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VCE Chemistry ¾
Galvanic Cells [0.7]
Workshop Solutions

Error Logbook:

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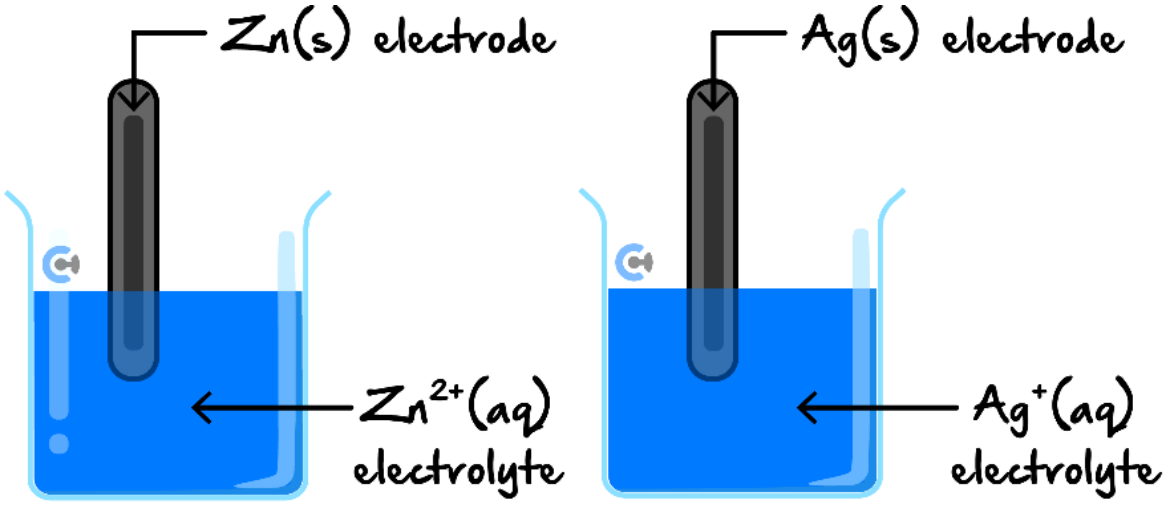
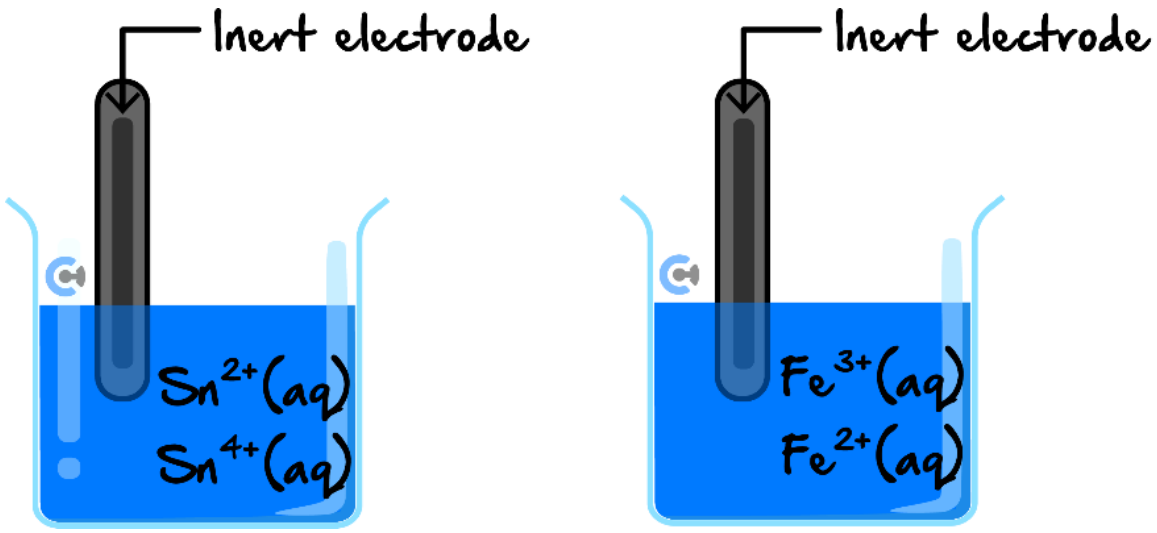
Section A: Recap (8 Marks)

Learning Objective: [1.8.1] – Identify Electrodes and Salt Bridge/Electron Movement During Galvanic

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

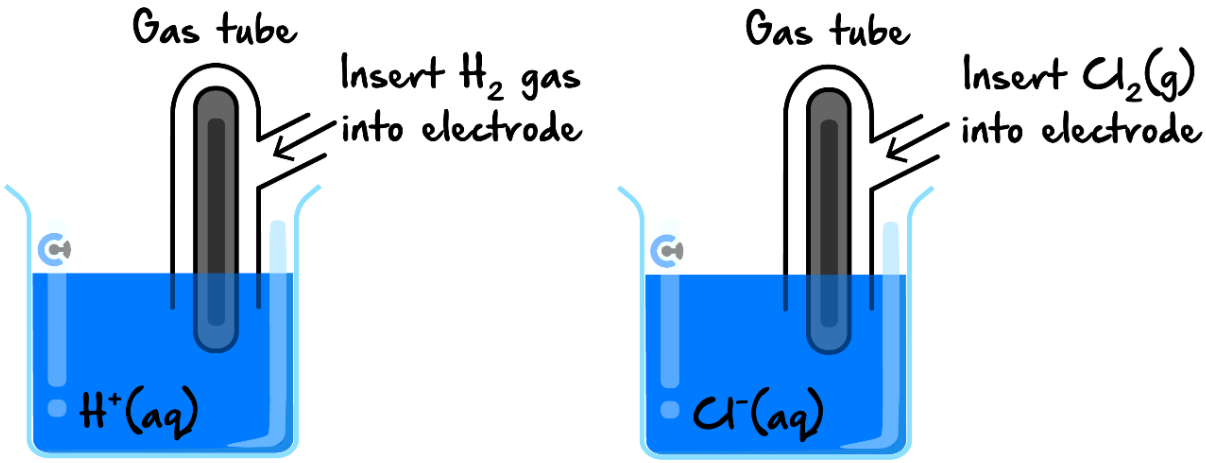


- Ion / Ion half-cell.
  - Example: $Sn^{4+}(aq)/Sn^{2+}(aq), Fe^{3+}(aq)/Fe^{2+}(aq)$



  - Material of electrode: _____ inert (Pt/C) _____

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

- **Internal Circuit (Salt Bridge)**
  - Purpose: To balance the ___ buildup of charge ___, to ___ complete the circuit ___.

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

  - **Properties of salt bridge:** ___ Soluble & inert ___
  - **The most common salt bridge used:** ___ $KNO_3(aq)$ ___

Learning Objective: [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 which 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
Chemical → Thermal	Chemical → Electrical

- Electromotive Force (EMF) formula:

$$\mathbf{EMF} = E^0(oxidant) - E^0(reductant)$$

- **Conditions:** ___ SLC ___ & ___ $1.0\ M$ ___

- 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.

