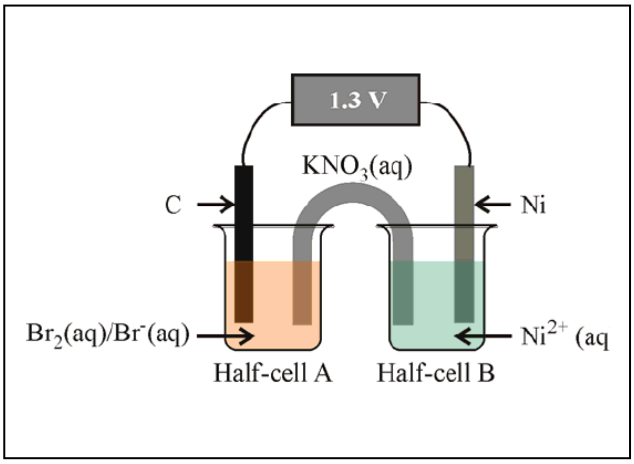
Question 1
The maximum voltage produced when standard Fe^{3+}(aq)/Fe^{2+} (aq) and Fe^{2+} (aq)/Fe(s) half-cells are combined to produce a galvanic cell, will be
A. 1.98 V
B. *1.21 V*
C. 1.10 V
D. 0.33 V

Question 2
The spontaneous reaction that occurs when a piece of zinc metal is placed in an aqueous solution of copper(II) sulfate can be represented by the thermochemical equation
$Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$ $\Delta H = -218$ kJ mol^{-1}
The same chemical reaction occurs when the Cu^{2+} (aq)/Cu(s) and Zn^{2+} (aq)/Zn(s) standard half-cells are combined to form a galvanic cell. Compared to the total energy produced when the materials are mixed directly, the total amount of energy produced from the galvanic cell will be
A. larger because of the more efficient conversion of chemical energy into electrical energy.
B. smaller because in the galvanic cell the process can take longer.
C. *the same because the reactants and products are the same*.
D. smaller because the electrons have to travel from the anode to the cathode through the external circuit.

Question 3
A galvanic cell was constructed by combining the Br_2 (aq) / Br-(aq) and Ni^{2+}(aq) / Ni(s) half-cells as shown in the diagram.



For the cell pictured above which option below is correct.

	Positive electrode	Oxidant	Potassium ions will flow into
A	Nickel	Bromine, Br_2 (aq)	Half-cell B
B	Carbon	Bromide ion, Br-(aq)	Half-cell A
C	Nickel	Nickel ions (Ni^{2+} (aq))	Half-cell B
D	*Carbon*	*Bromine, Br_2(aq)*	*Half-cell A*

Question 4
When comparing a fuel cell with a galvanic cell, one major difference is that in fuel cells
A. the charge on the cathode is the opposite of that on a galvanic cell.
B. the energy conversion is more efficient.
C. the potential difference will always be about 2V.

D. *the oxidant and reductant are continually being replaced*.

Question 5

The salt bridge and connecting wire are removed from an operating galvanic cell and discarded. If the contents of the two half-cells are placed in one container, what will be the likely outcome?

A. The reaction will continue as electrons can move through the liquid in the container.

B. *The reaction will continue as heat energy rather than electrical energy is produced*.

C. The reaction will stop as there is no salt bridge to transport the anions and cations.

D. The reaction will stop as there is no connecting wire allowing electrons to travel.

Question 6

The following equation shows the chemical reaction which occurs in a galvanic cell.

$$2\overset{0}{Cd}(s) + 2OH^-(aq) + \overset{+3}{Ni}_2O_3(s) + 3H_2O(l) \rightarrow 2\overset{+2}{Cd}(OH)_2(s) + 2\overset{+2}{Ni}(OH)_2(s)$$

Statements I to IV relate to the above chemical reaction.

I Cd undergoes oxidation.

II OH^- is the reducing agent.

III The oxidation number of nickel changes from +3 to +2.

IV The nickel ion undergoes reduction and so it is the oxidising agent.

Which of these statements are correct?

