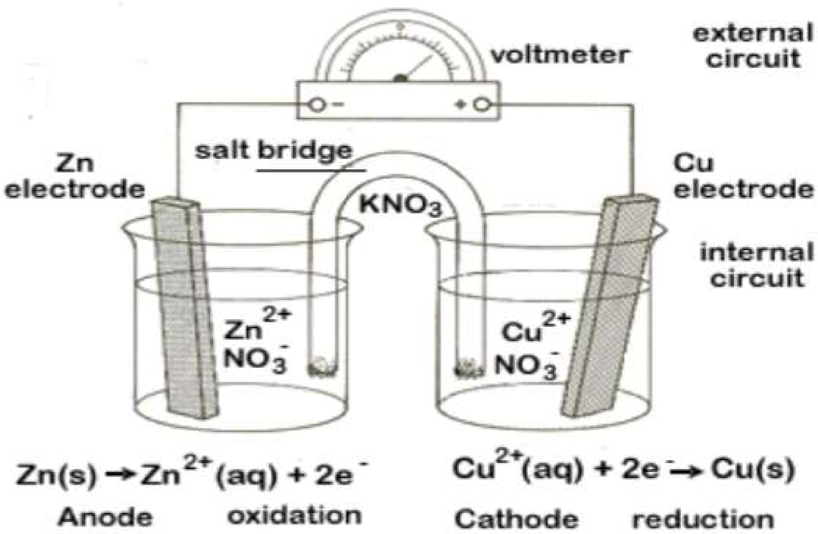
PHYSICAL SCIENCES

GRADE 12

ELECTROCHEMICAL REACTIONS



MOVING BEYOND THE LIMITS, SETTING THE STANDARD AND LEAVING NO CHILD BEHIND

ELECTROCHEMICAL REACTIONS

FORMULAE

$E^{\theta}_{cell} = E^{\theta}_{cathode} - E^{\theta}_{anode} \,/\, E^{\theta}_{sel} = E^{\theta}_{katode} - E^{\theta}_{anode}$

or

$E^{\theta}_{cell} = E^{\theta}_{reduction} - E^{\theta}_{oxidation} \,/\, E^{\theta}_{sel} = E^{\theta}_{reduksie} - E^{\theta}_{oksidasie}$

or

$E^{\theta}_{cell} = E^{\theta}_{oxidising agent} - E^{\theta}_{reducing agent} \,/\, E^{\theta}_{sel} = E^{\theta}_{oksideermiddel} - E^{\theta}_{reduseermiddel}$

TERMINOLOGY

Electrochemical reactions involve the transfer of electrons. In this type of chemical reaction, one chemical substance loses electrons while another receives them. The processes of losing and receiving electrons happen at the same time. All ionic compounds are made in redox reactions. The word "redox" refers to reduction and oxidation.

Oxidation is the loss of electrons.(OIL)

Oxidation: An increase in oxidation number

$Fe \longrightarrow Fe^{2+} + 2e^{-}$(out)

Iron lose two electrons and is oxidised to Fe^{2+}

$Cu \longrightarrow Cu^{2+} + 2\,e^{-}$

Cu is being oxidised, Cu is a reducing agent

Reduction is the gain of electrons.

Reduction: A decrease in oxidation number

The oxygen atom changes to O^{2-} by gaining two electrons

$\frac{1}{2}O_2 + 2e^{-}$ (In) $\longrightarrow O^{2-}$

Oxygen gains two electrons and is reduced to O^{2-}

$Br_2 + 2\,e^{-}$ (In) $\longrightarrow 2Br^{-}$

Gain electrons = reduction(RIG)

Br_2 is being reduced, Br_2 is an oxidising agent

Reducing agent is a substance that is oxidised/loses electrons. A strong reducing agent has a conjugate oxidising agent. Zn conjugate Zn^{2+}

Oxidising agent is a substance that is reduced/gains electrons. A strong oxidising agent has a weak reducing agent. Cu^{2+} has Cu has a weak conjugate reducing agent.

$HCl + Ag \nrightarrow$

H_2 is a strong reducing agent than Ag hence it is oxidised

Redox reaction is a reaction in which electrons are transferred from the reducing agent to the oxidising agent.

The anode is the electrode where oxidation takes place

The cathode is the electrode where reduction takes place.

HOW TO TELL IF A REACTION IS SPONTANEOUS?

Steps

1. Identify the chemical species on the table(circle only the chemical spices)
2. Pick the best reactants
3. Write out the reaction from the ones you have chosen
4. Balance the electrons and add the reactions.

Will the following combination of metals and ions react?

(a) Cu and $ZnSO_4$ (b) Zn and $CUSO_4$

Zn^{2+}, SO_4^{2-} Cu^{2+}, SO_4^{2-}

Part of redox table

$H_2O + 2e^- \rightleftharpoons H_2 + H_2 + 2OH^-$	-0,83
$Zn^{2+} + 2e^- \rightleftharpoons Zn$	-0,76
$Cr^{3+} + 3e^- \rightleftharpoons Cr^+$	-0,74
$Fe^{2+} + 2e^- \rightleftharpoons$ +ve slope Fe	-0,44
$Cu^{2+} + e^- \rightleftharpoons Cu^+$	+0,16
$SO_4^{2-} + 4H^+ + 2e^- \rightleftharpoons SO_2 + 2H_2O$	+0,17
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	+0,34

POSTIVE SLOPE SPONTANEOUS REACTION (Zn and $CUSO_4$)

Start from the chosen reactants to the opposite side when writing out the half reaction

$Zn \longrightarrow Zn^{2+} + 2e^-$(oxidation) Oil

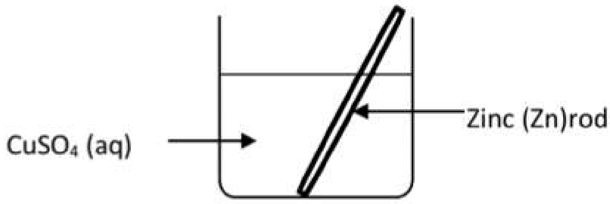
$Cu^{2+} + 2e^- \longrightarrow Cu$ (reduction) Rig

Over all equation

$Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

Direct Transfer of Electrons

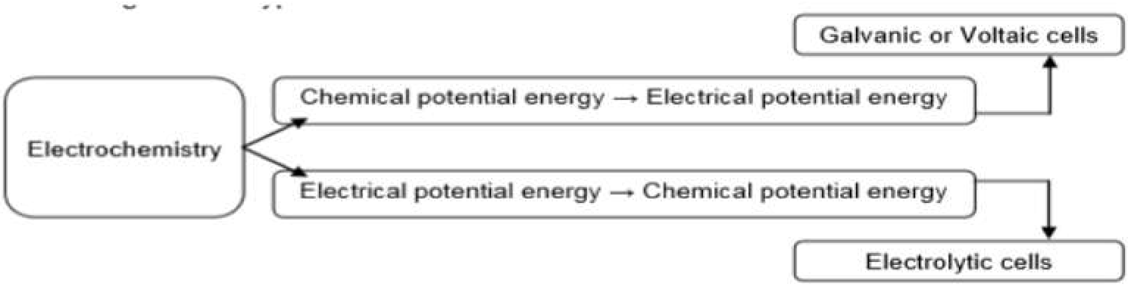
When a zinc plate is placed in a copper solution, the zinc atoms lose electrons. The oxidation process that occurs, is represented as follows; $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$. On the other hand, the copper ions in solution are receiving (gaining) the electrons from the zinc atoms. This reduction process is represented as follows: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$. If this oxidation-reduction process between zinc and copper ions is allowed to continue for a longer time, loss of mass in the zinc plate will be observed while the blue colour of the solution, due to the presence of Cu^{2+} ions, will lose its colour.



All chemical reactions involve a change in energy as well as a change in chemical composition. In redox reactions, electrons are transferred from one substance to another while the substances themselves undergo changes. All these processes deal with changes in energy. The changes in energy during redox reactions can be summarised as follows: Redox reactions are exothermic

Chemical potential energy $\xrightleftharpoons{\text{Redox reaction}}$ Electrical potential energy

We distinguish two types of electrochemical cells as summarised below.



3. Two types of cells: electrolytic and galvanic cells

Both electrolytic and galvanic cells make use of redox reactions.

Electrolytic cells use electrical energy to initiate chemical reactions. Electrical energy is converted to chemical energy.

Galvanic (voltaic) cells use chemical reactions to produce electrical energy. Chemical energy is converted to electrical energy.

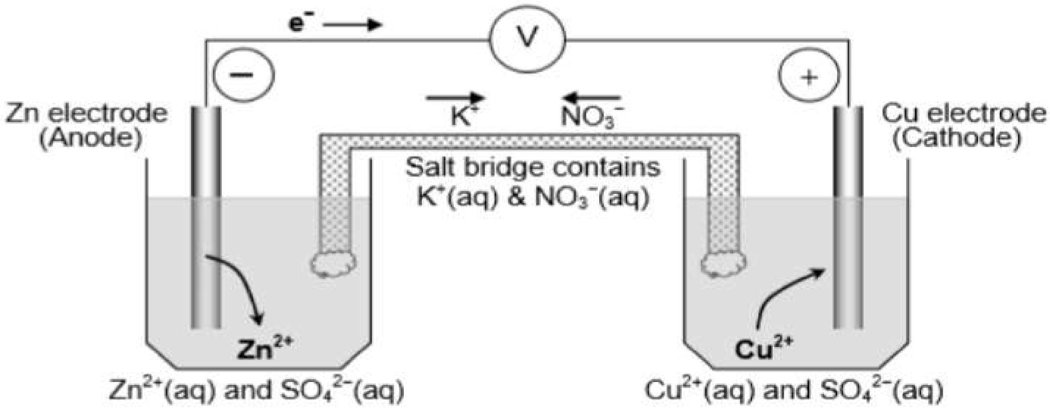
3.1 Galvanic cells (voltaic cells) indirect transfer of electrons

A galvanic cell is a cell in which chemical energy is converted to electrical energy. Galvanic cells make use of spontaneous redox reactions. The oxidising and reducing agents are separated and the electron transfer from the reducing agent to the oxidising agent takes place through an electrical conductor between them. The current in the conductor can be used to do useful work.