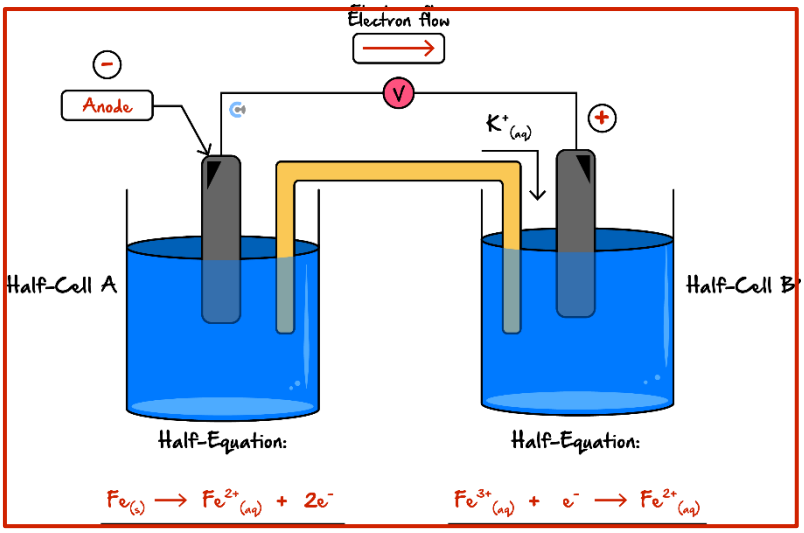
Learning Objective: [1.8.3] – Identify & Explain Observations During the Operation of Galvanic Cells

- **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 $\mathbf{H^+}$ or $\mathbf{OH^-}$ ions in half-equations).
  - _____ Bubbles _____ being formed (look out for **gases** being formed).
  - **Change in** _____ colour _____ (look out for **ions** being produced/consumed in solution).
- **Procedure**: Write out _ half-equations _ first with RC + and AO –, then figure out observations.

Space for Personal Notes

Question 1 (8 marks) **Walkthrough.**

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



a. In the spaces provided above, write the half-equations which occur at each half-cell. (2 marks)

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)

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

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

Yellow/brown ⟶ pale green

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

Colour of electrolyte becomes more intensely green.

Electrode decreases in size or mass.

f. Find the EMF of this cell. (1 mark)

$0.77 - (-0.44) = 1.21\ V$

Section B: Warm Up (12 Marks)

INSTRUCTION: 12 Marks. 8 Minutes Writing.

Question 2 (6 marks)

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

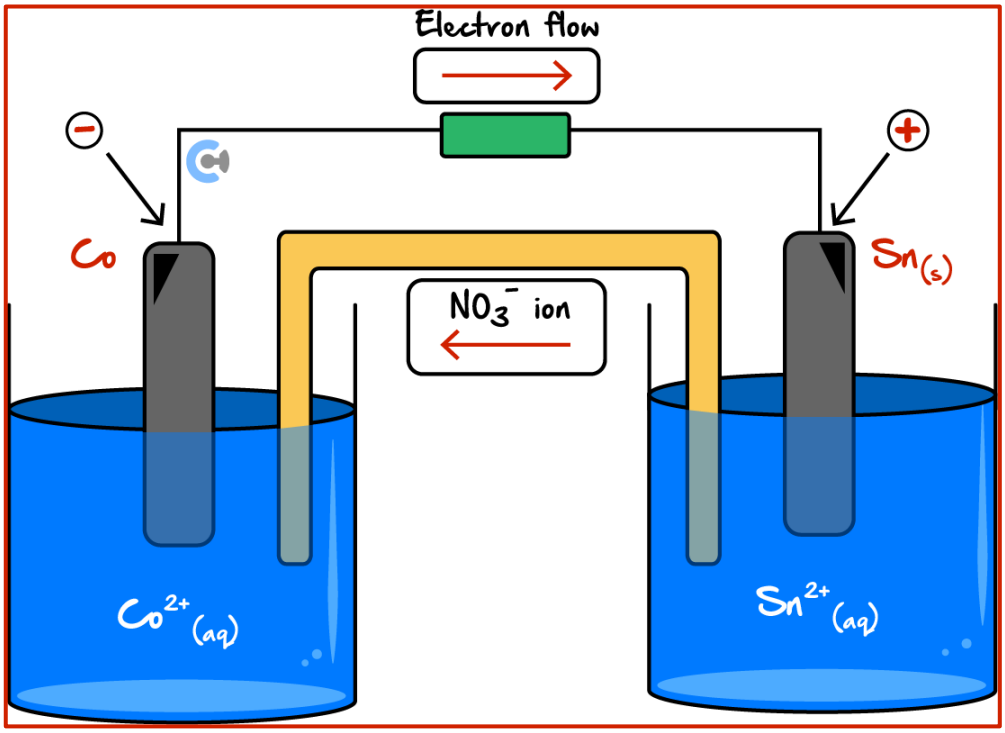
a. Write the half-equations which occur at the:

i. Negative electrode. (1 mark)

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

ii. Positive electrode. (1 mark)

$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$



b. Label the direction of movement of electrons and nitrate ions in the relevant boxes provided above. (1 mark)

c. Label the polarities of the electrodes in the circles provided above. (1 mark)

d. Given that tin (II) ions dissolved in water are yellow and cobalt (II) dissolved in water are pink, list four observations that will be observed as the reaction takes place. (2 marks)

- Increase in mass/size of Sn electrode.
- Decrease in mass/size of cobalt electrode.
- Decrease in intensity of yellow colour in reduction half-cell.
- Increase in intensity of pink colour in oxidation half-cell.

Question 3 (6 marks)

