CONTOUREDUCATION

Website: contoureducation.com.au | **Phone:** 1800 888 300
Email: hello@contoureducation.com.au

VCE Chemistry ¾
AOS 1 Revision II [1.12]
Workbook

Outline:

Section A: [1.6-1.7] - Introduction to Redox & Spontaneous Redox Reactions (45 Marks)

Sub-Section: Recap

Cheat Sheet

[1.6.1] - Apply Oxidation Numbers to Find Oxidant & Reductant

- Redox reactions must occur in pairs.
- Redox is the exchange of electrons - one substance gives away electrons, the other substance takes in electrons.

Oxidation Reaction	Reduction Reaction
Electrons are [Gained] / [Lost].	Electrons are [Gained] / [Lost].
Oxidation Number [Increases] / [Decreases].	Oxidation Number [Increases] / [Decreases].

- Oxidation Number Rules
  - Isolated Elements (e.g., H_2): 0
  - Ions (e.g., Na^+): charge (+1)
  - Ionic Compounds (e.g., NaCl): periodic table
  - Oxygen (O): -2
  - Hydrogen (H): +1
  - Sum of oxidation numbers in compound is equal to (e.g., H_2SO_4 or MnO_4^-): charge
- Oxidant: Causes [reduction] / [oxidation] to other species, itself undergoes [reduction] / [oxidation].
- Reductant: Causes [reduction] / [oxidation] to other species, itself undergoes [reduction] / [oxidation].
- In conjugate redox pairs, the [oxidant] / [reductant] is always written first.

oxidant / reductant

%

[1.6.2] - Apply KOHES to Write Balanced Half-Equations and Overall Equations in Acidic & Basic Conditions

- Balancing Equation Steps:
  - Balanced the K.
  - Balanced the O by adding H_2O.
  - Balanced the H by adding H^+.
  - Balanced the E by adding ______.
  - Included the S.
- Acronym: KOHES
- Balancing in Basic Conditions:
  1. Balance in acidic conditions first using KOHES.
  2. neutralise hydrogen ions (H^+) by adding OH^- to both sides.
- Number of electrons lost/gained should align with change in oxidation no.
- Forming Overall Equation: Cancel out electrons by finding LCM force.

$(Cu \rightarrow Cu^{2+} + 2e^-) \times 5$

$(MnO_4^- + 8H^+ + \boxed{5e^-} \rightarrow Mn^{2+} + 4H_2O) \times 2$

+7 → +2

-5

37/100

%

Cheat Sheet

H^+ SO_4^{2-}
H_2SO_4

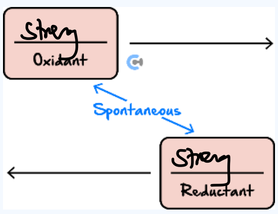
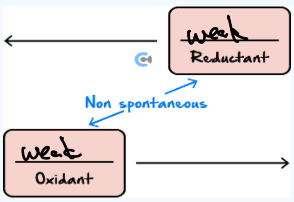
[1.7.1] – Apply the ECS to Predict Spontaneous Reactions

- Net Ionic Equation Definition: A balanced full equation with spectators omitted.
- **Spectator Ion:** Compound which is present but does not participate in reaction.

Reduction Reaction	Oxidation Reaction
[forward] / [reverse] reaction on ECS	[forward] / [reverse] reaction on ECS

Oxidants	Reductants
Positioned on the [left] / [right] side	Positioned on the [left] / [right] side

Strongest Oxidants	Strongest Reductants
Positioned [top] / [bottom] – [left] / [right]	Positioned [top] / [bottom] – [left] / [right]

Spontaneous Redox Reaction	Non-Spontaneous Redox Reaction
	

- Steps to predicting spontaneous reaction:
  1. Split all species into cations/anions. Some cations/anions are inert.
  2. Locate all species on the ECS. Draw a vertical line to splits oxidants and reductants apart.
  3. Draw a mini ECS.
  4. Find the strongest oxidant and strongest reductant.
  5. Check for downhill/negative gradient.
  6. Write out half-equations.
- When multiple oxidants/reductants are present, the strongest oxidant reacts with strongest reductant.
- The four ions which appear on both sides of the electrochemical series:
  - Fe^{2+}
  - Sn^{2+}
  - Mn^{2+}
  - Pb^{2+}

Definition

Learning Objective: [1.7.2] - Identify Differences Between Direct & Indirect Redox Reactions, & Features of ECS

- Standard Electrode Potential Definition: Method to measure EMF.
- Standard Hydrogen Electrode (SHE): $H^+(aq)/H_2(g)$ which has E^0 = 0 V.
- The electrochemical series does not predict the rate of reaction.

Direct Contact Spontaneous Redox Reaction	Indirect Contact Spontaneous Redox Reaction
chemical → thermal	chemical → electrical

Definition

Learning Objective: [1.7.3] - Find the Strongest Oxidants/Reductants by Constructing Your Own ECS

- Electrochemical series ordered from [lowest → highest] / [highest → lowest] E^0 value.

Strongest Oxidant	Strongest Reductant
[highest] / [lowest] E^0 value	[highest] / [lowest] E^0 value

- Creating electrochemical series yourself steps:

1. Draw a vertical line to separate oxidants and reductants.
2. Using information, place oxidants/reductants on this mini electrochemical series.

Spontaneous Reactions	Non-Spontaneous Reactions
[positive] / [negative] gradient	[positive] / [negative] gradient

3. Write the conjugate version of the oxidant/reductant.
4. Repeat for each piece of information.

Space for Personal Notes

A^+ | A

B^{2+} | B

Let's Walkthrough Together!

Question 1 (4 marks) **[1.6.2] Walkthrough.**

Balance the following equation in **alkaline** conditions:

$$Ni^{2+}(aq) + I_2(aq) \rightarrow Ni(s) + IO_3^-(aq)$$

$(Ni^{2+} + 2e^- \rightarrow Ni) \times 5$

KOHES

$I_2 + \cancel{6H_2O} \rightarrow 2IO_3^- + \cancel{12H^+} + 10e^-$ $\quad 6 \cancel{12H_2O}$

$+12OH^- \quad \cancel{+12OH^-}$

$I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O + 10e^-$

$5Ni^{2+}_{(aq)} + I_{2(aq)} + 12OH^-_{(aq)} \rightarrow 5Ni_{(s)} + 2IO^-_{3(aq)} + 6H_2O_{(l)}$

Question 2 (3 marks) **[1.7.1] Walkthrough.**

Fe^{2+} Sn^{2+}

Some copper metal is dipped into a solution that contains copper nitrate, tin (II) nitrate, zinc chloride and aluminium bromide.

Find the overall reaction that takes place.

$Cu^{2+}_{(aq)} + Sn^{2+}_{(aq)} \rightarrow Cu_{(s)} + Sn^{4+}_{(aq)}$

Space for Personal Notes

Question 3 (2 marks) **[1.7.3] Walkthrough.**

There are three unknown substances, P, Q and R. The following half-equations are given, but their E^0 values are not given.

Reaction
$P^{2+}(aq) + 2e^- \rightleftharpoons P(s)$
$Q^{2+}(aq) + 2e^- \rightleftharpoons Q(s)$
$R^{2+}(aq) + 2e^- \rightleftharpoons R(s)$

It is known that when P is mixed into a solution of R^{2+}, a reaction begins to occur.

It is also known that when Q is mixed into a solution containing R^{2+}, no reaction occurs.

Rank the three metals in terms of their decreasing oxidant strength.

- Decreasing oxidant strength: Q^{2+}, R^{2+}, P^{2+}

Space for Personal Notes

Q^{2+} Q

R^{2+} R

P^{2+} P

Sub-Section: Question Set A

INSTRUCTION: 10 Marks. 10 Minutes Writing.

