oxidant, the oxidation number of chlorine changes. Since $HClO_4$ is the oxidant, it causes oxidation and as a consequence is itself reduced. Hence the oxidation number of Cl should decrease.
Applying the general rules of oxidation numbers, the oxidation numbers of Cl in the four alternatives are
Cl_2: 0
Cl_2O: +1
ClO_2: +4
Cl_2O_7: +7
The oxidation number of Cl in $HClO_4$ is +7.
Cl_2O_7 will not be a reduction product of $HClO_4$ because the oxidation number of Cl is +7 in both.

Question 13 (1 mark)

The iron(II) ion, $Fe^{2+}(aq)$:

A. Can act as an oxidant but not

B. Can oxidise solid zinc and red

C. Can act as a reductant but not

D. Will always be reduced to Fe(s) in redox reactions.

According to the relevant equations from the electrochemical series
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$ 1.09 V
$Fe^{3+}(aq) + e^- \rightleftharpoons \mathbf{Fe^{2+}}(aq)$ 0.77 V
$\mathbf{Fe^{2+}}(aq) + 2e^- \rightleftharpoons Fe(s)$ -.0.44 V
$Zn^{2+}(aq) + 2e^- \rightleftharpoons \mathbf{Zn}(s)$ -0.76 V
These half-equations show that, depending on the species with which it is reacting, **Fe^{2+}(aq) can act as an oxidant** (oxidising Zn to Zn^{2+}) **and a reductant** (reducing Br_2 to Br^-). In these roles Fe^{2+}(aq) will be converted to Fe(s) and Fe^{3+}(aq) respectively.

Question 14 (1 mark)

A series of galvanic cells were constructed using four metals (A, B, C, D) and four 1 M solutions $(A^{2+}, B^{2+}, C^{2+}, D^{2+})$, it was found that:

- When the half-cell A/A^{2+} was connected to a B/B^{2+} half-cell, electrons flowed from metal A to metal B.
- When metal C was added to a solution of D^{2+}, a coating of metal D appeared on metal C.
- When the half-cell A/A^{2+} was connected to a D/D^{2+} half-cell, the mass of metal D decreased.

The ranking of metals from weakest reductant to strongest reductant would be:

A. B, A, D, C

B. A, B, C, D

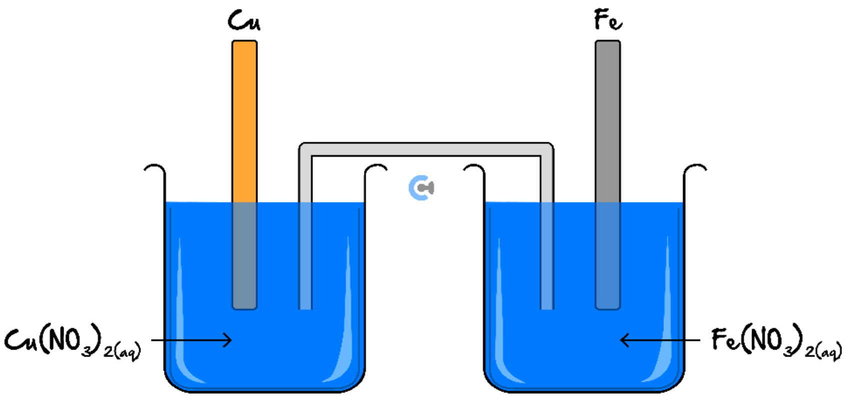
C. C, D, A, B

D. B, D, A, C

- A is below B on the electrochemical series
- C is below D on the electrochemical series
- D is below A on the electrochemical series

Question 15 (1 mark)

A student sets up a galvanic cell using two standard half-cells as illustrated below:



The solutions are connected to each other with a salt bridge consisting of an inverted U-tube containing an appropriate electrolyte.

Which species below could be used as the electrolyte for the salt bridge in the illustrated cell?

A. CH_3OH

B. $NH_4(NO_3)$

C. $AgNO_3$

D. KOH

Reaction occurs between Cu^{2+} (in cathode) and Fe (in anode).

CH_3OH is not an electrolyte.

Ag^+ will migrate from the salt bridge to the cathode. As Ag^+ is a stronger oxidant than Cu^{2+}, it will react, interfering with cell reactions. Therefore, we cannot use $AgNO_3$.

K^+ is a very weak oxidant, therefore, it will not interfere with reactions. OH^- will move into the anode where it reacts with Fe^{2+} to form a precipitate. Therefore, KOH cannot be used.