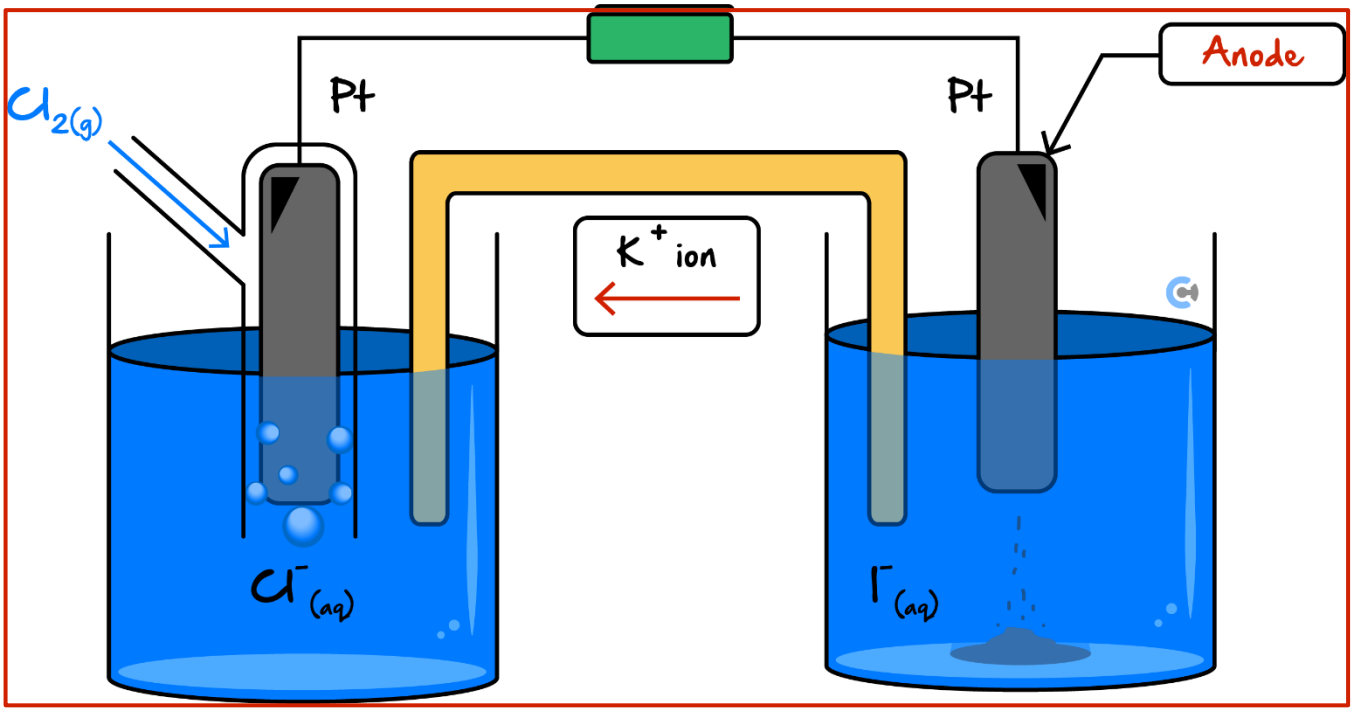
A galvanic cell contains the half-cells $Cl_2(g)/Cl^-(aq)$ and $I_2(s)/I^-(aq)$.

a. Write the half-equations which occur at each electrode. (2 marks)

Cathode: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

Anode: $2I^-(aq) \rightarrow I_2(s) + 2e^-$

b. The galvanic cell is shown below.



i. Label the direction of movement of potassium ions in the box provided above, and label the right electrode as the cathode or the anode. (2 marks)

ii. Explain the purpose of the potassium ions. (1 mark)

To maintain electric neutrality / to complete the circuit / to balance the build-up of charge.

iii. A solid sludge is seen to form at the bottom of the right electrode. Provide a reason for this observation. (1 mark)

$I_2(s)$ is formed, and does not stick to the electrode.

Space for Personal Notes

Section C: Ramping Up (11 Marks)

INSTRUCTION: 11 Marks. 8 Minutes Writing.

Question 4 (1 mark)

A galvanic cell has the following setup:

- Half-cell 1: Acidified 1 *M* Hydrogen Peroxide H_2O_2 solution with a Pt electrode.
- Half-cell 2: 1 *M* Nickel Nitrate, $Ni(NO_3)_2$ solution with a Ni electrode.

When the cell is operating, the oxidising agent is:

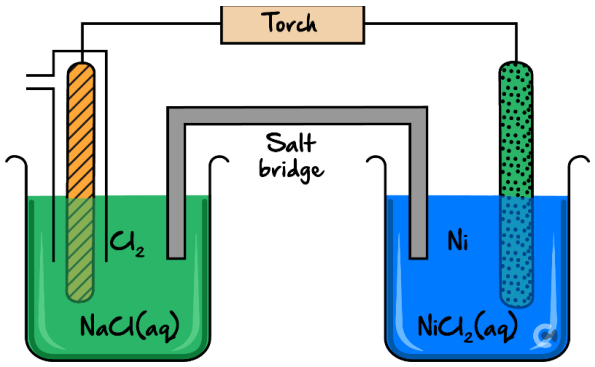
A. Ni

B. Ni^{2+}

C. H_2O

D. H_2O_2

Question 5 (1 mark)



For this cell, the:

A. Electrons will flow from the chlorine to the nickel half-cells.

B. Nickel electrode will be the negative electrode.

C. Concentration of nickel (II) ions in the solution will be decreasing.

D. Chlorine electrode will be the negative electrode.

The following information applies to the two questions that follow.

A galvanic cell is constructed from the following half cells, at 25°C.

	Electrode	**Half Cell Solution (all concentrations 1.0 *M*)**
Half cell 1	Silver	Colourless solution of $AgNO_3$
Half cell 2	Copper	Blue-coloured solution of $CuCl_2$

The half cells are connected with a salt bridge, and the electrodes are joined by a wire.

Question 6 (1 mark)

Which one of the following is likely to occur?

A. The copper electrode will increase in mass.

B. Bubbles of gas will form at the copper electrode.

C. The concentration of Silver ions in solution will increase.

D. The blue colour of the copper (II) chloride solution will become more intense.

Question 7 (1 mark)

When the current is flowing:

A. The anode is positive, and the cathode is negative.

B. An oxidation reaction occurs at the positive electrode.

C. Anions in the salt bridge move towards the negative electrode.

D. Electrons travel in the external circuit from the cathode to the anode.

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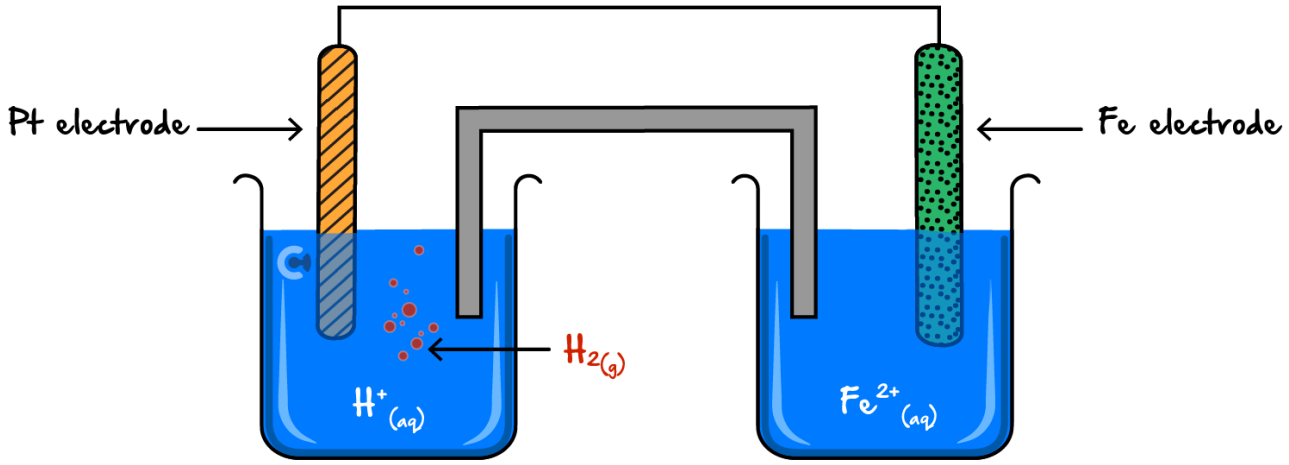
Question 8 (1 mark)

A diagram of a galvanic cell is shown below.