Question 4 (1 mark) **[1.6.1]**

The substance that gains electrons:

A. ~~Undergoes oxidation.~~

B. Is the reductant.

C. Decreases its oxidation number.

D. Gains oxygen.

Question 5 (1 mark)

Find the oxidation number for the specified element in the following compounds:

a. Chromium in CrO_4^-. (0.5 marks) **[1.6.1]**

$Cr + 4(-2) = -1$

$Cr = \boxed{+7}$

b. Sulphur in $S_2O_3^{2-}$. (0.5 marks) **[1.6.1]**

$2S + 3(-2) = -2$

$2S = +4$

$\boxed{S = +2}$

Space for Personal Notes

Question 6 (3 marks)

It is known that sulphate can react and turn into thiosulphate. KOHES

a. Complete the balanced half-equations in **acidic** conditions, and state whether it is a reduction or oxidation reaction. (2 marks) **[1.6.2]**

$2SO_4^{2-}{}_{(aq)} + 10H^+{}_{(aq)} + 8e^- \rightarrow S_2O_3^{2-}{}_{(aq)} + 5H_2O_{(l)}$

Type of Reaction: [Reduction] / [Oxidation]

b. Write the conjugate redox pair for the reaction. (1 mark) **[1.6.1]**

$SO_4^{2-}{}_{(aq)} / S_2O_3^{2-}{}_{(aq)}$

Question 7 (5 marks)

a. Complete the balanced half-equations in **alkaline** conditions, and state whether it is a reduction or oxidation reaction. States are not required.

i. Nitrate ions turning into nitrogen monoxide (NO). (2 marks) **[1.6.2]**

$2\cancel{4}H_2O$

$NO_3^- + \cancel{4H^+} + 3e^- \rightarrow NO + \cancel{2H_2O}$

$\cancel{+4OH^-}$ $+4OH^-$

$(NO_3^-{}_{(aq)} + 2H_2O_{(l)} + 3e^- \rightarrow NO_{(g)} + 4OH^-{}_{(aq)}) \times 2$

Type of Reaction: [Reduction]/[Oxidation]

ii. Copper turning into copper (II) ions. (1 mark) **[1.6.2]**

$Cu_{(s)} \rightarrow Cu^{2+}{}_{(aq)} + 2e^-$

Type of Reaction: [Reduction]/[Oxidation]

b. These two half-equations are combined to form an overall equation.

i. Write the balanced reaction for the overall reaction. (1 mark) **[1.6.2]**

$2NO_3^-{}_{(aq)} + 3Cu_{(s)} + 4H_2O_{(l)} \rightarrow 3Cu^{2+}{}_{(aq)} + 2NO_{(g)} + 8OH^-{}_{(aq)}$

ii. State the oxidant. (1 mark) **[1.6.1]**

$NO_3^-{}_{(aq)}$

Check off any learning objectives that obtained full marks from the "Contour Check" booklet!

Space for Personal Notes

Sub-Section: Question Set B

INSTRUCTION: 13 Marks. 13 Minutes Writing.

Question 8 (2 marks)

Lead (II) nitrate has potassium metal dipped inside of it.

a.

i. Write the reduction reaction that takes place. (0.5 marks) **[1.7.1]**

$Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$

ii. Write the oxidation reaction that takes place. (0.5 marks) **[1.7.1]**

$(K_{(s)} \rightarrow K^+_{(aq)} + e^-) \times 2$

b. Write the **full balanced ionic equation**. (1 mark) **[1.7.1]**

$Pb^{2+}_{(aq)} + 2K_{(s)} \rightarrow 2K^+_{(aq)} + Pb_{(s)}$

Space for Personal Notes

Question 9 (3 marks)

A mixture of nickel (II) nitrate, copper (II) nitrate and iron (II) nitrate are mixed together. A nickel coin is then dipped into the mixture.

a. Write the two half-equations that are expected to occur for the:

i. Reduction reaction. (1 mark) **[1.7.1]**

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

ii. Oxidation reaction. (1 mark) **[1.7.1]**

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

b. State the energy conversions that occur. (1 mark) **[1.7.2]**

chemical to thermal

Question 10 (1 mark) **[1.7.2]**

A tin rod is dipped into a solution containing 1.0 *M* copper (II) chloride at 25°C. A reaction is not observed to occur. Explain this observation.

ECS doesn't predict rate of reaction,
Reaction occurs very slowly

Space for Personal Notes

Question 11 (4 marks)

The following half-equations are given:

$$D^{3+}(aq) + 3e^- \rightleftharpoons D(s) \quad +1.19\,V$$

$$B^{2+}(aq) + 2e^- \rightleftharpoons B(s) \quad -0.85\,V$$

$$A^{2+}(aq) + 2e^- \rightleftharpoons A(s) \quad +1.29\,V$$

$$C^{+}(aq) + e^- \rightleftharpoons C(s) \quad +1.14\,V$$

A^{2+} | A
D^{3+} | D
C^{+} | C
B^{2+} | B

a. State the weakest oxidant and the weakest reductant. (2 marks) **[1.7.1]**

Weakest Oxidant	Weakest Reductant
$B^{2+}_{(aq)}$	$A_{(s)}$

b. A solution of $B^{2+}(aq)$ is mixed with some $C(s)$. Will a reaction occur? Explain why/why not, and if there is a reaction, write the overall reaction which takes place. (2 marks) **[1.7.1]**

No, B^{2+} is a weak oxidant, $C_{(s)}$ is a weak reductant relative to each other -

Space for Personal Notes

Question 12 (2 marks) **[1.7.3]**

Matthew is similarly given five metals and 1 M solution of nitrates of the metals.

The metals are labelled A, B, C, D and E, and the solutions are labelled $A^{2+}, B^{2+}, C^{2+}, D^{2+}$ and E^{2+}.

The student carried out several experiments and the results obtained are listed below.

- Metal A reacts with B^{2+} spontaneously.
- Metal C becomes coated with another metal when placed in each of the solutions A^{2+}, B^{2+}, D^{2+}, but not with E^{2+}.
- When metal A is dipped into a solution of D^{2+}, no reaction takes place.

Rank each of the 5 metals in order of increasing E^0 values.

E, C, D, A, B
(s) (s) (s) (s) (s)

B^{2+} B
A^{2+} $\boxed{A}$
D^{2+} D
C^{2+} C
E^{2+} E

Question 13 (1 mark) **[1.6.1] [1.7.1]**

The E^0 values for reduction reactions involving chromium are shown below.

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O$	$+1.33\ V$
$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	$-0.41\ V$
~~$Cr^{3+}(aq) + 3e^- \rightleftharpoons Cr(s)$~~	~~$-0.76\ V$~~
$Cr^{2+}(aq) + 2e^- \rightleftharpoons Cr(s)$	$-0.91\ V$

+6 +3 +3 +2 +2 0

Which of the following could be used to reduce chromium from an oxidation state of +6 to +3?

A. Cu

B. Al

C. Mn

D. Fe

Li^+ $Li_{(s)}$

Space for Personal Notes

Sub-Section: Additional

Question 14 (4 marks) **[1.6.2]**

Balance the overall reaction for the following equation.

$$6\,MnO_4^-(aq) + 5\,I^-(aq) + 18\,H^+(aq) \rightarrow 6\,Mn^{2+}(aq) + 5\,IO_3^-(aq) + 9\,H_2O(l)$$

Space for Personal Notes

Question 15 (9 marks)

Steel is made up mostly of iron and small amounts of cobalt and manganese.

a.

i. If a steel pan were to be filled with water, would a reaction be expected to occur? Justify your answer. (2 marks) **[1.7.1]**

Yes

ii. Why is a reaction often not observed in practice? (1 mark) **[1.7.2]**

Slow rate

b.

i. Now a mixture of $ZnSO_4$ and $Ni(NO_3)_2$ solutions are tossed into the steel pan. Write the overall equation taking place. (2 marks) **[1.7.1]**

$Ni^{2+} + Mn \rightarrow Ni + Mn^{2+}$

ii. If the $Ni(NO_3)_2$ runs out after some time, what would one expect to happen? If there is a new reaction, write the overall equation that would be taking place. (2 marks) **[1.7.1]**

$Zn^{2+} + Mn \rightarrow Zn + Mn^{2+}$

iii. Now if both $Ni(NO_3)_2$ and the Mn in the steel pan ran out, what would one expect to happen? If there is a new reaction, write the overall equation that would be taking place. (2 marks) **[1.7.1]**

No

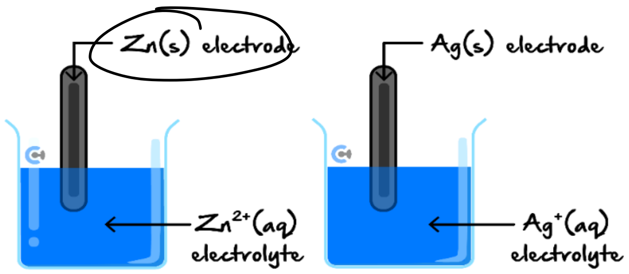
Section B: [1.8] - Galvanic Cells (23 Marks)