3.1.1 Composition and functioning of galvanic cells

Any galvanic cell consists of the following main parts: Two half-cells: The one half-cell contains the reducing agent and the other half-cell contains the oxidising agent. A salt bridge – an internal connection between the two half-cells. Ions can move freely through the salt bridge. An electrical conductor – an external connection between the two half-cells. Electrons flow through this conductor from the half-cell containing the reducing agent to the half-cell containing the oxidising agent.

A sketch of a copper-zinc galvanic cell is shown below.



The following components can be distinguished in the above galvanic cell:

- Zinc half-cell i.e. a Zn rod (called an electrode) placed into a Zn^{2+} solution - Zn is the reducing agent
- Copper half-cell i.e. a Cu rod (called an electrode) placed into a Cu^{2+} solution – Cu^{2+}(aq) is the oxidising agent

- Salt bridge – the internal circuit. The salt bridge consists of a U tube filled with an electrolyte e.g. KNO_3. It can also be a porous separation between the two half cells.
- An external circuit i.e. conducting wires connecting the two half-cells

In this cell, the Zn rod and the Cu rod are the electrodes. The reaction takes place between the Zn rod and the Cu^{2+} ions. The Cu rod serves as a link between the wire and the Cu^{2+} solution. No reaction will take place between the Cu rod and the Cu^{2+} solution. The Zn rod is placed in a Zn^{2+} solution – no reaction will take place between the Zn rod and the Zn^{2+} solution. The Zn electrode supplies electrons according to:

$Zn \rightarrow Zn^{2+} + 2e^-$ Therefore it is the negative terminal. Any electrode where oxidation takes place is called the anode in electrochemistry. Hence, the Zn electrode is the anode of this galvanic cell. The electrons flow from the Zn electrode through the metallic conductor to the Cu electrode. More Zn^{2+} ions enter the solution in the zinc half-cell, causing the Zn electrode to erode gradually (its mass decreases).

At the Cu electrode, the electrons react with the Cu^{2+} ions in the copper half-cell according to:

$Cu^{2+} + 2e^- \rightarrow Cu$ The Cu electrode is the positive terminal. The electrode where reduction takes place is called the cathode in electrochemistry. Hence, the Cu electrode is the cathode of this galvanic cell. The concentration of Cu^{2+} ions in the copper half-cell decreases, while the copper that is formed is deposited on the Cu electrode, causing the mass of the Cu electrode to gradually increase.

The reading on the voltmeter is positive if the negative terminal is connected to Zn.

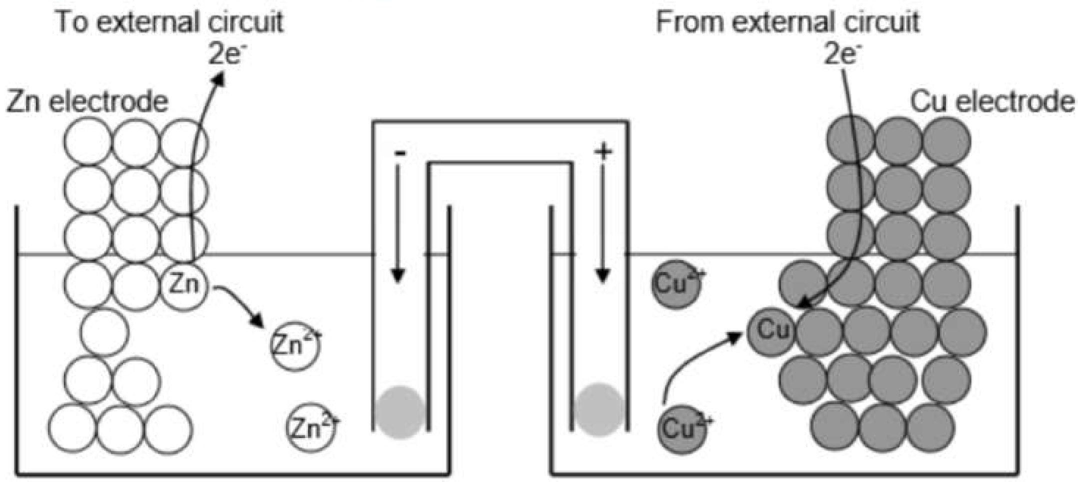
A positive reading on the voltmeter is therefore associated with a spontaneous redox reaction.

Functions of the salt bridge

- It completes the circuit. Electrons flow through the external conductor from the anode to the cathode of the cell.
- It maintain electrical neutrality within the solutions in the cell.
- It separates the two electrolytes

A salt bridge is a u-tuble consisting of soluble salts like $NaNO_3$, KNO_3, KCl

3.1.2 Microscopic events in a copper-zinc galvanic cell Electrons released during oxidation of Zn flow through the external conductor from the Zn electrode to the Cu electrode. At the Cu electrode, the Cu^{2+} ions accept the electrons to form Cu that is deposited on the Cu electrode. The diagram below illustrates these microscopic events in the cell.



Due to the oxidation of Zn, more Zn^{2+} ions enter the solution in the Zn half-cell. This will result in a built-up of positive ions in the Zn half-cell. Negative ions must be found to balance these newly generated positive ions. Otherwise the cell will stop functioning.

Reduction of Cu^{2+} ions in the Cu half-cell leaves behind negative ions that were associated with the Cu^{2+} ions. This will result in a built-up of negative ions in the Cu half-cell. Positive ions must be found to balance these excess in negative ions. The salt bridge is needed to prevent the building-up of positive and negative charges in the half-cells. It allows ions to move freely between the half-cells to keep both solutions neutral. Therefore, NO_3^- ions flow in the direction of the Zn electrode (the anode), while K^+ (and Zn^{2+} and Cu^{2+}) flow in the direction of the Cu electrode (the cathode). Positive ions move through the salt bridge to the cathode and negative ions move to the anode to keep the cell electrically neutral.

3.1.3 Cell reaction

The combination of the two reactions which take place in each of the half-cells produces the net cell reaction that takes place in the cell. negative electrode (anode):

$Zn \rightarrow Zn^{2+} + 2e^-$……………oxidation positive electrode (cathode):

$Cu^{2+} + 2e^- \rightarrow Cu$……………………..reduction net cell reaction:

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

3.1.4 Cell notation

A galvanic cell can be represented by abbreviated method, known as cell notation.

When writing cell notation, the following convention should be used:

- The $H_2|H^+$ half-cell is treated just like any other half-cell.
- Cell terminals (electrodes) are written on the outside of the cell notation.

- Active electrodes: reducing agent | oxidised species || oxidising agent | reduced species o Inert electrodes (usually Pt or C): Pt | reducing agent | oxidised species || oxidising agent | reduced species | Pt

Example: Pt |Cl^-(aq) | Cl_2(g) // F_2(g)/ F^-(aq) | Pt

The cell notation for the above copper-zinc galvanic cell is as follows:
Zn(s) | Zn^{2+}(aq) ‖ Cu^{2+}(aq) | Cu(s) Key parts of the notation are:

. If half-cell components are in the same phase e.g. H^+(aq), MnO_4^- (aq), they are separated by a comma.

3.1.5 STANDARD ELECTRODE POTENTIAL

The Standard Electrode Potential table ranks the half reactions from lowest to highest.

These are reduction potentials - the likelihood that the reaction would undergo reduction.

Increasing oxidising ability

Half-reactions			E^θ (V)
$Li^+ + e^-$	⇌	Li	− 3,05
$K^+ + e^-$	⇌	K	− 2,93
$Cs^+ + e^-$	⇌	Cs	− 2,92
$Ba^{2+} + 2e^-$	⇌	Ba	− 2,90
$Sr^{2+} + 2e^-$	⇌	Sr	− 2,89
$Ca^{2+} + 2e^-$	⇌	Ca	− 2,87
$Na^+ + e^-$	⇌	Na	− 2,71
$Mg^{2+} + 2e^-$	⇌	Mg	− 2,36
$A\ell^{3+} + 3e^-$	⇌	$A\ell$	− 1,66
$Mn^{2+} + 2e^-$	⇌	Mn	− 1,18
$Cr^{2+} + 2e^-$	⇌	Cr	− 0,91
$2H_2O + 2e^-$	⇌	$H_2(g) + 2OH^-$	− 0,83
$Zn^{2+} + 2e^-$	⇌	Zn	− 0,76
$Cr^{3+} + 3e^-$	⇌	Cr	− 0,74
$Fe^{2+} + 2e^-$	⇌	Fe	− 0,44
$Cr^{3+} + e^-$	⇌	Cr^{2+}	− 0,41
$Cd^{2+} + 2e^-$	⇌	Cd	− 0,40
$Co^{2+} + 2e^-$	⇌	Co	− 0,28
$Ni^{2+} + 2e^-$	⇌	Ni	− 0,27
$Sn^{2+} + 2e^-$	⇌	Sn	− 0,14
$Pb^{2+} + 2e^-$	⇌	Pb	− 0,13
$Fe^{3+} + 3e^-$	⇌	Fe	− 0,06
$2H^+ + 2e^-$	**⇌**	**$H_2(g)$**	**0,00**
$S + 2H^+ + 2e^-$	⇌	$H_2S(g)$	+ 0,14
$Sn^{4+} + 2e^-$	⇌	Sn^{2+}	+ 0,15
$Cu^{2+} + e^-$	⇌	Cu^+	+ 0,16
$SO_4^{2-} + 4H^+ + 2e^-$	⇌	$SO_2(g) + 2H_2O$	+ 0,17
$Cu^{2+} + 2e^-$	⇌	Cu	+ 0,34
$2H_2O + O_2 + 4e^-$	⇌	$4OH^-$	+ 0,40
$SO_2 + 4H^+ + 4e^-$	⇌	$S + 2H_2O$	+ 0,45
$Cu^+ + e^-$	⇌	Cu	+ 0,52

Increasing reducing ability

Notice how the reactions are ranked from left to right.