A. I and II only

B. II and III only

C. *I, III and IV only*

D. I, II, III and IV

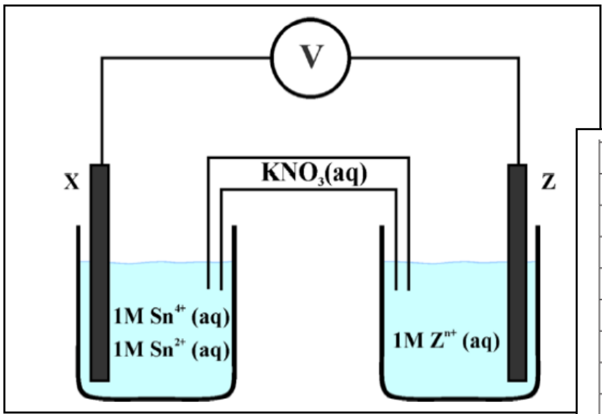
Oxidant = oxidising agent

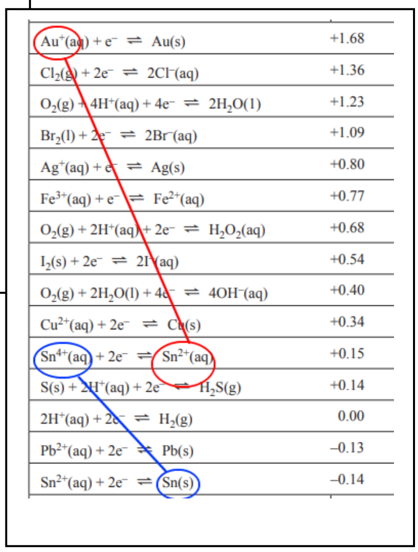
Reductant = Reducing agent

The oxidant is itself reduced at the cathode (+) while the reductant is itself oxidised at the anode(-). OH^- (aq) is not a reducing agent because it does not become oxidised.

The information below applies to Questions 7 and 8.

The diagram below represents a simple galvanic cell assembled in a laboratory.





$Au^+(aq) + e^- \rightleftharpoons Au(s)$	+1.68
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.09
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$	+0.68
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(g)$	+0.14
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14

Question 7

Which of the following elements would be **least** suitable for electrode X?

A. Carbon (graphite).

B. Silver.

C. Platinum.

D. *Tin.*

For this galvanic cell

(Sn^{4+}(aq)|Sn^{2+}(aq) ||Z(s)|Z^{n+}(aq))

we ideally place an inactive (inert) electrode in to the half-cell on the left. In this case however Sn and Sn^{4+}(aq) are likely to undergo some reaction, as indicated by the blue line on the E^0 series on the right.

Question 8

If the galvanic cell has a potential difference of 1.53 V at SLC,

A. the half-equation for the reaction at electrode Z is $Z(s) \rightarrow Zn^+(aq) + ne^-$.

B. *K^+(aq) ions are moving towards electrode Z.*

C. the cathode is electrode X.

D. number of tin(IV) ions in the cell is decreasing.

$Au^+(aq) + e \rightarrow Au$ +1.68

$Sn^{4+}(aq) + 2e \rightarrow Sn^{2+}$ +0.15

Emf = 1.68 – 0.15 = 1.53V

K^+ move to the cathode which is where the oxidant is (Au^+ (aq))

1. A lead-acid battery is made up of six cells connected in series. When the battery is providing energy, the reactions occurring at the electrodes of a single cell are:

$Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e^-$

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

a. i. Give an equation for the net reaction that occurs while a lead-acid battery is providing energy. 1 mark

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

ii. Give the formula of the oxidant and the formula of the reductant in the above reaction.

oxidant _______________$PbO_2(s)$____________________

reductant _______________*$Pb(s)$*____________________ 2 marks

b. What happens to the pH when the battery is being recharged? Explain 2 marks

When discharging the following reaction takes place at the cathode

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

When recharging, however, the reverse reaction takes place

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

Since H^+ is being formed during recharging the pH falls.

c. Write the equation occurring at the negative terminal when the battery is being recharged. 1 mark

During recharging the negative electrode is the cathode where reduction takes place.

$PbSO_4(s) + 2e^- \rightarrow Pb(s) + SO_4^{2-}(aq)$

Notice it is the opposite to the reaction taking place at the anode when the battery is discharging.

Revise secondary cells.