For the reaction in this galvanic cell:

A. $H^+(aq)$ is the oxidising agent, and the concentration of H^+ ions will increase.

B. $H^+(aq)$ is the oxidising agent, and the concentration of H^+ ions will decrease.

C. $H_2(g)$ is the reducing agent, and the concentration of H^+ ions will increase.

D. $H_2(g)$ is the reducing agent, and the concentration of H^+ ions will decrease.

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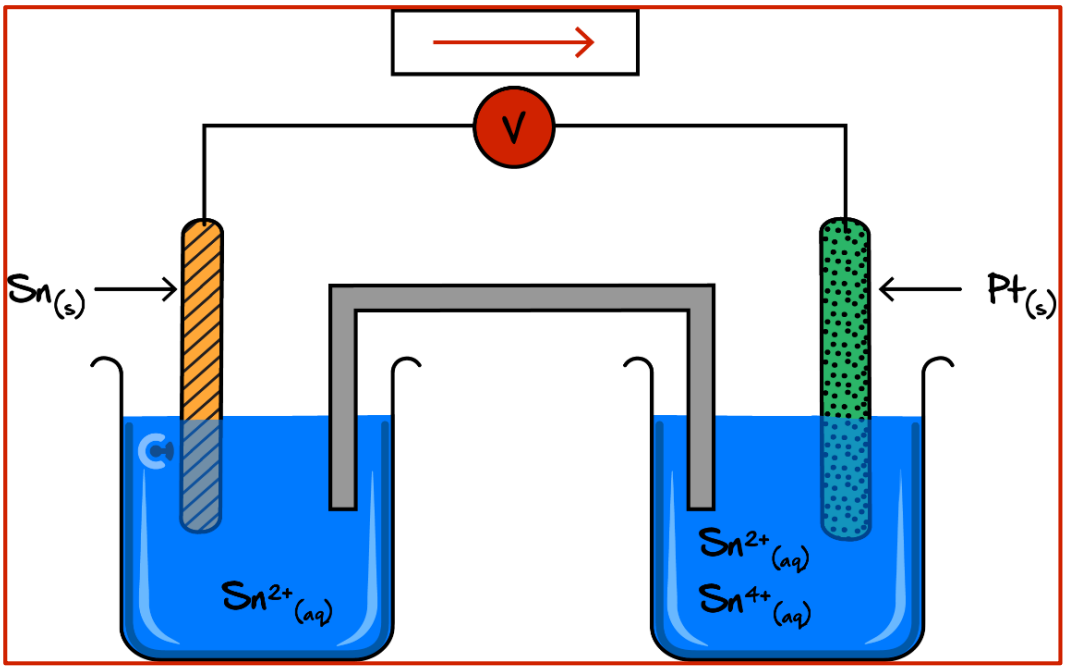
Question 9 (6 marks)

A chemistry teacher wanted to explore Tin and its ions by setting up the following galvanic cell:



a. On the diagram above, label the direction of electron flow. (1 mark)

b. Write the equation for the overall reaction which takes place. (1 mark)

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

c. Explain whether this is an example of a direct or an indirect redox reaction. Outline the energy transformations that would occur in this cell. (2 marks)

Indirect as reactants are not in direct contact with one another
Chemical E → Electrical + Thermal Energy (need both).

d. If this cell were constantly supplied with the required reactants, would it be able to produce electricity indefinitely? Explain your answer. (2 marks)

No, as the ions in the salt bridge would eventually run out, rendering the cell inactive.

Section D: Getting Trickier I (10 Marks)

INSTRUCTION: 10 Marks. 8 Minutes Writing.

Question 10 (3 marks)

In electrochemistry, a 'gas electrode' contains a conducting material over which a gas is bubbled or collected. A galvanic cell, constructed at standard conditions, contains two different gas electrodes. One of the relevant half-equations is shown below.

$$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(g) \qquad E^0 = +0.40\ V$$

The theoretical maximum cell voltage is $0.96\ V$.

a. Write the half-equation for the half-reactions which occur at each electrode. (2 marks)

Positive electrode: $Cl_2(g) + 2e^-(aq) \rightarrow 2Cl^-(aq)$

Negative electrode: $H_2S(g) \rightarrow S(s) + 2H^+(aq) + 2e^-$

b. Which of the following observations, made when the cell was operating, is most likely to be accurate? (1 mark)

A. A solid is produced at the anode.

B. Gas is produced at the cathode.

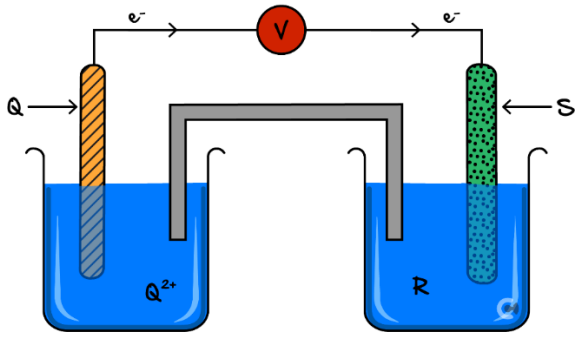
C. pH increases at the anode.

D. pH increases at the cathode.

Space for Personal Notes

Question 11 (7 marks)

A diagram of an electrochemical cell is shown below.



a. Write the equation occurring in the right half-cell, given R represents $H^+(aq)$. (1 mark)

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

b. List 3 observations that can be made based on the information provided thus far. (3 marks)

1. Bubbles ($H_2(g)$) at cathode.
2. $c(H^+)$ decreases → less acidic → pH increases at cathode.
3. Q electrode will get thinner/smaller/lighter
CANNOT TALK ABOUT ANODE SOLUTION BECOMING DARKER BECAUSE Q^{2+} COULD BE COLOURLESS.

c. To determine the identity of Q, the Q^{2+}/Q half-cell is now connected with a Ni^{2+}/Ni half-cell and it is observed that the cations in the salt bridge migrate to the Q^{2+} electrolyte.

Using this information, state the possible range of EMF values the Q^{2+}/Q half-cell may have. Explain your reasoning. (2 marks)

Q must be cathode and above Ni due to cations migrating to it.
Therefore $E^0 > -0.25\ V$
Since it was anode with H^+, $E^0 < 0.00\ V$
Therefore, $-0.25\ V < E^0 < 0.00\ V$

d. When tested at SLC, however, it turns out that the voltage recorded was not in the expected range. Suggest a possible reason why. (1 mark)

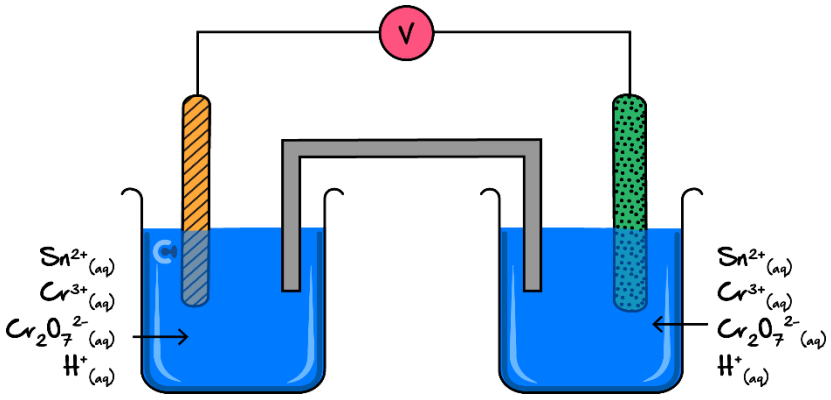
Concentration of Q^{2+} is not at 1.0 M.