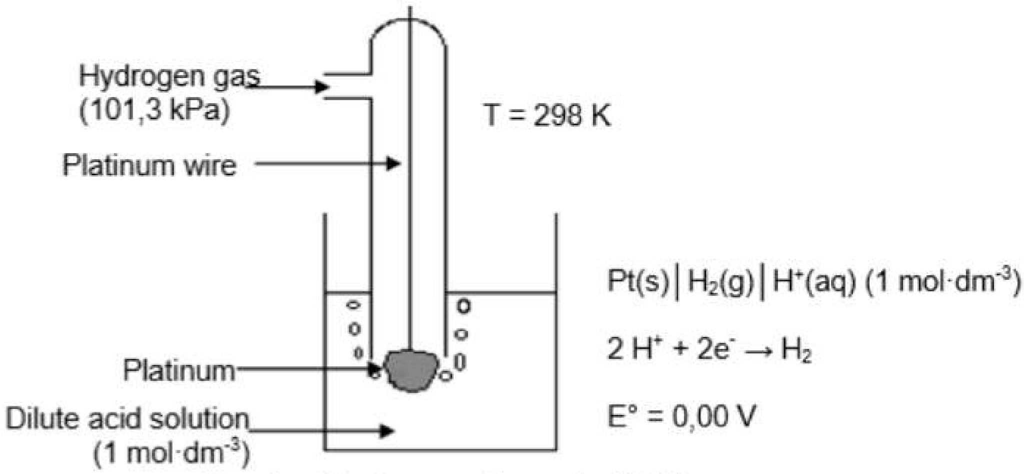
This table was manufactured under standard conditions. These conditions are:

- (Temperature) 25 °C/298 K
- (Concentration) 1 mol·dm^{-3}
- pressure 101,3 kpa

3.1.6 Standard hydrogen electrode

The standard hydrogen electrode is used as a reference electrode against which all other half-cells are measured. The reference half-reaction is the reduction of $H^+(aq)$ to $H_2(g)$ under standard conditions. A standard reduction potential of 0,00 V is allocated to this half-reaction: $2\,H^+(aq) + 2e^- \rightarrow H_2(g)$ oreduction E = 0,00 V

The standard hydrogen electrode, producing this half-reaction, consists of a platinum wire that is connected to a piece of platinum metal which serves as an inert electrode for the reaction. This electrode is in a glass tube. Hydrogen gas is bubbled over the platinum and through the 1 mol·dm^{-3} H^+ solution under standard conditions. A sketch of the hydrogen halfcell is shown below:



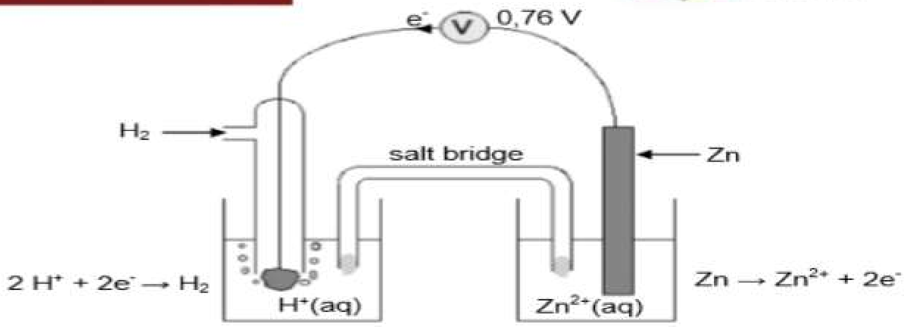
Standard hydrogen electrode (SHE)

Platinum is used as an electrode in the SHE because

- platinum is a conductor of electricity
- platinum is inert
- platinum has a high affinity for hydrogen

3.1.7 Compiling a list of standard reduction potentials

Standard reduction potentials are determined by connecting the half-cell in which a particular half-reaction takes place, to the standard hydrogen electrode (SHE). The diagram below shows a zinc half-cell connected to the SHE.



The spontaneous reaction, which takes place when a positive voltmeter reading is registered, involves the oxidation of Zn and the reduction of $H^+(aq)$. Zn is therefore the anode and the SHE is the cathode. The measured cell potential is + 0,76 V. Since the SHE has a reduction potential of 0,00 V the reduction potential of the Zn^{2+}/Zn half- $E^\ominus$cell = $E^\ominus$cathode – $E^\ominus$anode

$+ 0,76 = 0,00 - E^\ominus_{anode}$

$E^\ominus_{anode} = -0,76$ V

3.1.8 Cell potential

By using the numbers on the table, we can calculate the voltage, or emf, that we can expect from a combination of two electrodes.

$E_{cell} = E_{cathode} - E_{anode}$

So, first you need to find out where oxidation is taking place (anode) and then reduction (cathode).

Hint: Always substitute using brackets as the signs may need to change.

3.1.9 Reduction potentials.

Reduction potential is the tendency for a chemical species to gain electrons.

Oxidation potential is the tendency for a chemical species to lose electrons(to be oxidised) E_{cell} = reduction potential + oxidation potential

An oxidation half reaction is the reverse reaction of the particular reduction half reaction. Therefore the oxidation potential of a particular half reaction is equal to the negative of the reduction potential.

Oxidation potential for half reaction = – reduction potential for half reaction

3.1.9 Using the tables of standard reduction potentials to predict cell potential

With the help of the information from Table 4A or 4B, there are two ways to calculate the cell potential of a galvanic cell under standard conditions.

Method 1: Using the derived formula for cell potential

- Identify the relevant half-reactions
- Determine from this the relevant oxidising and reducing agents (or oxidation and reduction half-reactions, or the half-reactions taking place at the anode and cathode)
- Substitute the E° values as they appear on the table into one of the formulae above.

Example 1

A voltaic cell is constructed using $Ni^{2+}|Ni$ and $Cu^{2+}|Cu$ half-cells. Determine the cell potential at standard conditions and write down the cell notation.

Solution From the Table of Standard Reduction potentials:

$Ni^{2+} + 2e^- \leftrightharpoons Ni$ (-0,25 V)

$Cu^{2+} + 2e^- \leftrightharpoons Cu$ (+0,34 V)

Method 1:

The copper half-reaction has the more positive cell potential and thus Cu2+ is a stronger oxidising agent than Ni2+. The copper half-reaction will be reduction and the nickel half-reaction will be the oxidation.

The half-reactions will therefore take place as follows:

$Ni \rightarrow Ni^{2+} + 2e^-$ (+0,25 V) (Note: If the reduction potential is -0,25 V, the oxidation potential is +0,25 V

$Cu^{2+} + 2e^- \rightarrow Cu$ (+0,34 V)

E°cell = 0,59 V

Method 2: From this: Cu^{2+} is a stronger oxidising agent than Ni^{2+}. Therefore, Cu^{2+} behaves as the oxidising agent and Ni as the reducing agent.

E°cell = E°oxidising agent - E°reducing agent = +0,34 - (-0,25) = 0,59 V

For the cell notation: Ni undergoes oxidation; therefore Ni is the anode.

$Ni(s) \mid Ni^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu$

Example 2

Consider the following unbalanced equation that takes place in an electrochemical cell: $Ag(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Ag^+(aq)$

Use the cell potential to determine whether this reaction proceeds spontaneously from left to right at standard conditions.

Solution According to the information in the equation, Ag is oxidised to Ag+ and Cu2+ is reduced to Cu. Therefore, Ag is the reducing agent and Cu^{2+} the oxidising agent.

E°cell = E°oxidising agent - E°reducing agent

= +0,34 - (+0,80) = -0,46 V

OR The E° value for the oxidation of Ag is -0,80 V The E° value for the reduction of Cu2+ is +0,34 V

E°cell = -0,46 V

The reaction cannot take place spontaneously.

3.1.10 Factors which influence cell potential

The cell potential of a galvanic cell depends on the following factors:

- Specific reactions that take place at the anode and the cathode
- Concentrations of reactants and products
- Temperature

Specific reactions at the anode and cathode

The relative strength of oxidising and reducing agents influences the cell potential. The stronger the oxidising and reducing agent that are combined, the greater the cell potential will be.

Concentration of reactants and products Under standard conditions the concentration of all solutions in a galvanic cell is 1 $mol{\cdot}dm^{-3}$. The galvanic cell described by the cell reaction below has a cell potential of 0,48 V.