Question 16 (1 mark)

A galvanic cell can be constructed by combining a standard hydrogen half-cell to a standard Ni^{2+}/Ni half-cell. In this galvanic

A. The conc

B. Reductio

C. The pH c

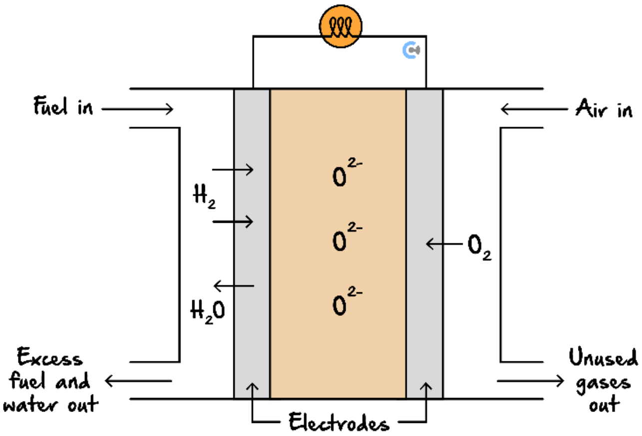
D. Hydrogen gas would be consumed at the anode.

From the order in the Electrochemical series:
H^+ / H_2
Ni^{2+} / Ni
The strongest oxidant is H^+ and the strongest reductant is Ni. Spontaneous reaction is expected.
Ni will react with H^+.
Concentration Ni^{2+} would increase. Not A.
Reduction occurs at the cathode which is part of the half cell H^+ / H_2 not Ni^{2+} / Ni. Not B.
H^+ is being used up so pH would increase. **Correct Answer: C**
H^+ is being used up and H_2 is being produced at the anode. Not D.

Space for Personal Notes

The following information applies to the two questions that follow

Solid oxide fuel cells (SOFC) are being considered for power generation and for use in space because of their high efficiency, high power density and extremely low pollution. A simplified diagram showing the key parts of a SOFC is shown below.



The electrode reactions are:

Reaction 1: $O_2(g) + 4e^- \rightarrow 2O^{2-}(s)$

Reaction 2: $H_2(g) + O_2(s) \rightarrow H_2O(l) + 2e^-$

Question 17 (1 mark)

When the SOFC is generating electricity, which of the following statements is most likely to be correct?

I. Chemical energy is completely converted into electrical energy.

II. The reduction of the oxidant consumes electrons.

III. The fuel undergoes oxidation and its oxidation number Increases.

A. I only.

B. II and III only.

C. I and II only.

D. III only.

Question 18 (1 mark)

In this cell:

A. Reaction 2 occurs at the anode which is negative.

B. Reaction 2 occurs at the anode which is positive.

C. Reaction 2 occurs at the cathode which is negative.

D. Reaction 2 occurs at the cathode which is positive.

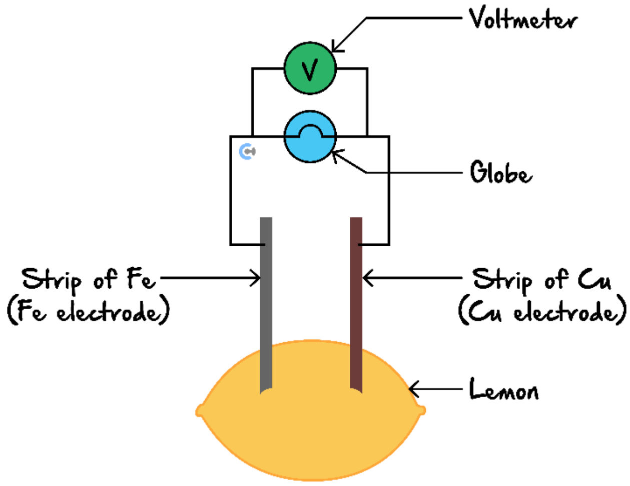
Section G: VCAA-Level Questions II (8 Marks)

INSTRUCTION: 8 Marks. 0.5 Minutes Reading. 7 Minutes Writing.

Question 19 (8 marks)

Some students set up a 'lemon battery' experiment using freshly sanded strips of copper, Cu, and iron, Fe, as the electrodes.

The students closed the switch, measured the voltage and determined the direction of the flow of electrons in the external circuit. A diagram depicting their experiment and a table of their results are shown below.



Results:

Voltage	0.35 V
Movement of electrons.	From Fe to Cu.

a. Lemon juice can be considered as an aqueous solution of H^+ ions.

Write balanced half-equations for the reactions at the electrodes in the copper-iron lemon battery. (2 marks)

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

Fe electrode: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

b. Identify the electrode that acts as the cathode. (1 mark)

Copper(Cu).

c.

i. Using the electrochemical series, find the theoretical voltage for this lemon battery. (1 mark)

$0.44\ V$

ii. Explain why the experimental value of the voltage is different from the theoretical voltage. (2 marks)

Acceptable responses included:

- The setup was not at standard conditions.
- The c(H^+) in a lemon is not 1.0 M; the temperature was not set at 25°C.
- Other reactions may have been occurring in the lemon.

d. The Cu electrode in the lemon battery is replaced with a freshly sanded strip of aluminium, Al.

For this aluminium-iron lemon battery, state the expected direction of electron flow. Justify your answer. (2 marks)

From Al (aluminium) to Fe (iron), i.e. electron flow will reverse and either:

- Al is a stronger reductant than Fe and so will be the site of oxidation.
- Al is oxidised, and H^+ (aq) is reduced at the Fe.

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