Sub-Section: Recap

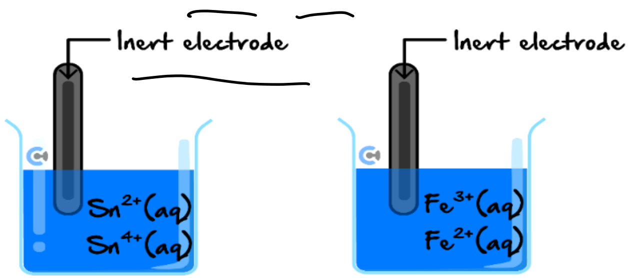
Cheat Sheet

[1.8.1] - Identify Electrodes, and Salt Bridge/Electron Movement During Galvanic

- Metal/Ion half-cell.
  - Example: $Zn^{2+}(aq)/Zn(s)$, $Ag^{+}(aq)/Ag(s)$

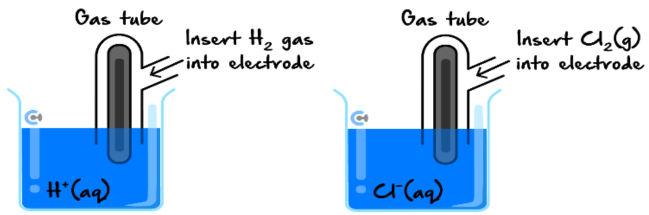


- Ion/Ion half-cell.
  - Example: $Sn^{4+}(aq)/Sn^{2+}(aq)$, $Fe^{3+}(aq)/Fe^{2+}(aq)$



  - Material of electrode: Pt(s), Au(s), graphite

- Gas/Ion half-cell.
  - Example: $H^{+}(aq)/H_2(g)$, $Cl_2(g)/Cl^{-}(aq)$



- Gas tube only required when gas is a [reactant] / [product].

Cathode	Anode
[reduction] / [oxidation] reaction	[reduction] / [oxidation] reaction
[positive] / [negative] charge	[positive] / [negative] charge
Acronym to Remember: RC⊕	Acronym to Remember: AO⊖

- Electron Flow Overall:

Anode	→	Cathode
[positive] / [negative]	e^- →	[positive] / [negative]
[reduction] / [oxidation] reaction [gains] / [loses] electrons Example: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$	e^-	[reduction] / [oxidation] reaction [gains] / [loses] electrons Example: $2H^{+}(aq) + 2e^- \rightarrow H_2(g)$

Cheat Sheet

- Internal Circuit (Salt Bridge)
  - Purpose: To balance the charges, to complete circuit.

Cations Move to	Anions Move to
[cathode] / [anode]	[cathode] / [anode]

  - **Properties of Salt Bridge:** inert, soluble
  - **Most common salt bridge used:** KNO_3

[1.8.2] - Write Reactions in Galvanic Cells & Calculate the Maximum EMF Produced

- Steps to draw galvanic cells.
- Identify the oxidants/reductants present.
- Find the strongest oxidant and strongest reductant.
- Write the half-equations that occur.
- Draw out the cells, and label the cathode and anode along with their respective polarities.
- Draw the flow of electrons and the ions in the salt bridge.
- Energy Conversions:

Direct Contact Spontaneous Redox Reaction	Indirect Contact Spontaneous Redox Reaction

- Electromotive Force (EMF) formula:

$$EMF = E^\circ(\text{oxidant}) - E^\circ(\text{reductant})$$

- **Conditions:** SLC & 1.0M.

- Creating electrochemical series, yourself steps:

1. Draw a vertical line to separate oxidants and reductants.
2. Using information, place oxidants/reductants on this mini electrochemical series.

Spontaneous Reactions	Non-Spontaneous Reactions
[positive] / [negative] gradient	[positive] / [negative] gradient

3. Write the conjugate version of the oxidant/reductant.
4. Repeat for each piece of information.

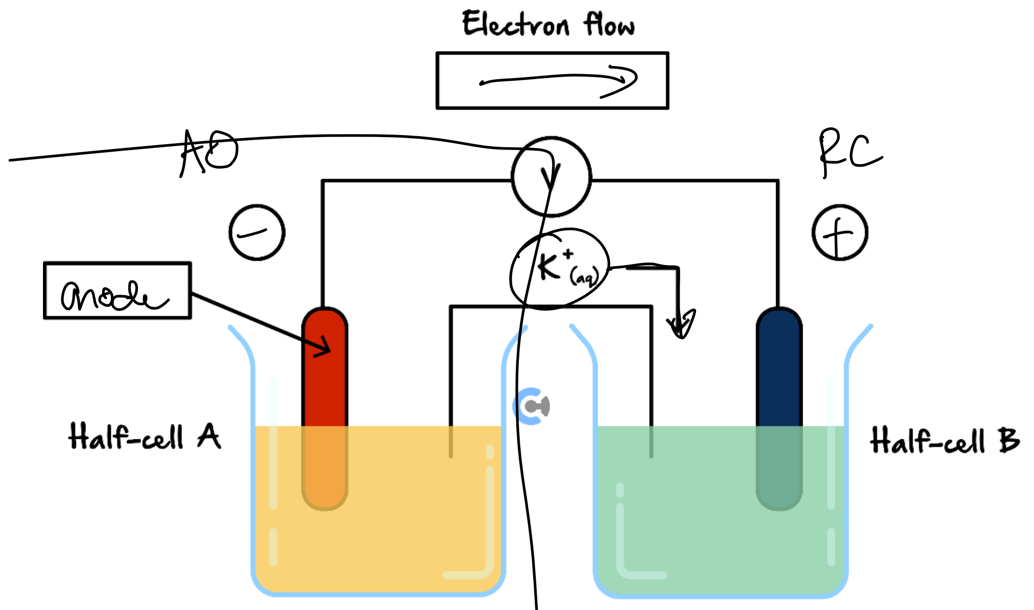
[1.8.3] – Identify & Explain Observations During the Operation of Galvanic Cells

- **Four types of observations made:**
  - Change in size/mass of the electrodes (look out for **solid** reactants/products in half-equations).
  - Change in pH (look out for H^+ or OH^- ions in half-equations).
  - bubbles being formed (look out for **gases** being formed).
  - **Change in** color (look out for **ions** being produced/consumed in solution).
- **Procedure**: Write out half-equations first with RC+ and AO−, then figure out observations.

Let's Walkthrough Together!

Question 16 (8 marks) **Walkthrough.**

A galvanic cell containing the half-cells $Fe^{3+}(aq)/Fe^{2+}(aq)$ and $Fe^{2+}(aq)/Fe(s)$ are constructed.



$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$

$Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$

a. In the spaces provided above, write the half-equations that occur at each half-cell. (2 marks) **[1.8.2]**

b. Label the direction of electron flow in the box provided above, and label the polarities of the electrodes in the circles provided above. (1 mark) **[1.8.1]**

c. Write whether the electrode in a half-cell *A* is the cathode or anode in the box provided above. (1 mark) **[1.8.1]**

d. As the cell reacts, a colour change is seen in the electrolyte of a half-cell *B*. State the colour change that would be observed. (1 mark) **[1.8.3]**

yellow/brown → pale green

e. List two observations that can be made in half-cell *A*. (2 marks) **[1.8.3]**