$2Al(s) + 3Mn^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Mn(s)$

If the concentration of the Mn^{2+} ions increases (greater than1 $mol{\cdot}dm^{-3}$) the forward reaction will be favoured as can be predicted by using Le Chatelier"s principle. Consequently the cell potential will be greater than 0,48 V.

If the concentration of the Al^{3+} ions increases (greater than 1 $mol{\cdot}dm^{-3}$) the reverse reaction will be favoured as can be predicted by using Le Chatelier"s principle. Consequently the cell potential will be less than 0,48 V.

Initially the concentration of the reactants in a galvanic cell is high and the cell potential is at its maximum. As the reaction progresses the concentration of reactants decreases and the concentration of products increases. The result is a decrease in cell potential.

We can therefore say that an increase in the concentration of the reactants will increase the cell potential and that an increase in the concentration of the products will decrease the cell potential. When equilibrium is reached, the cell potential has dropped to 0 V.

Temperature Electrochemical reactions in galvanic cells are all exothermic reactions. According to Le Chatelier"s principle an increase in temperature will favour the endothermic reaction, i.e. the reverse reaction, and lead to a drop in cell potential.

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TYPICAL EXAM QUESTIONS

QUESTION 1 MULTIPLE CHOICE QUESTIONS

1.1 Consider the cell notation of the galvanic cell below.

$$Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$$

Which ONE of the following statements regarding this cell is TRUE?

A Copper is formed at the cathode.

B Copper is formed at the anode.

C Zinc is formed at the anode.

D Zinc is formed at the cathode. (2)

1.2 Which ONE of the following metals will NOT react spontaneously with sulphuric acid?

A Zn B Mg C Cu D Fe (2)

1.3 Which ONE of the following CANNOT act as a reducing agent?
A Mg B Fe^{2+} C Br^- D MnO_4^- (2)

1.4 An electrochemical cell is used to electroplate an iron spoon with nickel.

Which ONE of the following half-reactions takes place at the positive electrode of this cell?

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

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

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

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

1.5 Consider the reaction represented by the balanced equation below:

$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

In the above reaction, Cu(s) is the ...

A oxidising agent and is reduced.

B oxidising agent and is oxidised.

C reducing agent and is reduced.

D reducing agent and is oxidised. (2)

1.6 The following half-reactions take place in a galvanic cell:

$Co^{3+} + e^- \rightleftharpoons Co^{2+}$
$A\ell^{3+} + 3e^- \rightleftharpoons A\ell$

Which ONE of the following is the cell notation for this cell?

A $A\ell \mid A\ell^{3+} \parallel Co^{3+}, Co^{2+}$

B $A\ell \mid A\ell^{3+} \parallel Co^{3+}, Co^{2+} \mid Pt$

C $A\ell \mid A\ell^{3+} \parallel$ Co2+, $Co^{3+} \mid Pt$

D $Pt \mid Co^{2+}, Co^{3+} \parallel A\ell^{3+} \mid A\ell$ (2)

1.7 Chlorine gas ($C\ell_2$) is bubbled through a potassium iodide solution (KI). The reducing agent in this reaction is:

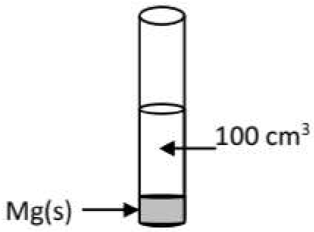
A Potassium ions

B Chlorine gas

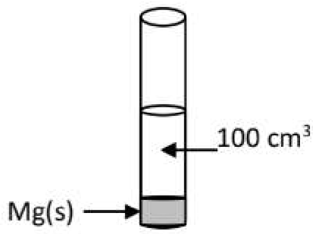
C Iodide ions

D Chloride ions (2)

1.8 Equal amounts of magnesium (Mg) powder react respectively with equal volumes and equal concentrations of $HC\ell(aq)$ and $H_2SO_4(aq)$, as shown below.



Test tube X



Test tube Y

The magnesium is in EXCESS.

Consider the following statements regarding these two reactions:

I: The initial rate of the reaction in test tube **X** equals the initial rate of the reaction in test tube **Y**.

II: After completion of the reactions, the mass of magnesium that remains in test tube **X** will be greater than that in test tube **Y**.

III: The amount of hydrogen gas formed in **X** is equal to the amount of hydrogen gas formed in **Y**.

Which of the above statements is/are TRUE?

A **I** only

B **II** only

C **III** only

D **I** and **III** only (2)

1.9 Which ONE of the following is a NON-SPONTANEOUS redox reaction? Refer to the Table of Standard Reduction Potentials (Table 4A or 4B).

A $Zn(s) + 2HC\ell(aq) \rightarrow ZnC\ell_2(aq) + H_2(g)$

B $Cu(s) + FeC\ell_2(aq) \rightarrow CuC\ell2(aq) + Fe(s)$

C $2AgNO3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

D $2A\ell(s) + 3Ni(NO_3)_2(aq) \rightarrow 2A\ell(NO_3)_3(aq) + 3Ni(s)$ (2)

1.10 A galvanic cell consists of the following half-cells: $Pt(s)\,|C\ell_2(g)\,|C\ell^-(aq)$ AND $Cu^{2+}(aq)\,|Cu(s)$. Which ONE of the following statements is TRUE while the cell is functioning?

A Cu(s) is oxidised.

B $C\ell^-(aq)$ is reduced.

C $C\ell_2(g)$ acts as reducing agent.

D Cu(s) acts as oxidising agent. (2)

QUESTION 2

An electrochemical cell consisting of half-cells A and B is assembled under standard conditions as shown below.

Half-cell **A**	Pt, $C\ell_2$ (101,3 kPa) | $C\ell^-$ (1 $mol{\cdot}dm^{-3}$)
Half-cell **B**	Mg^{2+} (1 $mol{\cdot}dm^{-3}$) | Mg(s)

2.1 At which half-cell, A or B, are electrons released into the external circuit? (1)

2.2 Write down the:

2.2.1 Reduction half-reaction that takes place in this cell (2)

2.2.2 NAME or FORMULA of the substance whose oxidation number DECREASES (1)

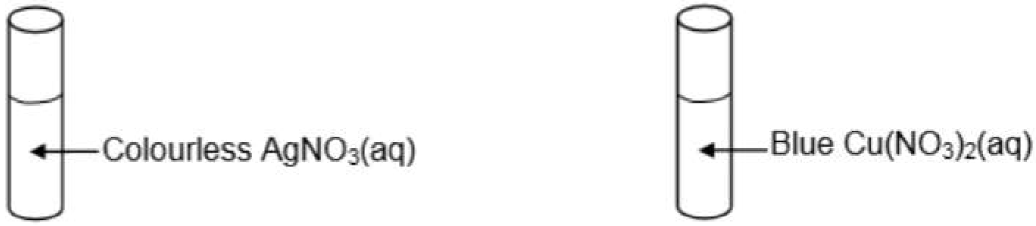
2.3 Calculate the initial cell potential of this cell when it is in operation. (4)

QUESTION 3

A learner conducts two experiments to investigate the reaction between copper (Cu) and a silver nitrate solution, AgNO3(aq).

EXPERIMENT 1 The learner adds a small amount of copper (Cu) powder to a test tube containing silver nitrate solution, $AgNO_3(aq)$. The solution changes from colourless to blue after a while.

Before addition of Cu(s) **After addition of Cu(s)**

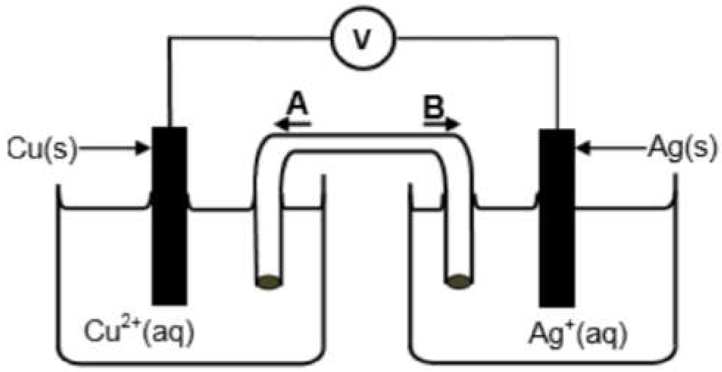


3.1 Define the term oxidising agent. (2)

3.2 Explain why the solution turns blue by referring to the relative strength of oxidising agents. (4)

EXPERIMENT 2

The learner now sets up a galvanic cell as shown below. The cell functions under standard conditions.