Section E: Getting Trickier II (14 Marks)

INSTRUCTION: 14 Marks. 13 Minutes Writing.

Question 12 (1 mark)

A diagram of a galvanic cell at standard conditions is shown below.



Some half-reactions relevant to this galvanic cell are given in the following table.

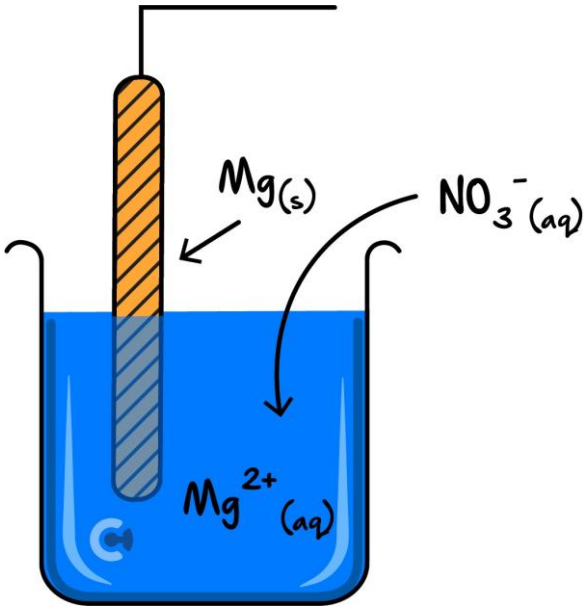
Half-Reaction	Standard Electrode Potential at 25°C
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	$+1.36$
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	-0.14
$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	-0.42
$Cr^{3+}(aq) + e^- \rightleftharpoons Cr(aq)$	-0.74

If the galvanic cell produces 1.5 V, the anode and cathode should be made from which of the following?

	Anode	Cathode
A.	Pt(s)	Sn(s)
B.	Sn(s)	Sn(s)
C.	Pt(s)	Pt(s)
D.	Sn(s)	Pt(s)

Question 13 (4 marks)

A particular metal/ion half-cell is shown below:



a. Write the half-equation occurring **and** state the polarity of this half-cell based on the information provided in the diagram. (2 marks)

Half-Equation	Polarity
$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$	Negative

b. Given the other half-cell involves chlorite ions, $ClO_2^-(aq)$, and ClO_2 gas, write the balanced half-equation that will be occurring. (1 mark)

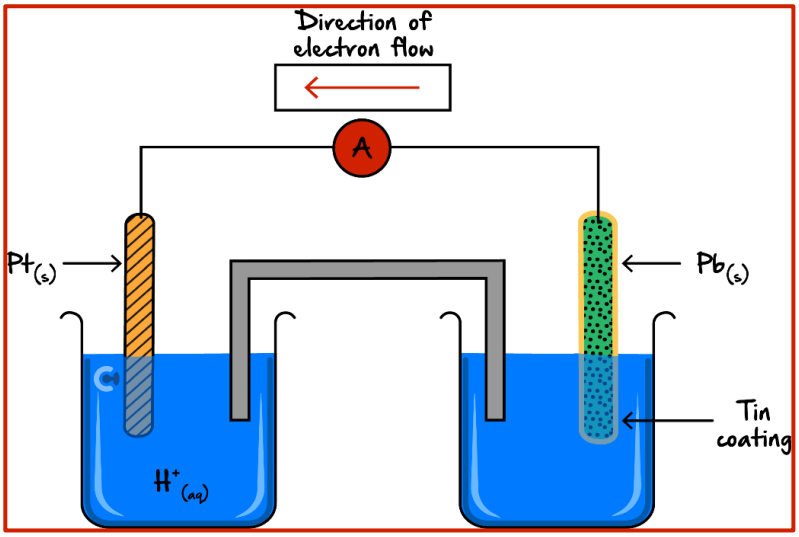
$ClO_2(g) + e^- \rightarrow ClO_2^-(aq)$

c. Assuming all solutions are 1 M concentration if this galvanic cell generates 3.32 V, calculate the standard electrode potential of the chlroite ion half-cell. (1 mark)

$$\text{EMF} = E^0(oxidant) - E^0(reductant)$$
$$E^0(oxidant) = \text{EMF} + E^0(reductant) = 3.32 - 2.37 = 0.95\ V$$

Question 14 (9 marks)

A galvanic cell is set below, whereby excess $H^+(aq)$ ions are set up in one half-cell with a platinum electrode, and a lead electrode which has been fully covered with tin on the outside has been used as the electrode on the other side, with no ions in the electrolyte.



a. Label the direction of electron flow on the box provided above. (1 mark)

b. In the first 5 minutes of the cell running, the same half-equations occur at each electrode. Write the balanced half-equation for the reaction which takes place at the:

i. Positive electrode. (1 mark)

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

ii. Negative electrode. (1 mark)

$Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$

c. After 5 minutes, the reaction at one of the electrodes seems to change. Write the new half-equation which takes place at this electrode, giving justification for your reasoning. (3 marks)

The anode half-equation changes to:
$Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$ instead

This is because the $Sn(s)$ runs out, exposing the $Pb(s)$on the inside, and resulting in this $Pb(s)$ oxidising instead.

d. The electromotive force generated by the cell also seems to change over time. Explain how the EMF produced changes over time, providing numerical values of the EMF each time it changes. (3 marks)

At the beginning as H^+ reduces and Sn oxidises, the EMF produced is $0.14\ V$. As time progresses, Sn runs out, and Pb oxidises instead, decreasing the EMF produced to $0.13\ V$. The EMF then remains at $0.13\ V$ until the cell stops running.

Let's take a BREAK!