decrease in size of electrode

electrolyte turns more intensely green

f. Find the EMF of this cell. (1 mark) **[1.8.2]**

$0.77 - (-0.44) = 1.21\text{ V}$

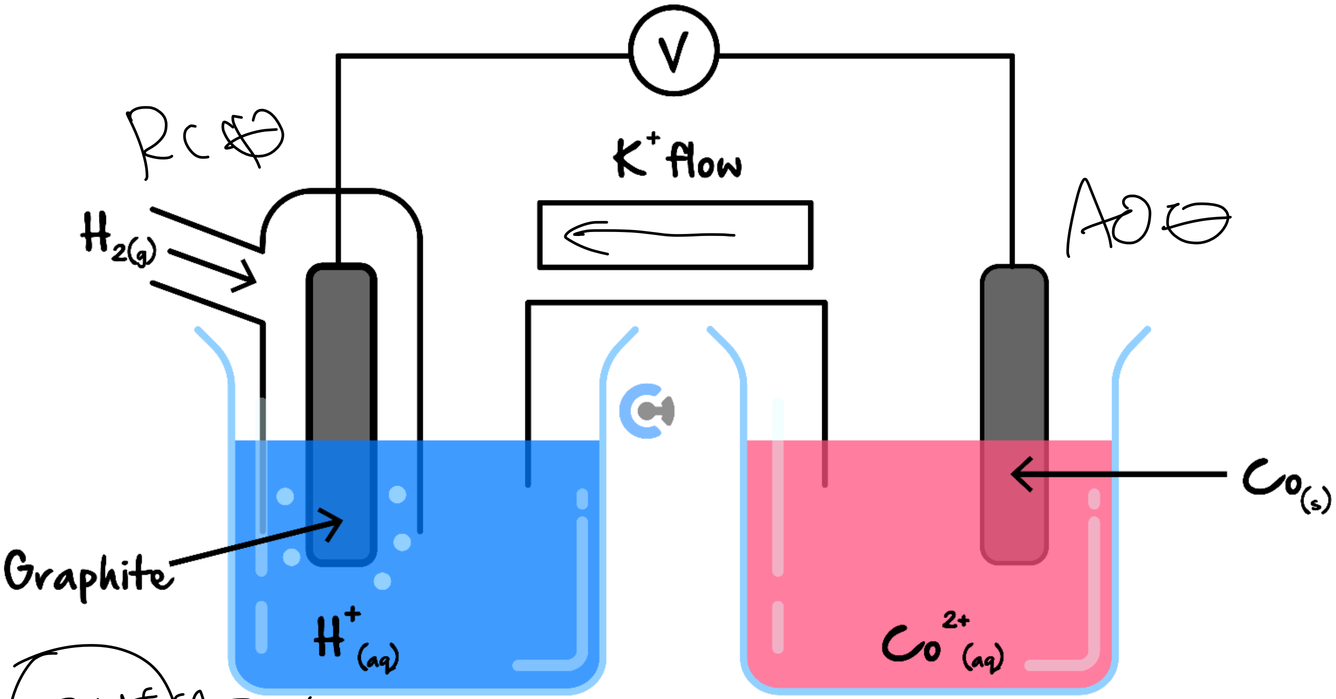
Space for Personal Notes

Sub-Section: Questions

INSTRUCTION: 10 Marks. 10 Minutes Writing.

Question 17 (3 marks)

The following galvanic cell is provided which contains the half-cells $Co^{2+}(aq)/Co(s)$ and $H^+(aq)/H_2(g)$.



$2H^+ + 2e^- \rightarrow H_2$

a. In the box in the diagram above, label the direction of the flow of potassium ions. (1 mark) **[1.8.1]**

b. Write the balanced half-equation for the reaction that occurs at the negative electrode. (1 mark) **[1.8.2]**

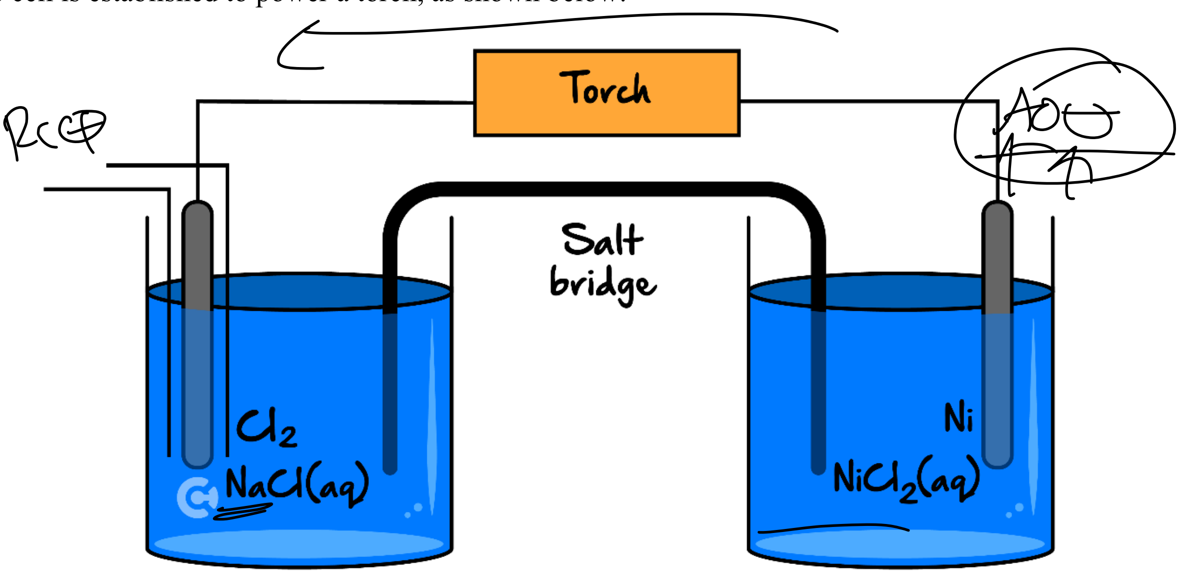
$Co_{(s)} \rightarrow Co^{2+}_{(aq)} + 2e^-$

c. A pH meter is inserted in the left half-cell. State the pH change that is expected to occur. (1 mark) **[1.8.3]**

pH increases

Question 18 (3 marks)

A galvanic cell is established to power a torch, as shown below.



a. Write a balanced equation for the overall reaction that will occur. (1 mark) **[1.8.2]**

$Cl_2 + 2e^- \rightarrow 2Cl^-$ $Ni \rightarrow Ni^{2+} + 2e^-$

$Cl_{2(g)} + Ni_{(s)} \rightarrow \cancel{2Cl^-_{(aq)} + Ni^{2+}_{(aq)}} + NiCl_{2(aq)}$

b. Explain the purpose of the salt bridge. (1 mark) **[1.8.1]**

to complete circuit

c. For this cell, the: (1 mark) **[1.8.1] [1.8.3]**

~~**A.** Electrons will flow from the chlorine to the nickel.~~

B. Nickel electrode will be the negative anode.

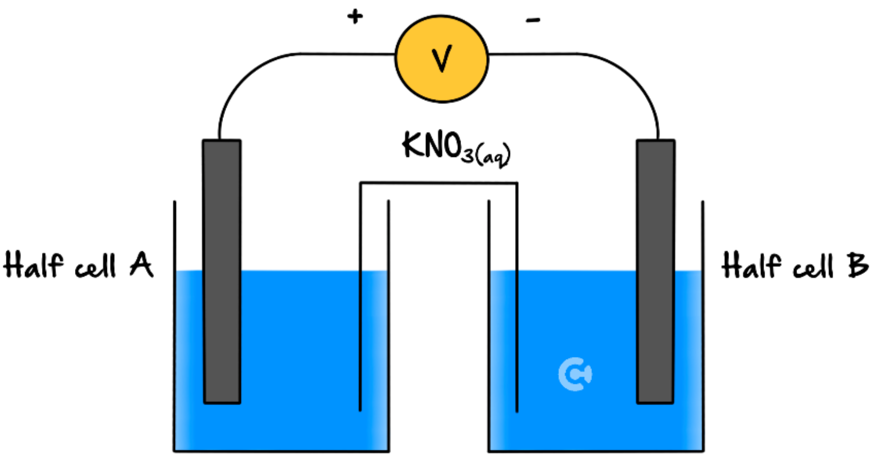
C. Concentration of nickel ions in the solution will be falling.

D. Chlorine electrode will be the negative cathode.

Space for Personal Notes

Question 19 (1 mark)

A galvanic cell was constructed using standard conditions as shown below.