3.3 Write down the energy conversion that takes place in this cell. (1)

3.4 In which direction (A or B) will ANIONS move in the salt bridge? (1)

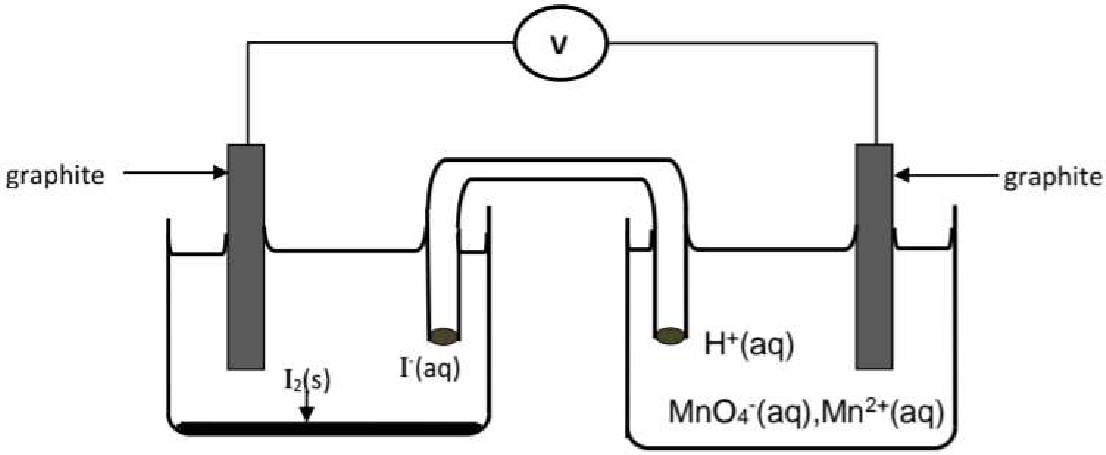
3.5 Calculate the emf of the above cell under standard conditions. (4)

3.6 Write down the balanced equation for the net cell reaction that takes place in this cell. (3)

3.7 How will the addition of 100 cm^3 of a 1 $mol \cdot dm^{-3}$ silver nitrate solution to the silver half-cell influence the initial emf of this cell? Write down only INCREASES, DECREASES or REMAINS THE SAME. (1)

QUESTION 4

The voltaic cell represented below functions at standard conditions.



4.1 Write down the concentration of the $H^+(aq)$ in the one half-cell. (1)

4.2 Solids present in half-cells are usually used as electrodes. Give a reason why I2(s) is not suitable to be used as electrode. (1)

4.3 Write down TWO properties, other than being a solid, of graphite that makes it suitable to be used as electrodes in the above voltaic cell. (2)

4.4 For the above voltaic cell, write down the:

4.4.1 NAME of the oxidising agent (1)

4.4.2 Net cell reaction (3)

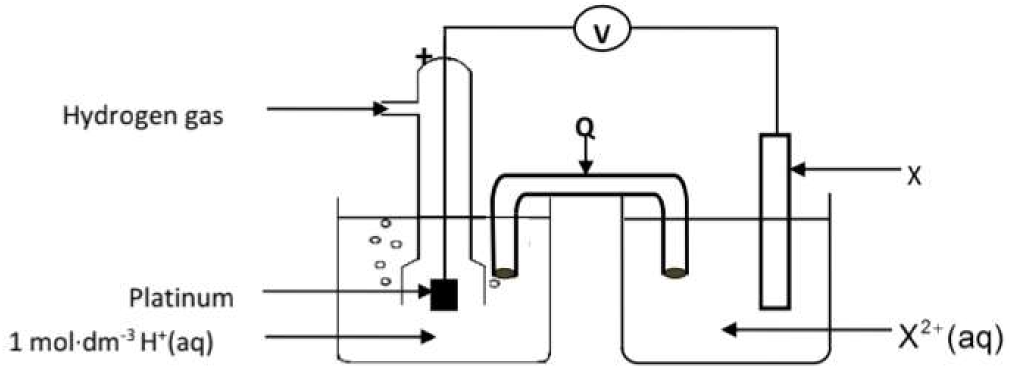
4.4.3 Cell notation (3)

4.5 Calculate the cell potential of the above cell. (4)

4.6 How will the reading on the voltmeter be affected if the concentration of decreases? Only write down INCREASES, DECREASES or NO EFFECT. (1)

QUESTION 5

A standard electrochemical cell is set up using a standard hydrogen half-cell and a standard $X|X^{2+}$ half-cell as shown below. A voltmeter connected across the cell, initially registers 0,31 V.



5.1 Besides concentration write down TWO conditions needed for the hydrogen half-cell to function under standard conditions. (2)

5.2 Give TWO reasons, besides being a solid, why platinum is suitable to be used as electrode in the above cell. (2)

5.3 Write down the:

5.3.1 NAME of component **Q** (1)

5.3.2 Standard reduction potential of the $X|X^{2+}$ half-cell (1)

5.3.3 Half-reaction that takes place at the cathode of this cell (2)

5.4 The hydrogen half-cell is now replaced by a $\mathbf{M}|\mathbf{M}^{2+}$ half-cell. The cell notation of this cell is:

$$\mathbf{M}(s) \mid \mathbf{M}^{2+}(aq) \parallel X^{2+}(aq) \mid X(s)$$

The initial reading on the voltmeter is now 2,05 V.

5.4.1 Identify metal **M**. Show how you arrived at the answer. (5)

5.4.2 Is the cell reaction EXOTHERMIC or ENDOTHERMIC? (1)

5.5 The reading on the voltmeter becomes zero after using this cell for several hours. Give a reason for this reading by referring to the cell reaction. (1)

[15]

QUESTION 6

Learners are given the following two unknown half-cells:

Half-cell 1: Q^{2+} (aq) | Q(s)

Half-cell 2: Pt | R_2(g) | R^-(aq)

During an investigation to identify the two half-cells, the learners connect each half-cell alternately to a Cd^{2+}(aq) | Cd(s) half-cell under standard conditions. For each combination of two half-cells, they write down the net cell reaction and measure the cell potential.

The results obtained for the two half-cell combinations are given in the table below.

COMBINATION	NET CELL REACTION	CELL POTENTIAL
I	$\mathbf{Q}^{2+}$(aq) + Cd(s) → Cd^{2+}(aq) + **Q**(s)	0,13 V
II	$\mathbf{R}_2$(g) + Cd(s) → Cd^{2+}(aq) + 2$\mathbf{R}^-$(aq)	1,76 V

6.1 Write down THREE conditions needed for these cells to function as standard cells. (3)

6.2 For **Combination I**, identify:

6.2.1 The anode of the cell (1)

6.2.2 **Q** by using a calculation (5)

6.3 For **Combination II**, write down the:

6.3.1 Oxidation half-reaction (2)

6.3.2 NAME or FORMULA of the metal used in the cathode compartment (1)

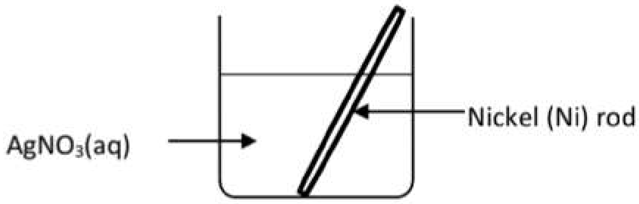
6.4 Arrange the following species in order of INCREASING oxidising ability:

Q^{2+} ; R_2 ; Cd^{2+}

Explain fully how you arrived at the answer. A calculation is NOT required. (4)

QUESTION 7

7.1 A nickel (Ni) rod is placed in a beaker containing a silver nitrate solution, $AgNO_3(aq)$ and a reaction takes place.



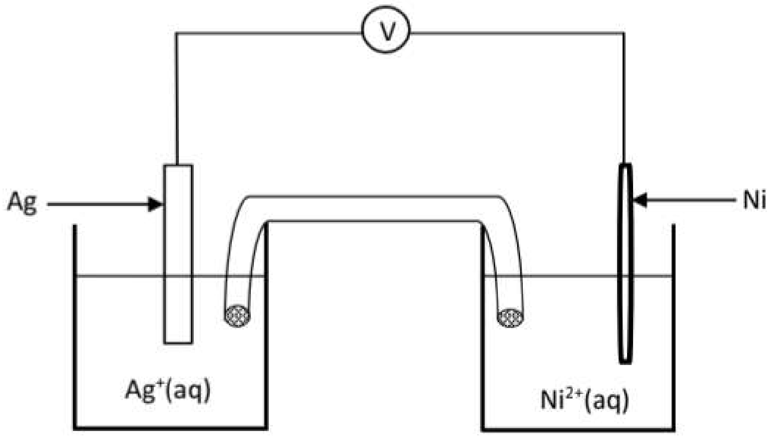
Write down the:

7.1.1 NAME or FORMULA of the electrolyte (1)

7.1.2 Oxidation half-reaction that takes place (2)

7.1.3 Balanced equation for the net (overall) redox reaction that takes place (3)

7.2 A galvanic cell is now set up using a nickel half-cell and a silver half-cell.



7.2.1 Which electrode (**Ni** or **Ag**) must be connected to the negative terminal of the voltmeter? Give a reason for the answer. (2)

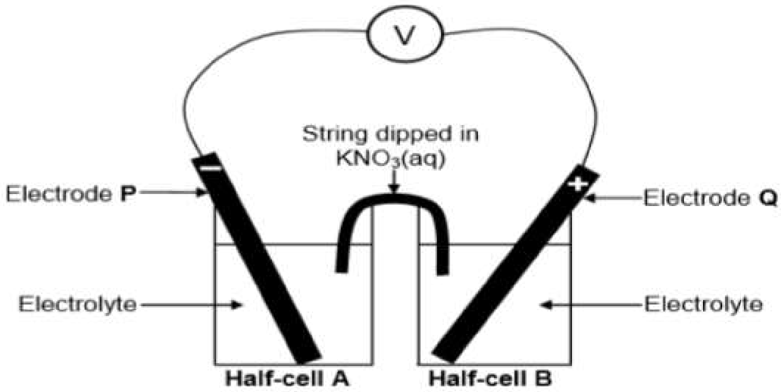
7.2.2 Write down the cell notation for the galvanic cell above. (3)

7.2.3 Calculate the initial reading on the voltmeter if the cell functions under standard conditions. (4)

7.2.4 How will the voltmeter reading in QUESTION 7.2.3 be affected if the concentration of the silver ions is increased? Choose from INCREASES, DECREASES or REMAINS THE SAME. (1)

QUESTION 8

Learners set up an electrochemical cell, shown in the simplified diagram below, using magnesium and lead as electrodes. Nitrate solutions are used as electrolytes in both half-cells.