2. NiCad batteries are secondary cells. The chemical reaction that occurs when a NiCad cell is being recharged can be represented by the chemical equation below

   $Cd(OH)_2(s) + 2\ Ni(OH)_2(s) \rightarrow Ni_2O_3(s) + Cd(s) + 3\ H_2O(l)$

   i. What is the reductant when the NiCad cell is discharging? 1 mark

   When recharging $Cd(OH)_2(s)$ goes to Cd (Cd is reduced from 2+ to 0)

   So during discharging this reaction is reversed and hence oxidation of Cd(s) takes place to $Cd(OH)_2$

   ii. When a NiCad cell is being recharged what terminal of the external power supply must be attached to the cadmium electrode? 1 mark

   The negative terminal of the battery is attached to the negative terminal of the recharger.

   At the cadmium electrode the following reaction takes place when the cell is recharging.

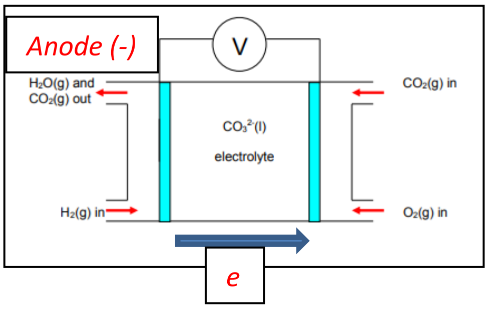
   $Cd(OH)_2(s) + 2e \rightarrow Cd(s) + 2OH^-(aq)$

   This is a reduction reaction and hence takes place at the cathode (-) terminal of the power source.

   iii. Write a balanced chemical reaction, with states, that occurs at the cathode during recharge.

The reaction that occurs at the cathode is always a reduction reaction.
$Cd(OH)_2(s)$ + 2e-> $Cd(s) + 2OH^-$ (aq)

3. A molten carbonate fuel cell (MCFC) uses a molten mixture of lithium carbonate, Li_2CO_3 and sodium carbonate, Na_2CO_3 as the electrolyte. Hydrogen gas is passed over one electrode and a combination of oxygen gas and carbon dioxide gas is passed over the other electrode, as shown in the diagram below. The net overall reaction is H_2 (g) + O_2 (g) -> H_2O (g). There is no net gain or loss of the electrolyte.



a. Write a half equation for the overall cell reaction at the anode.
___________*$H_2(g) + CO_3^{-2}(l)$ -> $H_2O(g) + CO_2(g)$ + 2e*_______________________________ 1 mark

b. Write a half equation for the overall cell reaction at the cathode.

___________ *$2CO_2(g) + O_2$ + 4e -> $2CO_3^{-2}(l)$* ________________________ 1 mark

c. On the diagram above, label

i. the anode and its polarity.

ii. the direction of electron flow. 2 marks

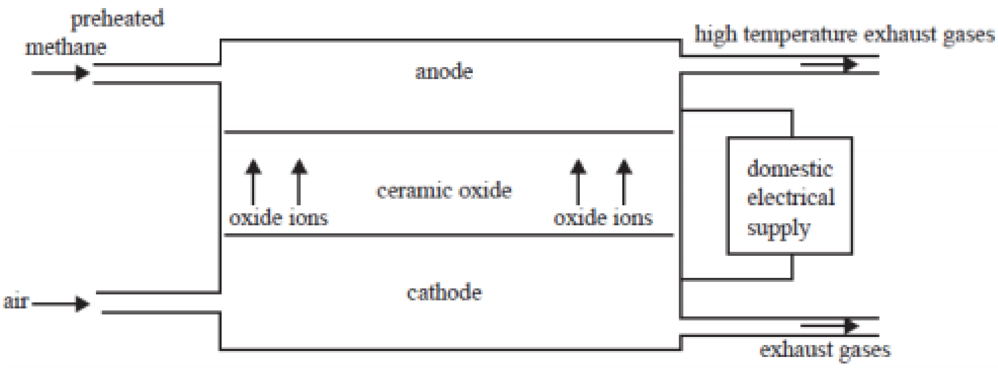
d. What is the net overall effect on the molten carbonate electrolyte as the cell produces energy?
Net zero change in CO_3^{-2} concentration in the electrolyte

The overall cell reaction is given below
($H_2(g) + CO_3^{-2}(l)$ -> $H_2O(g) + CO_2(g)$ + 2e) X 2
$2CO_2(g) + O_2$ + 4e -> $2CO_3^{-2}(l)$

$2CO_2(g) + O_2 + 2H_2(g) + 2CO_3^{-2}(l)$ -> $2CO_3^{-2}(l) + 2H_2O(g) + 2CO_2(g)$

Cancel the substances that appear on both sides (in blue)
and the overall reaction is given as $O_2(g) + 2H_2(g)$ -> $2H_2O(g)$

4. A Victorian company produces solid oxide fuel cells for use in the home. These fuel cells use natural gas to produce electricity through an electrochemical process summarized in the diagram below.