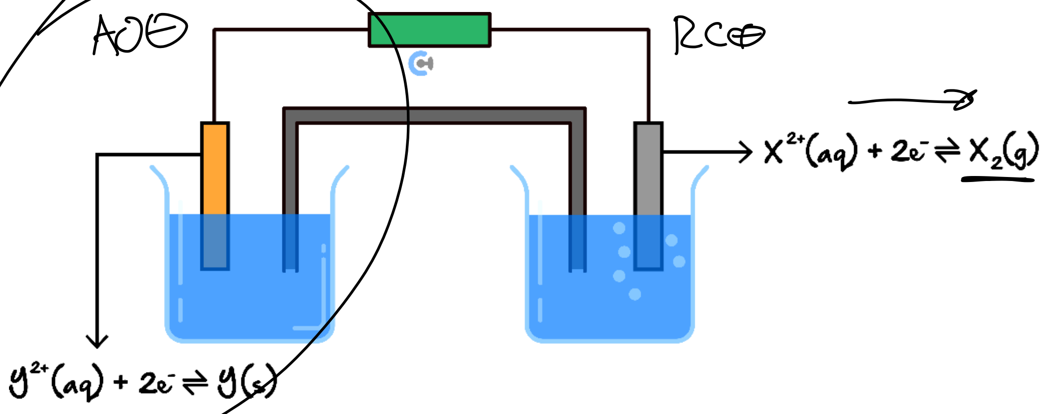
Each half-cell is made up of an inert electrode and an aqueous solution. The reading on the voltmeter is 1.00 V.
The half-cells could consist of:

	Half Cell A	Half Cell B
A.	$H_2O_2(aq)$	$Ni^{2+}(aq), Ni(s)$
B.	$Ni^{2+}(aq), Ni(s)$	$Fe^{3+}(aq), Fe^{2+}(aq)$
C.	$H_2O_2(aq), H^+(aq)$	$Fe^{3+}(aq), Fe^{2+}(aq)$
D.	$H_2O_2(aq)$	$Fe^{3+}(aq), Fe^{2+}(aq)$

Space for Personal Notes

Question 20 (1 mark) **[1.8.1] [1.8.3]**

Consider the following galvanic cell constructed as below during operation:



Select the correct option.

	Anode	Electron Flow
A.	X^{2+}/X_2	$X \rightarrow Y$
B.	X^{2+}/X_2	$Y \rightarrow X$
C.	Y^{2+}/Y	$X \rightarrow Y$
D.	Y^{2+}/Y	$Y \rightarrow X$

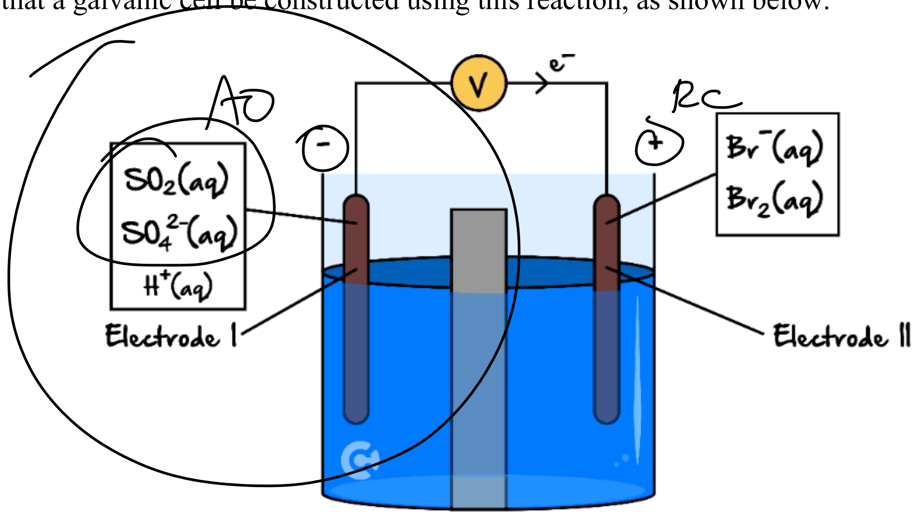
Space for Personal Notes

Question 21 (2 marks)

The chemical reaction represented below proceeds readily and rapidly in aqueous solution at room temperature.

$$Br_2(aq) + SO_2(aq) + 2H_2O(l) \rightarrow 2Br^-(aq) + 4H^+(aq) + SO_4^{2-}(aq)$$

It is proposed that a galvanic cell be constructed using this reaction, as shown below.



a. When this cell is operating, electrode I will be: (1 mark) **[1.8.1] [1.8.2]**

A. Anode, at which SO_2 undergoes oxidation.

B. Anode, at which Br_2 undergoes oxidation.

C. ~~Cathode~~, at which SO_2 undergoes reduction.

D. ~~Cathode~~, at which Br_2 undergoes reduction.

b. When the cell is operating, the pH close to electrode I would initially be expected to: (1 mark) **[1.8.3]**

A. Increase.

B. Decrease.

C. Remain unchanged.

D. Fluctuate rapidly.

Check off any learning objectives that obtained full marks from the "Contour Check" booklet!

Sub-Section: Additional

Question 22 (5 marks)

In a problem-solving exercise, a student was provided with four half-cells under standard conditions. The objective was to place the half-reactions in the correct order in an electrochemical series. The half-cell reduction reactions, in random order, are shown in the table below.

Half-cells	Reduction half-equations
A	$Cr^{2+}(aq) + 2e^- \rightleftharpoons Cr(s)$
B	$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightleftharpoons NO(g) + 2H_2O(l)$
C	$Au^{3+}(aq) + 3e^- \rightleftharpoons Au(s)$
D	$MnO_4^-(aq) + e^- \rightleftharpoons MnO_4^{2-}(aq)$

In a series of experiments, two half-cells were connected at a time and experimental observations were made as follows:

Half-cells	Experimental observations
A and B	Gas bubbles are produced at one electrode.
C and D	The Gold electrode increases in mass.
B and C	The pH near one electrode decreases.

a.

i. Give the formula of an oxidant from the experiment which is known to be stronger than acidified $NO_3^-(aq)$. (1 mark)

$Au^{3+}(aq)$

ii. Give the formula of a reductant from the experiment that is known to be a stronger reductant than $NO(g)$. (1 mark)

$Cr(s)$

b. Write the balanced equation for the cell reaction produced by connecting half-cells *A* and *B*. (1 mark)

$2NO_3^- + 8H^+ + 3Cu \rightarrow 2NO + 4H_2O + 3Cu^{2+}$

c. The standard hydrogen electrode (SHE) was not used in the problem-solving exercise.

Explain how the SHE could have been used in this exercise to gain further information about the correct order of the half-reactions in an electrochemical series. (2 marks)

normal flipped

electrolysis rates/equil

Space for Personal Notes

Section C: [1.9] - Fuel Cells (22 Marks)

Sub-Section: Recap

Cheat Sheet

[1.9.1] - Write Fuel Cell Half & Overall Reactions in Acidic Conditions

- Reactant is a fuel.
- Overall Reaction: Usually a combustion reaction.

- **Hydrogen** Fuel Cell Equations (acidic and basic) in ECS: *(Label Below.)*

Reaction	Standard electrode potential ($E°$) in volts at 25℃
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	$+2.87$
$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$	$+1.77$
$MnO_4^-(aq) + 8H^+ + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	$+1.51$
$PbO_2(aq) + 4H^+ + 2e^- \rightleftharpoons Pb^{2+}(aq) + 2H_2O(l)$	$+1.47$
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	$+1.36$
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	$+1.36$
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	$+1.23$
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	$+1.09$
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	$+0.80$
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	$+0.77$
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$	$+0.68$
$I_2(s) + e^- \rightleftharpoons 2I^-(aq)$	$+0.54$
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	$+0.40$
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	$+0.34$
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	$+0.15$
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	-0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	-0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	-0.25
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	-0.28
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	-0.44
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	-0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	-0.83
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$	-1.18
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	-1.66

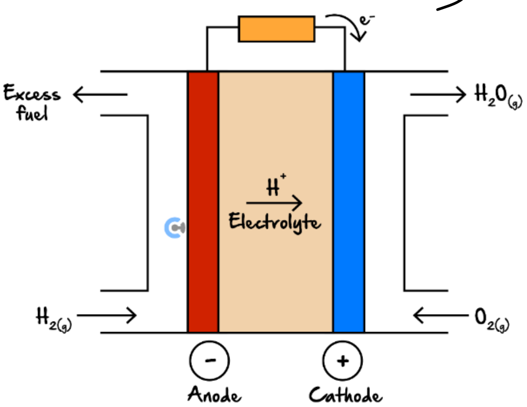
Maximum theoretical **EMF** produced: $+1.23\ V$

Cheat Sheet

- Balancing Half-Equations for Carbon Containing Fuels:
  - Unless otherwise specified, assume carbon-containing fuel oxidises into CO_2.
  - Method used to balance: KOHES
- All fuel cells have:
  - O_2: [reduces] / [oxidises] at the [cathode] / [anode].
  - Fuel: [reduces] / [oxidises] at the [cathode] / [anode].
- Checking Overall Equation: LCM ↔ stoich.
  - The electrolyte should always cancel out.