8.1 What type of reaction (NEUTRALISATION, REDOX or PRECIPITATION) takes place in this cell? (1)

8.2 Which electrode, P or Q, is magnesium? Give a reason for the answer. (2)

8.3 Write down the:

8.3.1 Standard conditions under which this cell functions (2)

8.3.2 Cell notation for this cell (3)

8.3.3 NAME or FORMULA of the oxidising agent in the cell (1)

8.4 Calculate the initial emf of the cell above under standard conditions. (4)

8.5 How will the voltmeter reading change if the: (Write down only INCREASES, DECREASES or REMAINS THE SAME.)

8.5.1 Size of electrode P is increased (1)

8.5.2 Initial concentration of the electrolyte in half-cell B is increased (1)

QUESTION 9

Magnesium (Mg) reacts with a dilute hydrochloric acid solution, HCℓ(aq), according to the following balanced equation:

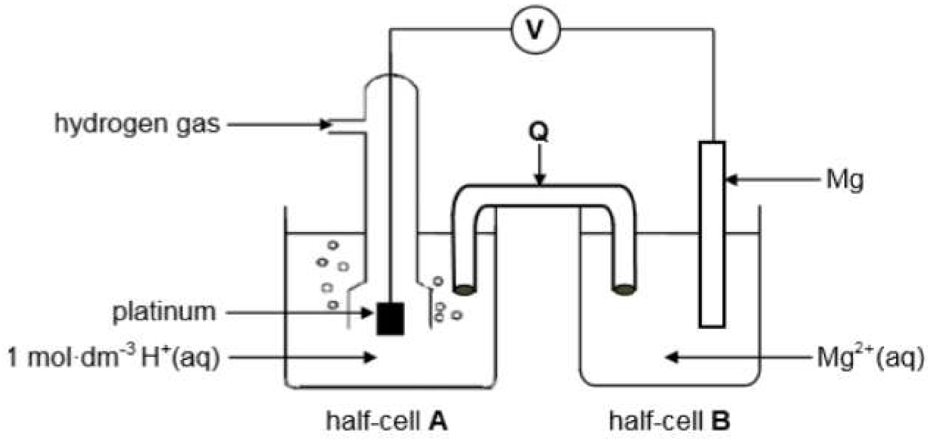
$Mg(s) + 2HC\ell(aq) \rightarrow MgC\ell_2(aq) + H_2(g)$

9.1 Give a reason why the reaction above is a redox reaction. (1)

9.2 Write down the FORMULA of the oxidising agent in the reaction above. (1)

It is found that silver does not react with the hydrochloric acid solution.

9.3 Refer to the relative strengths of reducing agents to explain this observation. (3)
The reaction of magnesium with hydrochloric acid is used in an electrochemical cell, as shown in the diagram below. The cell functions under standard conditions.



9.4 What is the function of platinum in the cell above? (1)

9.5 Write down the:

9.5.1 Energy conversion that takes place in this cell (1)

9.5.2 Function of Q (1)

9.5.3 Half-reaction that takes place at the cathode (2)

9.5.4 Cell notation of this cell (3)

9.6 Calculate the initial emf of this cell. (4)

9.7 How will the addition of concentrated acid to half-cell A influence the answer to QUESTION 9.6? Choose from INCREASES, DECREASES or REMAINS THE SAME. (1)

Electrolytic cells

The electrolytic cell is a cell in which electrical energy is converted into chemical energy. The process during which electrical energy is converted into chemical energy is called 'electrolysis'.

Electrolysis: The chemical process in which electrical energy is converted to chemical energy OR the use of electrical energy to produce a chemical change.

An electrolyte as a solution/liquid/dissolved substance that conducts electricity through the movement of ions

In electrolytic cells, oxidation-reduction reactions take place in a direction in which they do NOT OCCUR SPONTANEOUSLY. Electrical energy is used to drive a non-spontaneous reaction.

For electrolysis to take place, the following is necessary:

- Two electrodes, either inert (does not take part in the reaction) or active (takes part in the reaction)
- An electrolyte - a substance that dissolves in water to form positive (cations) and negative ions (anions) OR a substance having free positive (cations) and negative ions (anions) when melted
- A source of direct electric current, e.g. cells or a battery

The battery acts as an 'electron pump', pulling electrons from one electrode and pushing electrons into the other electrode. Withdrawing electrons from an electrode gives the electrode a positive charge, and adding electrons to an electrode gives it a negative charge.

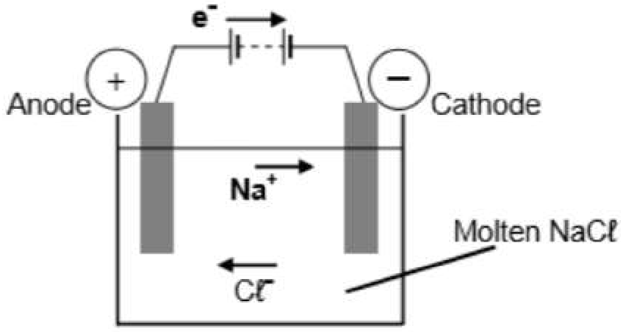
In an electrolytic cell the battery therefore determines which electrode is negative and which one is positive (polarity).

Ionic substances in the solid state cannot conduct an electric current. The electrolyte is obtained by melting an ionic substance or by dissolving it in water in order to enable the cations (positive ions) and anions (negative ions) to move freely.

3.2.1 Electrolysis of molten salts using inert electrodes Inert electrodes do not take part in the electrolysis reaction. Carbon or platinum are usually used as inert electrodes.

3.2.1.1 Electrolysis of molten sodium chloride

The diagram below shows a simple electrolytic cell using molten sodium chloride as electrolyte. Wires from a battery are connected to electrodes dipped into molten sodium chloride. The electrodes are inert and do not take part in the reaction.



$2C\ell^- \rightarrow C\ell_2 + 2e^-$ $\quad$ $2Na^+ + 2e^- \rightarrow 2Na$

- In the above diagram, electrons flow from the negative terminal of the battery. The electrode connected to the negative terminal of the battery is the negative electrode. The electrode connected to the positive terminal of the battery is the positive electrode.
- The electrolyte (NaCℓ) contains positive (Na^+) and negative ($C\ell^-$) ions that are, when melted, free to move around in the solution. These are the only ions present in the melt.
- When electrons flow in the external circuit, negative ions in the electrolyte move to /are attracted to the positive electrode. In the above cell, the $C\ell^-$ ions will move to the positive electrode.
- When electrons flow in the external circuit, positive ions in the electrolyte move to /are attracted to the negative electrode. In the above cell, the Na^+ ions will move to the negative electrode.
- At the positive electrode: Negative ions lose electrons at the positive electrode according to the following half-reaction: $2C\ell^- \rightarrow C\ell_2 + 2e^-$
  A loss of electrons is oxidation. Oxidation takes place at the anode – therefore the positive electrode in an electrolytic cell is the anode.
- Chlorine is a product of the oxidation. Chlorine is a gas – bubbles will thus be observed at the positive electrode.
- At the negative electrode: The negative electrode has an excess of electrons. The positive ions gain electrons at the negative electrode according to the following halfreaction: $Na^+ + e^- \rightarrow Na$ A gain of electons is a reduction. Reduction takes place at the cathode – therefore the negative electrode in an electrolytic cell is the cathode.
- Sodium is a product of the reduction. Sodium is a solid and will be deposited on the negative electrode. The mass of the negative electrode will increase.
- The net cell reaction is obtained by balancing the two half-reactions:

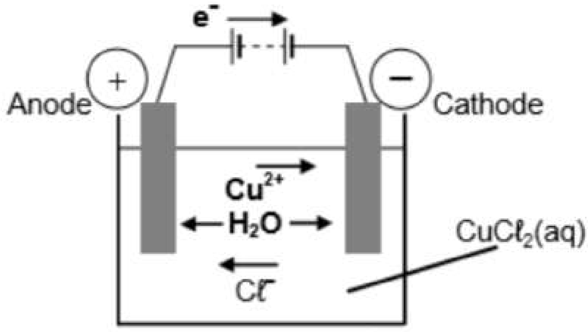
$2Na^+ + 2e^- \rightarrow 2Na$ (Reduction; with Na^+ the oxidising agent)
$2Cℓ^- \rightarrow Cℓ_2 + 2e^-$ (Oxidation; with $Cℓ^-$ the reducing agent)

$2Na^+(ℓ) + 2Cℓ^-(ℓ) \rightarrow 2Na(ℓ) + Cℓ_2(g)$

- The products of the electrolysis of molten NaCℓ are liquid sodium metal (Na) and chlorine gas ($Cℓ_2$). A specially designed cell, called the 'Downs cell' is used to prepare these products commercially.

3.2.1.2 Electrolysis of concentrated copper (II) chloride solution

In the simplified electrolytic cell below, an electric current passes through a solution of copper (II) chloride in water. Cu^{2+} ions, $Cℓ^-$ ions and water molecules are present in the solution. Carbon rods are used as electrodes.