They operate at temperatures in excess of 500^{o}C.

a. What are the exhaust gases produced from this fuel cell?

____________*$CO_2(g)$*______________________ ________________$H_2O(g)$___________________ 2 marks

b. Write the overall cell reaction from this fuel cell including states. 2 marks

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

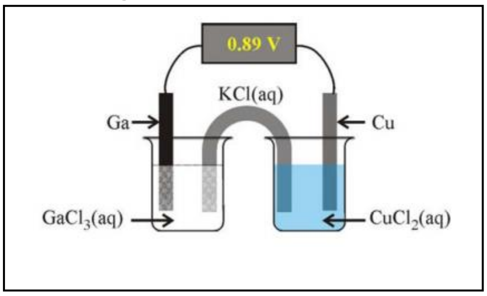
c. Give one advantage of this type of electricity production over a coal fired power station. 1 mark

It is more efficient due to the reduced number of energy conversion required to produce elecetrical energy.

d. How could the high temperature waste exhaust gases produced from this reaction best be utilised? 1 mark

Supply heat energy

5. a. A galvanic cell was assembled by combining the Cu^{2+}(aq)/Cu(s) and Ga^{3+}(aq)/Ga(s) standard half-cells as shown in the diagram below.



The cell potential was measured at 0.89 V, with the copper electrode gaining mass when the cell was discharging.

i. Write an appropriate half-equation for the process that would be occurring at the gallium electrode when the cell is discharging. 1 mark

Ga(s) -> Ga^{3+}(aq) + 3e

ii. Write an appropriate chemical equation for the overall reaction that would occur in this cell when it is discharging. 1 mark

2Ga(s) + 3Cu^{2+}(aq)-> 2Ga^{3+}(aq) + 3Cu(s)

iii. Determine the standard electrode potential (E^o) for the Ga^{3+}(aq)/Ga(s) standard half-cell. 1 mark

Cu^{2+}(aq) + 2e -> Cu(s) *+0.34V*

Ga^{3+}(aq) + 3e -> Ga(s) *-0.54V*

iv. On the diagram above clearly label the direction of flow of the potassium ions in the salt bridge 1 mark

From the gallium half-cell to the copper half-cell

b. When students conducting an experiment mixed aqueous solutions of potassium iodide and iron(III) sulfate in a test tube, they observed that a reaction had occurred.

i. Identify the oxidant in this reaction. 1 mark

Fe^{3+}(aq)

$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.771
$I_2 + 2e^- \rightarrow 2I^-$	+0.535

ii. Write an appropriate chemical equation, with states, to describe the reaction that has occurred. 1 mark

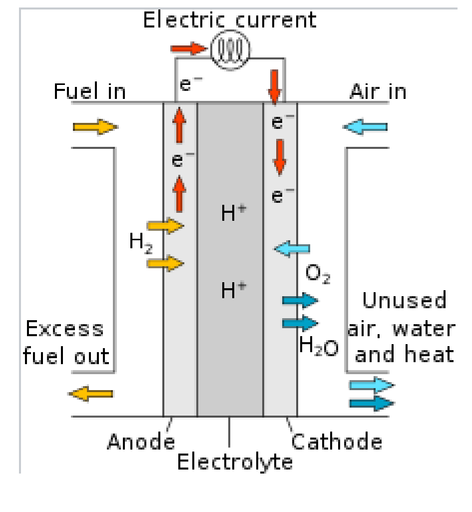
2Fe^{3+}(aq) + 2I^-(aq) -> I_2(s) + 2Fe^{2+}(aq)

6. A hydrogen-oxygen proton exchange membrane fuel cell operates at 65% efficiency.

a. Write appropriate chemical half-equations for the reaction occurring at the cathode 1 mark

Reaction occurring at the cathode

$O_2 + 4H^+ + 4e \rightarrow 2H_2O$



b. Calculate the volume of hydrogen gas, at 75°C and 120 kPa, that would be required for the fuel cell to produce 400 MJ of electrical energy. Be sure your answer is to the correct number of significant figures. 3 marks

Step 1 calculate the amount of energy needed to be supplied to produce 400MJ of electrical output.