[1.9.2] - Identify Key Features of Fuel Cell Including Continuous Supply, Electrolyte Movement and Properties of Electrodes (PICCY)

- **Key Characteristic:** continuous supply



- **Energy Conversion:** chem → electric
- **Electrolyte Purpose**: Similar to salt bridge - to balance the charge.
- **Electrolyte Movement**: cations → [cathode] / [anode], anions → [cathode] / [anode].
- **Electron Flow:** anode → cathode.
- **Properties of Electrodes Acronym:** Stands for... *(Label Below.)*

porous, inert, catalyse, conduct electricity

PICCY

- Reasons for porous electrodes:
  - ↑SA → ↑rate
  - → ↑ eff of cell

[1.9.3] - Explain the Advantages & Disadvantages of Fuel Cells with Reference to Green Chemistry Principles

- **Energy efficiency and pollution in a fuel cell Sample Response**
  - direct energy conversion from chemical → electrical energy.
  - [more] / [less] / [same] energy loss, [more] / [less] / [same] energy efficient.
  - [more] / [less] / [same] fuel is required in a fuel cell to produce the [more] / [less] / [same] amount of energy.
  - Less overall CO_2 emissions overall.

Advantages	Disadvantages
Main Advantage: More energy efficient, less fuel required	expensive (PICCY)
quiet	
low start/stop time	H_2 (safety/storage)

- Green Chemistry Principles:
  - use of renewable feedstocks
  - design for energy eff

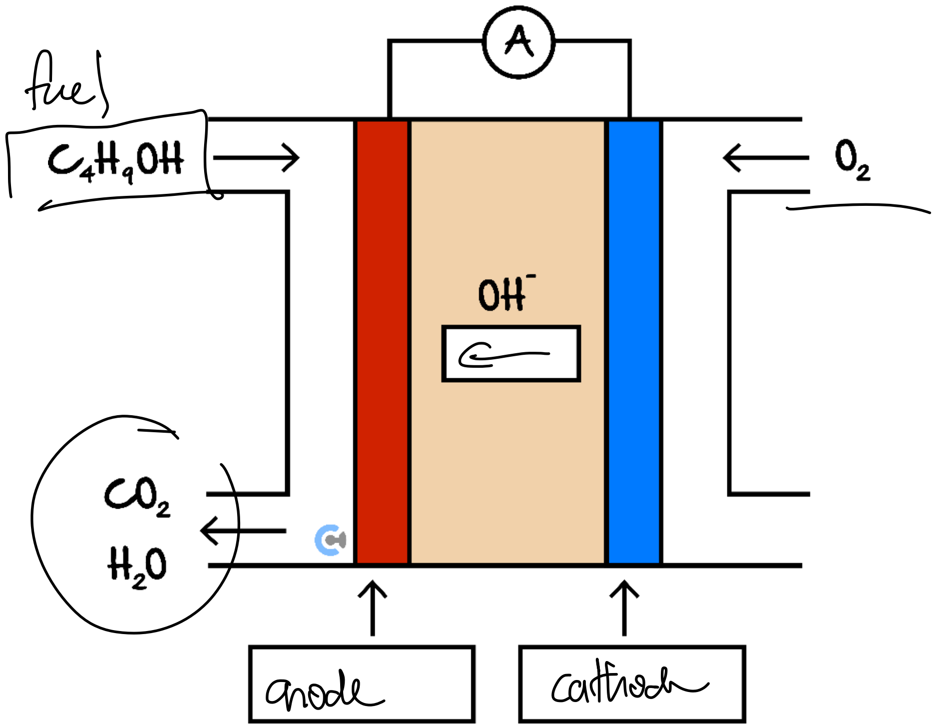
[1.9.4] - Write Fuel Cell Equations in Non-Acidic Conditions

- Steps to Balance in Alkaline/Basic Conditions:
  - Balance in acidic conditions using acidic.
  - Balance equation in basic conditions by adding OH^- to both sides.

Let's __Walkthrough__ Together!

Question 23 (5 marks) **Walkthrough.**

The following fuel cell is considered, including butanol as the fuel which produces carbon dioxide and water at SLC.



a. Label the cathode and anode in the boxes above. (1 mark) **[1.9.2]**

b. Label the direction of movement of the electrolyte above. (1 mark) **[1.9.2]**

CHO

c. Write the balanced equation for the overall reaction which takes place. (1 mark) **[1.9.1]**

$C_4H_9OH_{(l)} + 6O_{2(g)} \rightarrow 4CO_{2(g)} + 5H_2O_{(l)}$

d. Write the balanced half-equation for the reaction taking place at the:

i. Cathode. (1 mark) **[1.9.1]**

$O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)}$

//

KOHES 17

$\cancel{24}H_2O$

ii. Anode. (1 mark) **[1.9.1]**

$C_4H_9OH + \cancel{7H_2O} \rightarrow 4CO_2 + 2\cancel{4}H^+ + 24e^-$

$+24OH^- \qquad \cancel{+24OH^-}$

$C_4H_9OH_{(l)} + 24OH^-_{(aq)} \rightarrow 4CO_{2(g)} + 17H_2O_{(l)} + 24e^-$

Space for Personal Notes

Sub-Section: Questions

INSTRUCTION: 15 Marks. 15 Minutes Writing.

Question 24 (1 mark) **[1.9.2]**

The range of energy transformations that occur in devices includes the following:

I Chemical energy to electrical energy. 70%

~~II~~ Electrical energy to chemical energy.

III Chemical energy to heat energy.

Which of these transformations occur in a fuel cell?

A. I only.

B. I and II only.

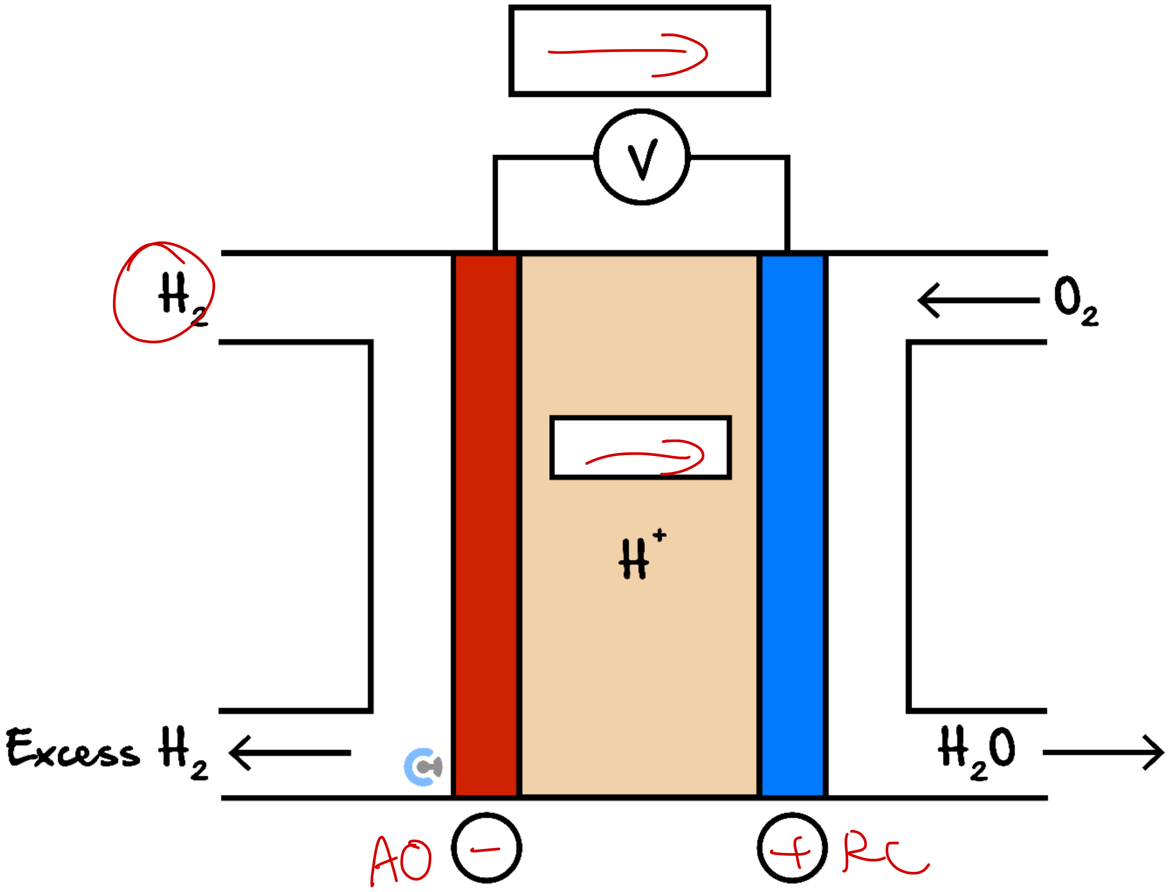
C. I and III only.