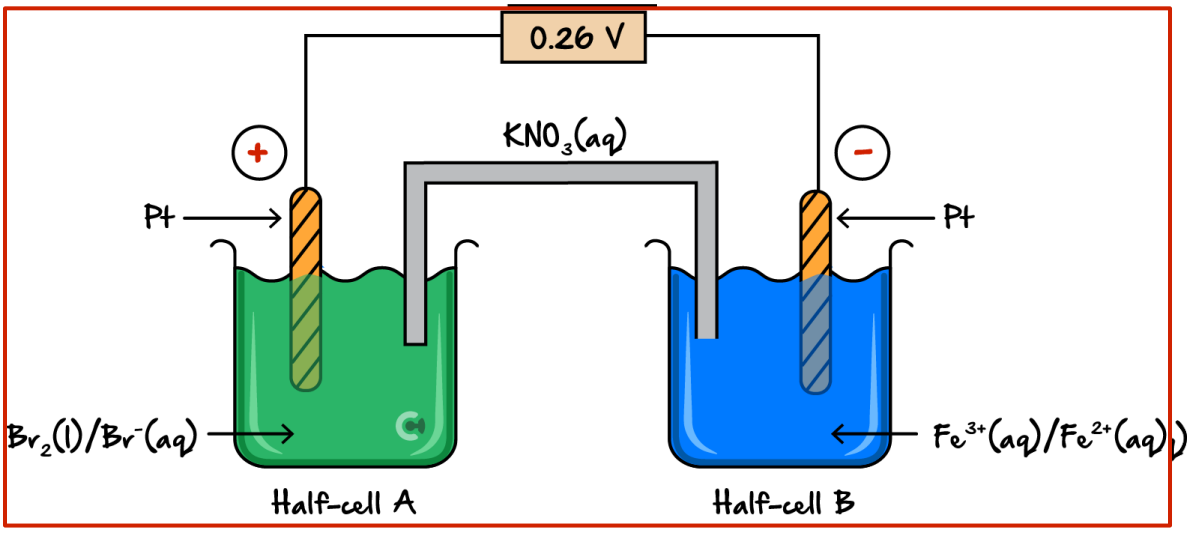
Space for Personal Notes

Section F: VCAA-Level Questions I (13 Marks)

INSTRUCTION: 13 Marks. 30 Seconds Reading. 12 Minutes Writing.

Question 15 (8 marks)

A student, as part of an experimental investigation of galvanic cells, constructed a cell using $Fe^{3+}(aq)/Fe^{2+}(aq)$ and $Br_2(l)/Br^-(aq)$ half-cells as shown in the figure below.



a. Label the charge of the cathode and anode on the diagram above. (1 mark)

b. Write the balanced equation for the half-reactions that would occur in:

i. Cathode. (1 mark)

$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$

ii. Anode. (1 mark)

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

c. Justify which way the cations and anions would flow through the salt bridge. (1 mark)

Since half-cell A contains the cathode then the cations, $K^+(aq)$, will flow through the salt bridge into this half-cell, and the anions, $NO_3^-(aq)$, will flow into half-cell B. (1 mark)

Bromide ions, Br^- (aq), are being formed in half-cell A, therefore to maintain charge neutrality in this solution, cations must flow into this half-cell. In half-cell B, iron(II) ions, $Fe^{2+}(aq)$, are being oxidised to iron(III)ions, $Fe^{3+}(aq)$, so anions must flow into this half-cell to maintain charge neutrality.

d. Determine the maximum voltage that this cell would deliver. (1 mark)

The maximum voltage that the cell can produce is the difference between the two standard electrode potentials.
$E = 1.09 - 0.77 = 0.32\ V$

e. The student observed that the cell they constructed delivered $0.26\ V$.

Provide an explanation as to why this was different to the voltage determined in **part d**. (1 mark)

The maximum voltage is calculated from the standard electrode potentials at SLC. In this cell it would require that the concentrations of the four species involved have concentrations of $1.0\ M$.

f. Explain the effect that replacing the Platinum electrode in half-cell B with an iron electrode would have on:

i. The maximum voltage that the cell could deliver. (1 mark)

Referring to - The Electrochemical Series, Table 2: VCE Chemistry Data Book, it can be determined that changing the platinum electrode in half-cell B to an iron electrode will now have iron as the stronger reductant.
$Fe^{2+}(aq) + 2e^- \rightleftarrows Fe(s) \qquad E^0 = -0.44\ V$
The maximum voltage that could be produced would be
$E = 1.09 - (-0.44) = 1.53\ V$. (1 mark)

ii. The reaction that would occur in the cell. Write any new half-equations which may occur. (1 mark)

The reaction in half-cell A will remain unchanged, while in half-cell B the iron will be oxidised to iron(II) ions. (1 mark).
Half-cell B: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
Overall: $Br_2(l) + Fe(s) \rightarrow 2Br^-(aq) + Fe^{2+}(aq)$
Some oxidation of the iron(II) ions to iron(III) ions will still occur, however the oxidation of the iron will be the dominant reaction as this is the stronger reductant.

Space for Personal Notes

Question 16 (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 which 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)

$$2NO_3^-(aq) + 8H^+(aq) + 3Cr(s) \rightarrow 2NO(g) + 4H_2O(l) + 3Cr^{2+}(aq)$$

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)

Connect each half-cell in turn to the standard hydrogen half-cell, noting the voltage and polarity of each. **(1 mark)**

Those half-cells which are positive in the cell are given positive voltages and those which are negative are given negative voltages. Half-cell reactions can then be placed in order of decreasing voltage to generate an electrochemical series. **(1 mark)**

Space for Personal Notes

Section G: Multiple Choice Questions (7 Marks)

INSTRUCTION: 7 Marks. 7 Minutes Writing.

Question 17 (1 mark)

A galvanic cell consists of two connected half-cells that can produce an electron flow.

Which combination of standard half-cell pairs would be expected to result in a cell potential of 1.41 V?

A.	Al electrode with $Al(NO_3)_3$	Ag electrode with $AgNO_3$
B.	Zn electrode with $Zn(NO_3)_2$	Ni electrode with $Ni(NO_3)_2$
C.	Ni electrode with $Ni(NO_3)_2$	Al electrode with $Al(NO_3)_3$
D.	Ag electrode with $AgNO_3$	Zn electrode with $Zn(NO_3)_2$

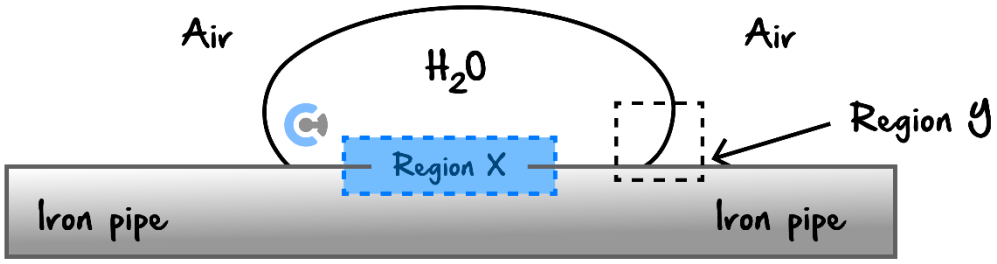
Space for Personal Notes

Question 18 (1 mark)

Inspired from VCAA Chemistry Exam 2009
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2009chem2-w.pdf#page=6

Iron pipes are used to transport natural gas to cities. Corrosion occurs when water droplets sit on the outer surface of the iron pipe.

Miniature galvanic cells are created, with regions such as those shown below, that act as anodes and cathodes.