=> $Energy_{input}$ X 0.65 = 400,000 kJ

=> $Energy_{input}$ = 615385 kJ

Step 2 Calculate the amount of mol of H_2 gas needed to deliver 615385 kJ of energy

=> 615385 / 282 (Molar heat of combustion of H_2) = 2182 mol

Step 3 Calulate the volume using PV=nRT

=> V = (2182 X 8.31 X 348) / 120 = 5.26 X 10^4 (3 sig figs)

On the right is what VCAA states about significant figures

c. Explain one major issue associated with the use of

hydrogen as a fuel. 1 mark

Any one of the points below.

- expensive infrastructure due to high pressures and low temperatures required for its storage.

-industrial quantities of hydrogen are obtained from fossil fuels through a process known as steam reformation.

$CH_4(g) + 2\ H_2O_{(g)} \rightarrow CO_2(g) + 4H_2(g)$

CO_2 is a product of the formation of H_2.

Non-zero digits in data are always considered significant. Leading zeros are never significant whereas following zeros and zeros between non-zero digits are always significant. For example, 075.0210 contains six significant figures with the zero at the beginning not considered significant. A whole number may be a counting number or a measurement and determination of significant figures varies in the literature. For the purpose of the *VCE Chemistry Study Design*, whole numbers will have the same significant figures as number of digits, for example 400 has three significant figures while 400.0 has four.

Using a significant figures approach, one can infer the claimed accuracy of a value. For example, 400 is closer to 400 than 399 or 401. Similarly 0.0675 is closer to 0.0675 than 0.0674 or 0.0676.

Columns of data in tables should have the same number of decimal places, for example, measurements of lengths in centimetres or time intervals in seconds may yield the following data: 5.6, 9.2, 11.2 and 14.5. Significant figure rules should then be applied in subsequent data analysis.

Calculations in chemistry often involve numbers having different numbers of significant figures. In mathematical operations involving:

- addition and subtraction, the student should retain as many digits to the right of the decimal as in the number with the fewest significant digits to the right of the decimal, for example: 386.38 + 793.354 - 0.000397 = 1179.73
- multiplication and division, the student should retain as many significant digits as in the number with the fewest significant digits, for example: 326.95 x 10.2 ÷ 20.322 = 164.

Intermediate results in calculations should retain at least one significant figure more than such analysis suggests until the final result is ascertained.

7. An experimental rechargeable galvanic cell is being trialled that uses lithium metal and sulphur as reactants. The overall equation, states not shown, for this cell is shown below.

$$16Li + S_8 \rightarrow 8Li_2S$$

A polymer electrolyte is used rather than an aqueous electrolyte. This cell produces 2.4 volts, is relatively cheap and is light in weight.

a) Give an explanation as to why an aqueous solution is not used in this cell. Provide a chemical equation to support your explanation.

A spontaneous reaction will occur between water and Li metal. The overall reaction produces hydrogen gas which is explosive.

$2Li(s) + 2H_2O(l) \rightarrow H_2(g) + 2Li^+(aq) + 2OH^-(aq)$

$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$	−1.18
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	−2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$	−2.87
$K^+(aq) + e^- \rightleftharpoons K(s)$	−2.93
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	−3.04

spontaneous reaction

b) Write balanced half-equations for the reactions occurring at the cell is discharging:

i. Anode *$Li \rightarrow Li^+ + e$*

ii. Cathode *$16e + S_8 \rightarrow 8S^{-2}$*

c) Write the balanced half-equation of the reaction occurring at the cathode during recharging.

Reduction always occurs at the cathode whether the cell is being recharged or discharged. Since all half-reactions are reversed at their respective electrodes during recharging the reaction that now takes place at the cathode, that used to be the anode when the cell was discharging, is given below.

$Li^+ + e \rightarrow Li$