D. I, II and III.

Space for Personal Notes

Question 25 (9 marks)

Use the diagram of the hydrogen fuel cell below for the following questions.



a.

i. Label the polarity of the electrodes. (1 mark) **[1.9.2]**

ii. Draw arrows in the boxes provided to indicate the direction of: ion flow **and** electron flow. (1 mark) **[1.9.2]**

b. Write the two half-equations occurring at each electrode: (2 marks) **[1.9.1]**

Cathode: $O_{2(g)} + 4H^+_{(aq)} + 4e^- \rightarrow 2H_2O_{(l)}$

Anode: $H_{2(g)} \rightarrow 2H^+_{(aq)} + 2e^-$

c. State the EMF produced by this cell. (1 mark) **[1.9.1]**

+1.23V

d. Fuel cells are compared to gas-fired power stations.

i. State **one** major advantage of using hydrogen in a fuel cell rather than a gas-fired power station with reference to **one** green chemistry principle. Justify your answer. Refer to Item **26. ii.** of the Data Book. (3 marks) **[1.9.3]**

1

Fuel cells have a direct energy conversion from chemical to electrical, whereas gas-fried power station has multiple energy conversions, as a result, fuel cells have less energy loss, and thus has greater energy efficiency.

1

This relates to the green chemistry principle of "Design for energy efficient."

1

As a result, to produce the same amount of energy as in a gas-fired power station, because fuel cell is more efficient, not as much fuel is required.

As less fuel is reacted, less pollutants are produced, thereby minimising negative environmental impact.

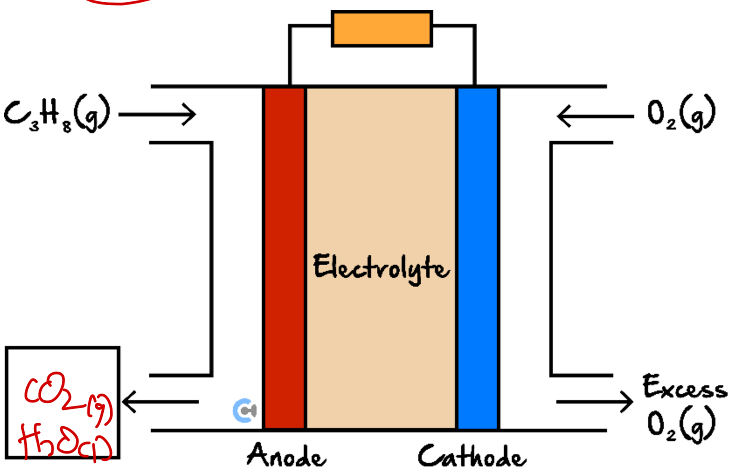
ii. State another advantage of a fuel cell compared to a gas-fired power station, other than the one discussed in **part e. ii.** (1 mark) **[1.9.3]**

quiet

Space for Personal Notes

Question 26 (5 marks)

Propane is often used as a fuel in **alkaline** fuel cells. The diagram below depicts the fuel cell in operation at SLC.



a. Write the balanced redox equation for:

i. The overall complete combustion reaction. (1 mark) **[1.9.1]**

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

ii. The half-reaction occurring at the anode. (2 marks) **[1.9.4]**

14
$\cancel{20}H_2O$

$C_3H_8 + \cancel{6H_2O} \rightarrow 3CO_2 + \cancel{20H^+} + 20e^-$

$+20OH^-$ $\qquad$ $+\cancel{20OH^-}$

$C_3H_8(g) + 20OH^-(aq) \rightarrow 3CO_2(g) + 14H_2O(l) + 20e^-$

b. In the diagram above, label the expected product(s) at the anode. (1 mark) **[1.9.1]**

c. Electrodes in fuel cells are typically porous. State **one** reason for this. (1 mark) **[1.9.2]**

Increase surface area ⇒ increases cell efficiency.

Check off any learning objectives obtained full marks from the "Contour Check" booklet!

Sub-Section: Additional

Question 27 (1 mark)

Fuel cells do not store energy; they convert energy directly and continuously to electrical energy. A fuel cell may be constructed using methanol and oxygen contained in an alkaline electrolyte. The overall reaction for the methanol/oxygen fuel cell is:

$$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$$

If the reaction at the cathode is $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$, then the reaction at the anode would be:

A. $CH_3OH(l) + H_2O(l) \rightarrow CO_2(g) + 6H^+(aq) + 6e^-$

B. $CH_3OH(l) + 6OH^-(aq) \rightarrow CO_2(g) + 5H_2O(aq) + 6e^-$

C. $CH_3OH(l) + 3O_2(g) \rightarrow CO_2(g) + 3H_2O(l)$

D. $CH_3OH(l) \rightarrow CO(g) + 4H^+(aq) + 4e^-$

Question 28 (1 mark)

Which statement best describes an advantage of fuel cells over traditional galvanic cells?

A. Fuel cells have a finite lifespan due to electrode degradation.

B. Fuel cells can operate indefinitely as long as fuel and oxidant are supplied.

C. Fuel cells generate more waste products.

D. Fuel cells are less efficient at converting chemical energy to electrical energy.

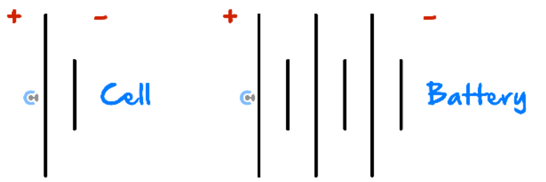
Space for Personal Notes

Section D: [1.10] - Primary Cells & Faraday's Laws (23 Marks)

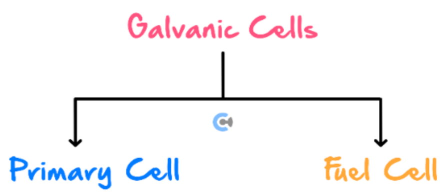
Sub-Section: Recap

Cheat Sheet

[1.10.1] – Identify Features of Primary Cells & How They Operate



- **Primary Cell:** Galvanic cell which produces electrical energy and cannot be recharged.
  - Electrolyte purpose: Allows for the movement of ions to balance charges around the electrodes.
  - Provides the internal circuit in the cell.
- Comparison of Primary & Fuel Cells:



Type of Cell	Primary Cells	Fuel Cells
Type of Reaction	[Spontaneous]/[Non-Spontaneous]	[Spontaneous]/[Non-Spontaneous]
Reactants	[Stored in each half-cell]/ [Continuous supply of reactants]	[Stored in each half-cell]/ [Continuous supply of reactants]
Electrodes	CY	Pt CY
Reactivity of Electrodes	Electrodes [may]/[will not] participate in the chemical reaction.	Electrodes [may]/[will not] participate in the chemical reaction.
Price	[Cheap]/[Expensive]	[Cheap]/[Expensive]
Electrolyte	[Each half-cell has its electrolyte]/ [shared electrolyte]	[Each half-cell has its electrolyte]/ [shared electrolyte]
Energy Efficiency	[High]/[Low] Energy Efficiency	[High]/[Low] Energy Efficiency

Cheat Sheet

[1.10.2] – Apply Faraday's First & Second Law and $Q = It$ & $Q = n(e^-)F$ to Calculations

- Electric Charge (Q): The amount of charged particles (e.g., electrons) which are present, measured in Coulombs (C).
- Electric Current (I): The rate of flow of charged particles per second, measured in amperes (A).
- Electric Charge Formula:

$$Q = It$$

- Faraday's First Law of Electrochemistry:

"The amount of chemical change being produced by a current at an electrode-electrolyte interface is proportional to the quantity of electricity used."