The type of region and reaction occurring at *X* in the cell is:

	Region	Reaction
A.	Anode	$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
B.	Cathode	$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
C.	Anode	$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
D.	Cathode	$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$

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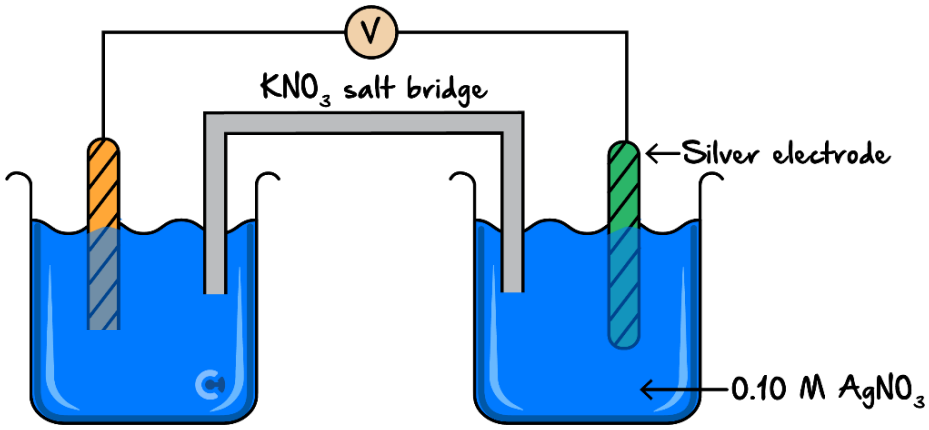
Inspired from VCAA Chemistry Exam 2014
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2014/2014chem-amd-w.pdf#page=13

The following information applies to the two questions that follow.

The following setup is used, with a $Sn^{4+}(aq)/Sn^{2+}(aq)$ half-cell used for the left half-cell.



Question 19 (1 mark)

Which one of the following statements about the cell above is true as the cell discharges?

A. The silver electrode is the negative electrode.

B. The concentration of Ag^+ ions will increase.

C. The nitrate ions will go into the $Sn^{4+}(aq)/Sn^{2+}(aq)$ half-cell to form $Sn(NO_3)_4$.

D. The maximum theorteical EMF generated is $+0.65\ V$.

Question 20 (1 mark)

What should be observed at the $Sn^{4+}(aq)/Sn^{2+}(aq)$ half-cell as the cell discharges?

A. Bubbles will form over the surface of the electrode.

B. The electrode will become thinner and pitted.

C. The half-cell becomes more positively charged over time.

D. The combined concentration of $Sn^{4+}(aq)$ and $Sn^{2+}(aq)$ stays the same.

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Question 21 (1 mark)

Consider the following half cells which are set up under standard conditions.

Half cell	Electrode	Electrolyte
I	Metal A	$A^{2+}(aq)$
II	Platinum	$B^{2+}(aq)$ and $B^{3+}(aq)$
III	Metal C	$C^{+}(aq)$

- When a galvanic cell is constructed from half cell I and half cell II, the electrode in half cell II is negative.
- When a galvanic cell is constructed from half cell II and half cell III, the electrode in half cell III is negative.

The strongest oxidant is:

A. $A^{2+}(aq)$

B. $B^{2+}(aq)$

C. $B^{3+}(aq)$

D. $C^{+}(aq)$

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Even if a student was unsure about the nature of the reaction of Ca with water, they should have known that A was correct because:

- (B) – the Pt|Sn^{4+}, Sn^{2+} half cell does not require Sn
- (C) – there are many soluble compounds of calcium
- (D) – the half cell containing the Sn^{4+} is separated from the proposed half cell containing Ca with a salt bridge.

Question 22 (1 mark)

A student is planning to set up a demonstration of a galvanic cell using half cells constructed as follows:

- Half cell 1: A calcium electrode in a beaker containing an aqueous solution of Ca^{2+} ions.
- Half cell 2: A platinum electrode in a beaker containing an aqueous solution of a mixture of Sn^{4+} and Sn^{2+} ions.

A salt bridge would connect the two beakers. The electrodes would be attached to a voltmeter.

This particular cell is impractical because:

A. Solid calcium (Ca) will react directly to reduce water to hydrogen gas.

B. There is no solid tin (Sn) in the half cell containing $Sn^{4+}(aq)$ and $Sn^{2+}(aq)$.

C. There are no known ionic compounds of calcium that are soluble in water.

D. $Sn^{4+}(aq)$ will be in contact with Ca and will oxidise it to $Ca^{2+}(aq)$.

Question 23 (1 mark)

It is predicted from the electrochemical series that the voltage of the galvanic reaction shown below is $1.76\,V$.

$$2VO_2{}^{+}(aq) + 4H^{+}(aq) + Zn(s) \rightarrow 2VO^{2+}(aq) + 2H_2O(l) + Zn^{2+}(aq)$$

When the galvanic cell was constructed, the actual voltage was $1.89\,V$.

What is the most likely explanation for this discrepancy?

A. The rate of the reaction cannot be predicted from the electrochemical series.

B. The conditions used to set up the galvanic cell were not standard conditions.

C. Voltage predictions from the electrochemical series have a margin of error of 5-10%.

D. A highly purified sample of Zinc metal was used in the galvanic cell.

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Section H: VCAA-Level Questions II (11 Marks)

INSTRUCTION: 11 Marks. 30 Seconds Reading. 10 Minutes Writing.

Question 24 (6 marks)

A galvanic cell was set up with Cl_2/Cl^- and Fe^{3+}/Fe^{2+} as the two half-cells.

a. Write the half-equation occurring at the anode. (1 mark)

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

b. Propose a suitable material for the cathode. Justify your answer. (1 mark)

Pt(s) or Graphite. Needs to be a conductor but inert.

c. With reference to this galvanic cell, explain why cations in the salt bridge do not move to the negative electrode. (2 marks)

At the anode, the electrolyte becomes more positive (Fe^{2+} is being replaced by Fe^{3+}), so negative anions are needed to balance this build up of charge.
Ion flow has nothing to do with electrode polarity.

d. If, for every 5 mol of $Cl_2(g)$ being pumped in, 112 L of volume was required.

i. Find the molar volume of the chlorine gas. (1 mark)

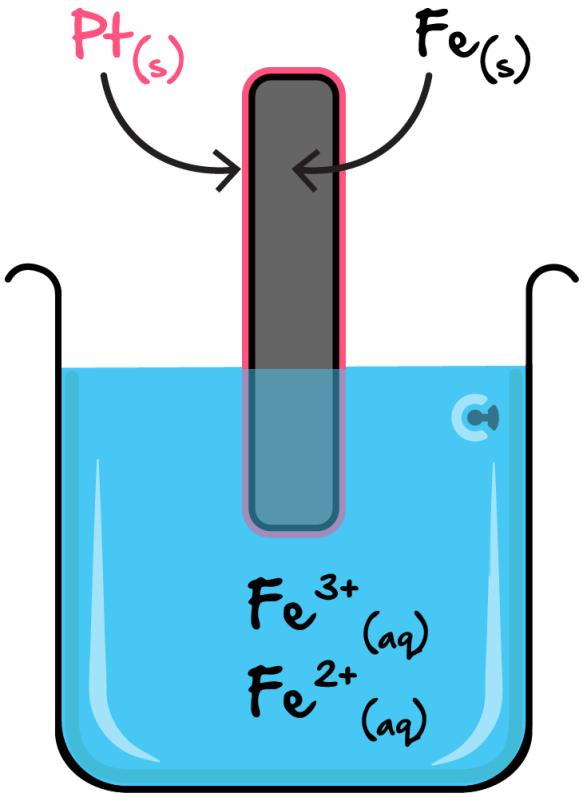
$$V_m = \frac{V}{n} = \frac{112}{5} = 22.4\ L/mol$$

ii. Explain whether the electromotive force of the cell can be accurately predicted using the electrochemical series in **Item (1)** of the Data Book. (1 mark)

As the molar volume is not 24.8, not SLC → EMF cannot be calculated accurately.