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

Part of redox Table

H_2O	$+ 2e^- \rightleftharpoons$	$H_2 + 2OH^-$	-0,83
Zn^{2+}	$+2e^- \rightleftharpoons$	Zn	-0,76
Cr^{3+}	$+3e^- \rightleftharpoons$	Cr^+	-0,74
Fe^{2+}	$+2e^- \rightleftharpoons$	Fe	-0,44
Cu^{2+}	$+ e^- \rightleftharpoons$	Cu^+	+0,16
$SO_4^{2-} + 4H^+$	$+2e^- \rightleftharpoons$	$SO_2 + 2H_2O$	+0,17
Cu^{2+}	$+ 2e^- \rightleftharpoons$	Cu	+0,34
Cl_2	$+ 2e^- \rightleftharpoons$	$2Cl^-$	+1,36

Two half-reactions are:

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

$2Cℓ^- \rightarrow Cℓ_2 + 2e^-$ (Oxidation)

The net cell reaction is therefore:

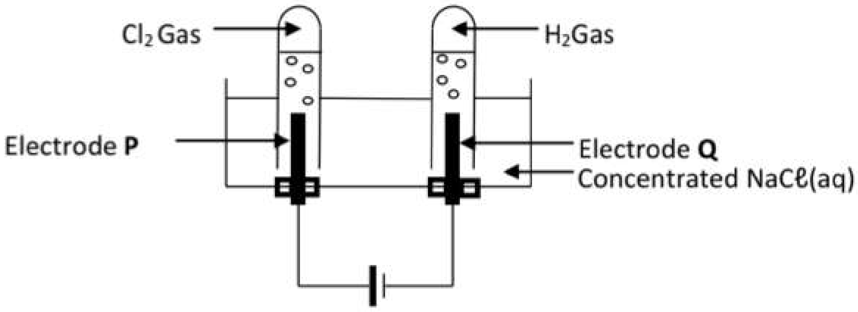
$$Cu^{2+} + 2e^- \rightarrow Cu$$

$$2C\ell^- \rightarrow C\ell_2 + 2e^-$$

$$Cu^{2+}(aq) + 2C\ell^-(aq) \rightarrow Cu(s) + C\ell_2(g)$$

3.2.1.3 Electrolysis of concentrated solution of sodium chloride (brine)

Na^+ ions, $C\ell^-$ ions and water molecules are present in the solution. Carbon rods are used as electrodes.



Part of redox Table

$Na^+ + e^-$	$\rightleftharpoons$	Na	-2,71
$Mg^{2+} + 2e^-$	$\rightleftharpoons$	Mg	-2,36
$H_2O + 2e^-$	$\rightleftharpoons$	$H_2 + H_2 + 2OH^-$	-0,83
$Zn^{2+} + 2e^-$	$\rightleftharpoons$	Zn	-0,76
$Cr^{3+} + 3e^-$	$\rightleftharpoons$	Cr^+	-0,74
$Fe^{2+} + 2e^-$	$\rightleftharpoons$	Fe	-0,44
$Cu^{2+} + e^-$	$\rightleftharpoons$	Cu^+	+0,16
$SO_4^{2-} + 4H^+ + 2e^-$	$\rightleftharpoons$	$SO_2 + 2H_2O$	+0,17
$Cu^{2+} + 2e^-$	$\rightleftharpoons$	Cu	+0,34
$Cl_2 + 2e^-$	$\rightleftharpoons$	$2Cl^-$	+1,36
$H_2O + 2H^+ + 2e^-$	$\rightleftharpoons$	$2H_2O$	+1,77

$H_2O + 2e^- \rightleftharpoons H_2 + H_2 + 2OH^-$ (Reduction) Cathode

$2Cl^- \rightleftharpoons Cl_2 + 2e^-$ (Oxidation) Anode

Net Reaction

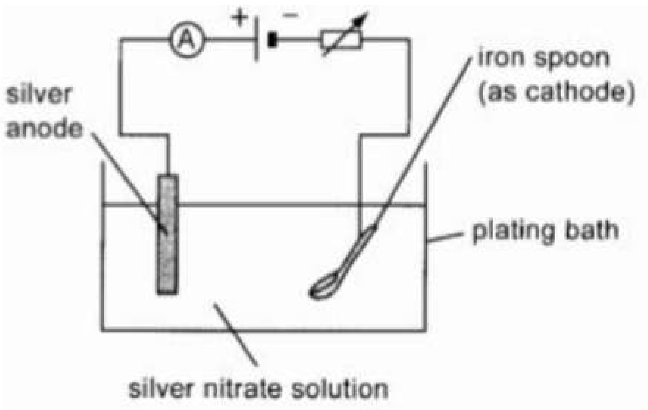
$2H_2O + 2Cl^- \longrightarrow H_2 + Cl_2 + 2OH^-$

Major reactions in the chlor-alkali industry

Industrial Applications of Electrolysis

Electroplating: It is a process to plate one metal with a thin, even layer of another metal. A metal is electroplated to protect against corrosion and/or to improve appearance.

Refer to the following diagram to plate an iron spoon with silver:

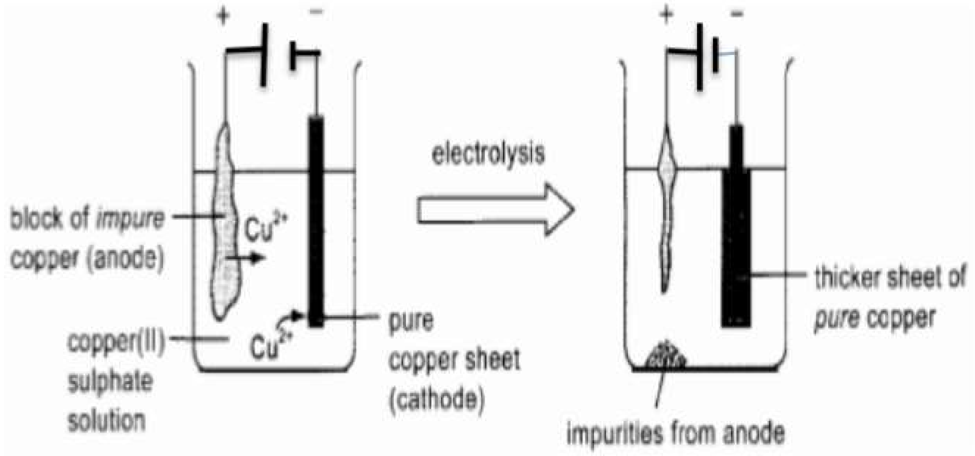


- The metal object to be electroplated (the iron spoon) is made the cathode.
- The plating metal (silver) is made the anode.
- The electrolytic cell contains the plating metal solution (silver nitrate) as the electrolyte.
  At the anode
  $Ag(s) \rightarrow Ag^+(aq) + e^-$
  At the cathode
  $Ag^+(aq) + e^- \rightarrow Ag(s)$

PURIFICATION OF METALS

Electrolytic refining could be used to purify metals such as lead and copper.

In the refining of copper, an aqueous solution of copper (II) sulphate is used as the electrolyte. A block of impure copper is used as the anode. A sheet of pure copper is used as the cathode. Refer to the following diagram:



At the anode

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

At the cathode

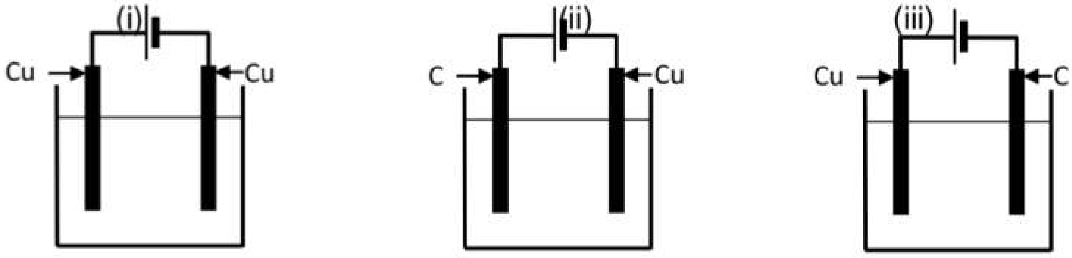
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Impure metals from the impure copper block will fall down the electrolyte as sludge.

TYPICAL EXAM QUESTIONS

QUESTION 1 MULTIPLE CHOICE QUESTIONS

1.1 In each of the electrolytic cells below, copper (II) sulphate is used as the electrolyte. The electrodes are either carbon (C) or copper (Cu).



In which cell(s) will the concentration of the electrolyte remain constant during electrolysis?

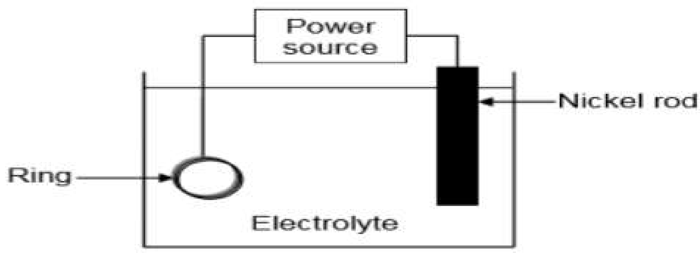
A (i) only

B (i) and (ii) only

C (i) and (iii) only

D (ii) and (iii) only (2)

1.2 A learner wants to electroplate a copper ring with nickel. He uses the experimental set-up shown in the simplified diagram below.



Which ONE of the following is CORRECT?

	ANODE	CATHODE	ELECTROLYTE
A	Copper ring	Nickel rod	$CuSO_4$
B	Nickel rod	Copper ring	$CuSO_4$
C	Copper ring	Nickel rod	$NiSO_4$
D	Nickel rod	Copper ring	$NiSO_4$

(2)

1.3 A sample of silver contains impurities of gold. During purification by electrolysis, the impure silver is made into an electrode.