  - **Relationship:** $Q \propto m$
  - **Law Simplified:** As more electrical charge passes through the cell, the amount of substances that react/are produced **[increases]/[decreases]/[same].**

- $F =$ 96500 $C\,mol^{-1}$.
- Faraday's Constant Formula:

$$Q = n(e^-)F$$

- **Faraday's Second Law:**

"If the same amount of electricity is passed through different electrolytes, the masses of ions deposited at the electrodes are directly proportional to their chemical equivalents."

  - **Law Simplified**: Consider the stoichiometric ratio – $n(e^-) : n(metal)$

- Steps for calculation:
  1. Write any half-equation.
  2. $Q = It$
  3. $n(e^-) = \frac{Q}{F}$
  4. Stoich Ratios: $n(metal) : n(e^-)$
  5. $m(metal) = n \times M_r$

[1.10.3] – Calculate the Charge of a Metal

- Steps to find the charge of ion:
  1. Find $n(metal)$ by $n(metal) = \frac{m}{M_r}$
  2. Find $n(e^-)$ by $n(e^-) = \frac{Q}{F}$
  3. Find the stoichiometric ratio by dividing by the smallest no. present.

$n(metal) : n(e^-)$

Let's Walkthrough Together!

$Q=It$ $n(e^-)=\frac{Q}{F}$

Question 29 (6 marks) **Walkthrough.**

A current of 7.50 A is passed through for 2.50 minutes in a $Cu^{2+}(aq)/Cu(s)$ and a $Ni^{2+}(aq)/Ni(s)$ galvanic cell.

I t

a. Find the change in mass of the anode. (4 marks) **[1.10.2]**

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

$Q=It=7.5A \times 2.5 \times 60s = 1125C$

$n(e^-)=\frac{Q}{F}=\frac{1125C}{96500C/mol}=0.0117mol$

$n(Ni)=\frac{1}{2}n(e^-)=0.00583mol$

$m(Ni)=n \times Mr = 0.00583 \times 58.7 = 0.342g$ → $\boxed{\text{decrease by } 0.342g}$

b. Find the amount of electrical energy which is produced. (2 marks) **[1.10.2]**

$E=VIt$ $EMF=0.34-(-0.25)=0.59V$

$E=0.59V \times 7.5A \times 2.5 \times 60s = \boxed{664J}$

Space for Personal Notes

Sub-Section: Questions

INSTRUCTION: 10 Marks. 10 Minutes Writing.

Question 30 (5 marks)

Crystal measures that after letting a galvanic cell operate for half an hour, 5.0 *g* of solid lead was evolved at one of the electrodes.

a. Write the appropriate balanced half-equation aligning with her observations. (1 mark) **[1.8.1]**

$Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$

b. Calculate the amount, in *mol*, of electrons produced in this cell. (2 marks) **[1.10.2]**

$n(Pb) = \frac{m}{M} = \frac{5}{207.2} = 0.024$ mol

$n(e^-) = 2n(Pb) = 0.048$ mol

c. Hence, calculate the current generated by this cell. (2 marks) **[1.10.2]**

$Q = n(e^-)F = 0.048 \times 96500 = 4658$ C

$I = \frac{Q}{t} = \frac{4658\ C}{30 \times 60} = 2.58$ A

Space for Personal Notes

Question 31 (1 mark) **[1.10.3]**

A galvanic cell which contains a chromium electrode and electrolyte for one half-cell runs for 6.0 minutes and produces a current of 2.80 A. The mass of the chromium electrode is seen to increase by 0.136 g. The charge on the chromium ion is:

$Q = It = 2.8A \times 6 \times 60 = 1008\ C$ $\quad n(Cr) = \frac{m}{M} = \frac{0.136}{52} = 0.00262 mol$

$n(e^-) = \frac{Q}{F} = \frac{1008}{96500} = 0.0104 mol$

n(metal) : n(e⁻)

0.00262 : 0.0104

A. +1

B. +2

C. +3

D. +4

1 : 4

Question 32 (4 marks)

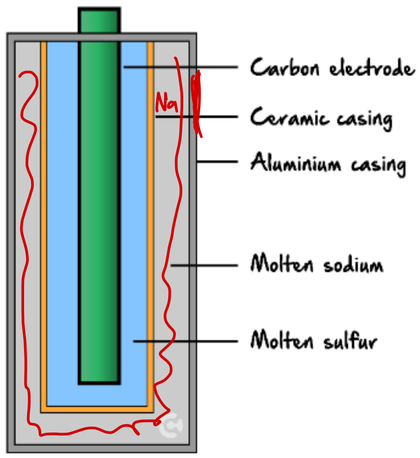
The sodium-sulfur cell is a galvanic cell with the following overall reaction during discharge:

oxidising

0 +1

electrolyte →
$$2Na(l) + S(l) \rightarrow Na_2S(l)$$

The design of one version of the cell is shown in the simplified diagram below.



The solid ceramic is made of a special material that allows only sodium ions to pass through it.

a.

i. Which component of the cell acts as the anode? (1 mark) **[1.10.1]**

aluminium casing

ii. Write the half-equation for the reduction reaction. (1 mark) **[1.10.1]**

$S_{(l)} + 2e^- \rightarrow S^{2-}_{(l)}$

b. In a normal galvanic cell, the half-cells are connected by a salt bridge.

i. What is the purpose of the salt bridge? (1 mark)

to complete circuit

ii. Which component of the sodium-sulfur cell acts as the salt bridge? (1 mark) **[1.10.1]**

ceramic casing

Check off any learning objectives obtained full marks from the "Contour Check" booklet!

Space for Personal Notes

Sub-Section: Additional

Question 33 (7 marks)

Aashan is a bit of a sceptic and wishes to verify Faraday's constant. To do so, he constructs an electrochemical cell between Ag^{+}/Ag and Sn^{4+}/Sn^{2+}.

a. Aashan calculates, using the current read by the ammeter and the time over which he ran the cell, that the cell produces 750 J of energy. Calculate the charge, in C, passing through the cell. (2 marks) **[1.10.2]**

V=0.80 –(0.15)=0.65 V
Q=E/V=750/0.65=1154 C

b. Aashan measures 1.294 g of silver being deposited onto the cathode. Using this information, what amount of electrons must have arrived at the cathode, in mol? (2 marks) **[1.10.2]**

n(Ag)=m/M=1.294/107.9=0.012 mol
n(e^-)=n(Ag)=0.012 mol

c.

i. Hence, estimate Faraday's constant to five significant figures. (1 mark) **[1.10.2]**

F=Q/n(e^-)=1154/0.012=96213 C/mol

ii. Thus, use your answer to **part c. i.** and Avogadro's number, N_A, to calculate the elementary charge, e, on an electron to 3 decimal places. (1 mark) **[1.10.2]**

e^-=F/N_A=96213/(6.02×10^23)= -1.598 x 10^(-19) C

d. Propose why there is a discrepancy between your values for the constants *F* and *e* and those found in the literature. (1 mark) **[1.10.2]**

The mass of silver measured was incorrect; the conditions were not SLC or 1.0 M and thus voltage was incorrect; the ammeter read an incorrect value; etc.

Space for Personal Notes

CONTOUREDUCATION

Website: contoureducation.com.au | Phone: 1800 888 300 | Email: hello@contoureducation.com.au

VCE Chemistry ¾

Free 1-on-1 Consults

<u>What are 1-on-1 Consults?</u>

- **Who Runs Them?** Experienced Contour tutors (45 + raw scores and 99 + ATARs).
- **Who Can Join?** Fully enrolled Contour students.
- **When Are They?** 30-minute 1-on-1 help sessions, after-school weekdays, and all-day weekends.
- **What To Do?** Join on time, ask questions, re-learn concepts, or extend yourself!
- **Price?** Completely free!
- **One Active Booking Per Subject:** Must attend your current consultation before scheduling the next :)

<u>SAVE THE LINK, AND MAKE THE MOST OF THIS (FREE) SERVICE!</u>

Booking Link

bit.ly/contour-chemistry-consult-2025