Question 25 (5 marks)

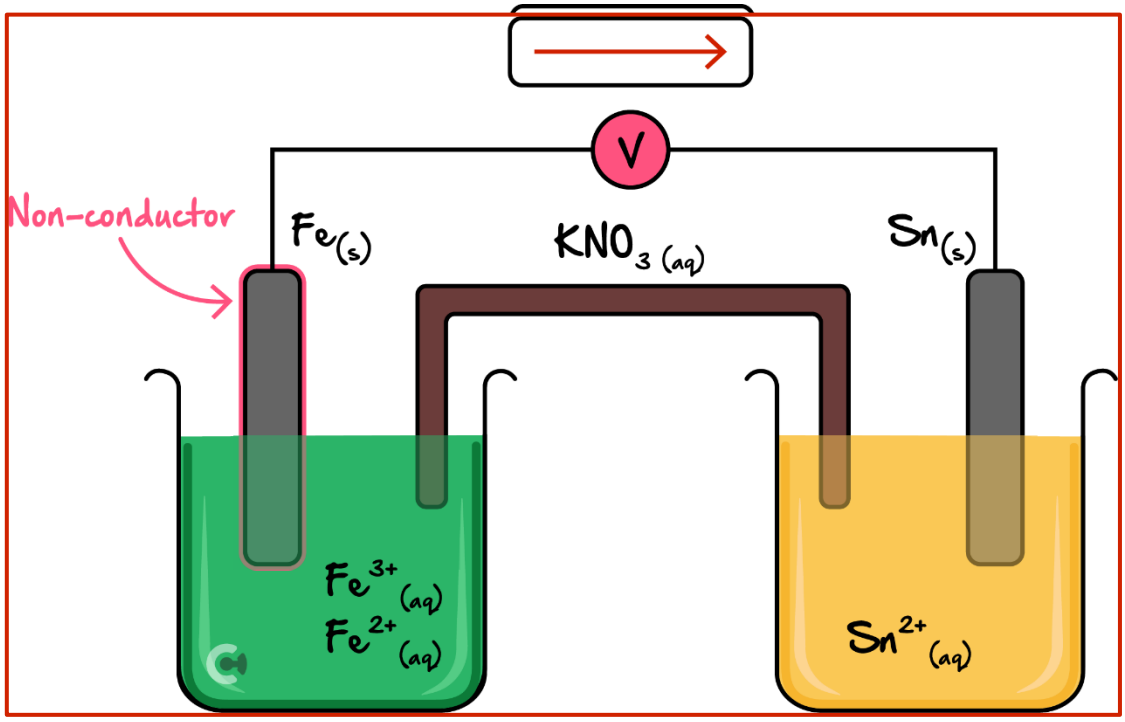
An iron electrode is completely covered in a non-conductor of electricity on the outside and is dipped into an electrolyte containing iron (II) and iron (III) ions as shown below.



a. Explain whether a spontaneous reaction will occur. (1 mark)

No, as there is no direct contact with $Fe^{3+}(aq)$ ions and Fe(s) ions.

b. The half-cell constructed above is then connected through an external and internal circuit to a $Sn^{2+}(aq)/Sn(s)$ half-cell as shown below.



i. Label the direction of electron flow in the box provided above. (1 mark)

ii. Write the reduction half-equation. (1 mark)

$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$

iii. Write the oxidation half-equation. (1 mark)

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

iv. Write the energy conversions which take place. (1 mark)

Chemical → electrical energy

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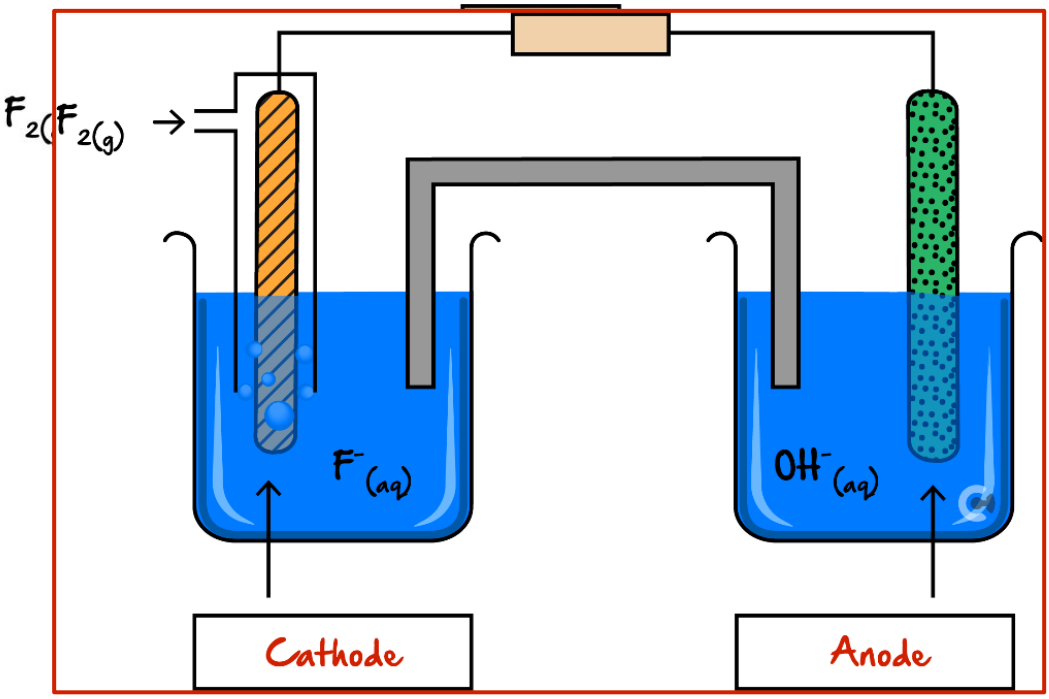
Section I: Extension Questions (5 Marks)

Question 26 (5 marks)

The following galvanic cell is set up.



a. Label the cathode and anode in the boxes provided above. (1 mark)

b. Write the half-equations which occur at the:

i. Cathode. (1 mark)

$F_2(g) + 2e^- \rightarrow 2F^-(aq)$

ii. Anode. (1 mark)

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

c. As the cell runs, bubbles are observed to be produced at an electrode. Determine whether this electrode is the positive or negative electrode, identifying the gas which causes the bubbles. (1 mark)

Negative electrode – Oxygen gas is produced.

d. A pH metre is placed into both half-cells, and a change in pH is measured in only one of the half-cells. Determine whether this electrode is a positive or negative electrode, explaining what happens to the pH at this electrode. (1 mark)

Negative electrode – OH^- ions are consumed, making it less basic. This decreases the pH at this electrode.

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