Which ONE of the following is the best choice of anode and cathode for this process?

	Cathode	Anode
A	Pure gold	Impure silver
B	Impure silver	Pure gold
C	Pure silver	Impure silver
D	Impure silver	Pure silver

(2)

1.4 Which ONE of the equations below represents the half-reaction occurring at the CATHODE of an electrochemical cell that is used to electroplate an object?

A $Ag \rightarrow Ag^{+} + e^{-}$

B $Cr^{3+} + 3e^{-} \rightarrow Cr$

C $Cr^{3+} + e^{-} \rightarrow Cr^{2+}$

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

1.5 Consider the reaction represented by the balanced equation below:

Cu(s) + 2Ag+(aq) → Cu2+(aq) + 2Ag(s)

In the above reaction, Cu(s) is the ...

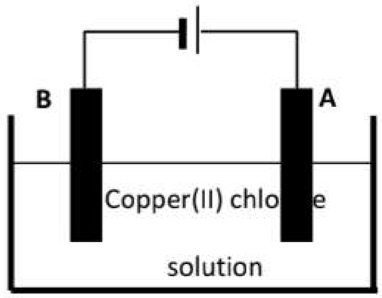
A oxidising agent and is reduced. B oxidising agent and is oxidised.

C reducing agent and is reduced. D reducing agent and is oxidised. (2)

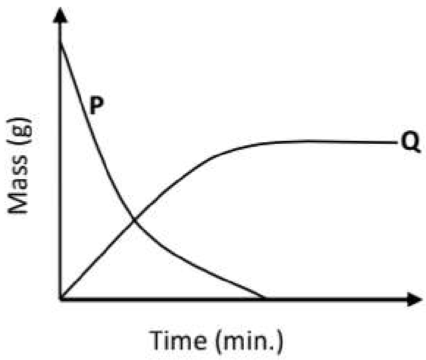
QUESTION 2

The electrochemical cell below is set up to demonstrate the purification of copper.

2.1 Write down the type of electrochemical cell illustrated above. (1)



The graphs below show the change in mass of the electrodes whilst the cell is in operation.



2.2 Define a reducing agent in terms of electron transfer. (2)

2.3 Which graph represents the change in mass of electrode A? (1)

2.4 Write down the half-reaction that takes place at electrode A. (2)

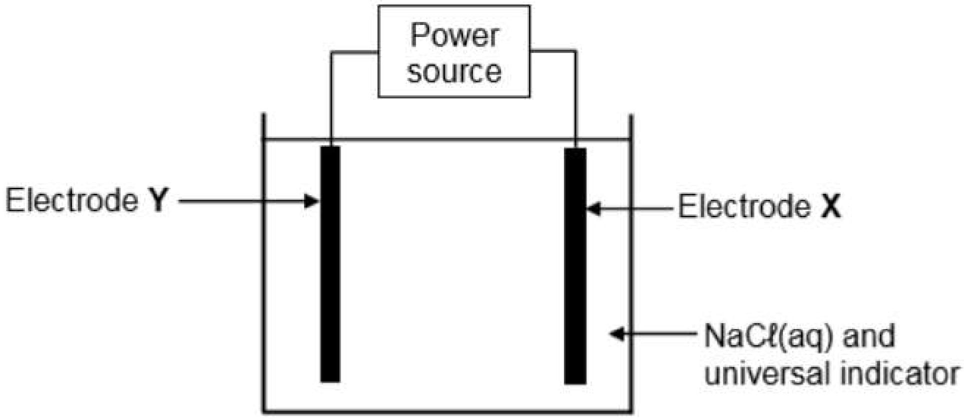
2.5 Electrodes A and B are now replaced by graphite electrodes. It is observed that chlorine gas (Cl_2) is released at one of the electrodes.

At which electrode (A or B) is chlorine gas formed? Fully explain how it is formed.

QUESTION 3

The apparatus below is used to demonstrate the electrolysis of a concentrated sodium chloride solution. Both electrodes are made of carbon. A few drops of universal indicator are added to the electrolyte. The equation for the net cell reaction is:

$2NaC\ell(aq) + 2H_2O(\ell) \rightarrow C\ell_2(g) + H_2(g) + 2NaOH(aq)$



Initially the solution has a green colour. Universal indicator becomes red in acidic solutions and purple in alkaline solutions.

3.1 Define the term electrolyte. (2)

When the power source is switched on, the colour of the electrolyte around electrode Y changes from green to purple.

3.2 Write down the:

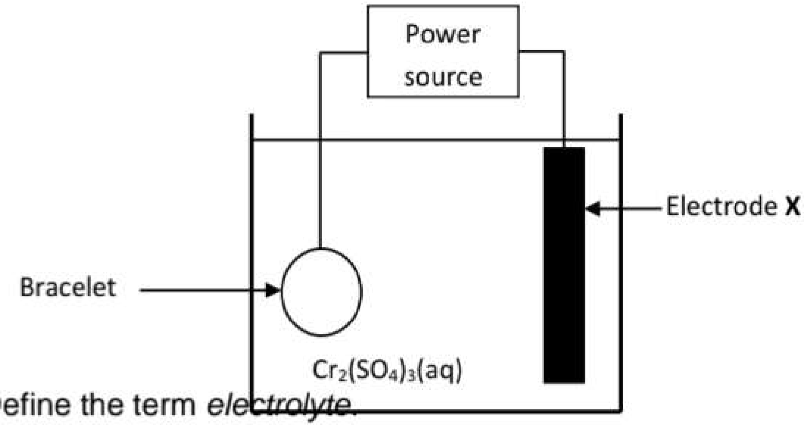
3.2.1 Half-reaction that takes place at electrode Y (2)

3.2.2 NAME or FORMULA of the gas released at electrode X (1)

3.3 Refer to the Table of Standard Reduction Potentials to explain why hydrogen gas, and not sodium, is formed at the cathode of this cell. (2)

QUESTION 4

A technician is plating a bracelet with chromium in an electrolytic cell containing $Cr_2(SO_4)_3(aq)$. The simplified diagram of the electrolytic cell is shown below.



41 Define the term *electrolyte*. (2)

4.2 Which electrode, the BRACELET or **X**, is the cathode? (1)

4.3 Write down the:

4.3.1 Metal of which electrode **X** is made (1)

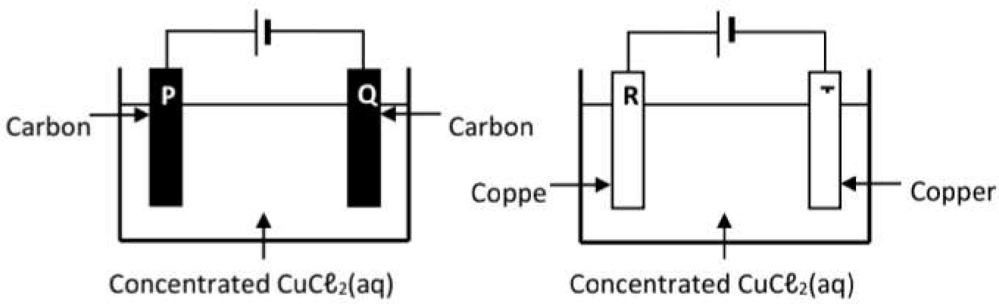
4.3.2 Reduction half-reaction (2)

4.4 During the process, the bracelet is plated with 0,86 g chromium. Calculate the number of electrons transferred during the process. (6)
[12]

QUESTION 5

The simplified diagrams below represent two electrochemical cells, **A** and **B**. A concentrated copper(II) chloride solution is used as electrolyte in both cells.

ELECTROCHEMICAL CELL A **ELECTROCHEMICAL CELL B**



5.1 Are **A** and **B** ELECTROLYTIC or GALVANIC cells? (1)

5.2 Which of the electrodes (**P**, **Q**, **R** or **T**) will show a mass increase? Write down a half-reaction to motivate the answer. (4)

5.3 Write down the NAME or FORMULA of the product formed at:

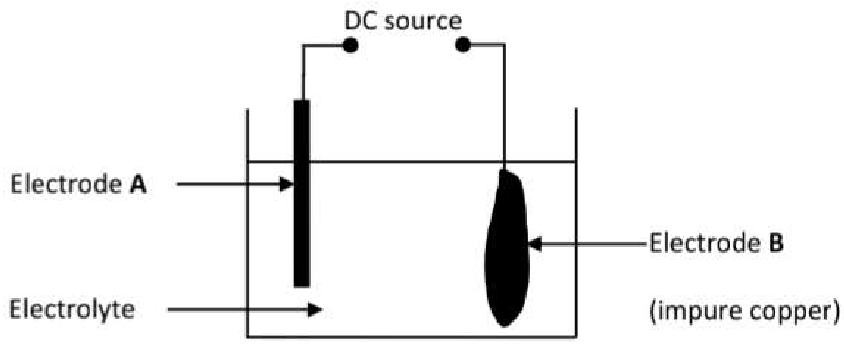
5.3.1 Electrode **P** (1)

5.3.2 Electrode **R** (1)

5.4 Fully explain the answer to QUESTION 5.3.2 by referring to the relative strengths of the reducing agents involved. (3)
[10]

QUESTION 6

The simplified diagram below represents an electrochemical cell used for the purification of copper.



6.1 Define the term *electrolysis*. (2)

6.2 Give a reason why a direct-current (DC) source is used in this experiment. (1)

6.3 Write down the half-reaction which takes place at electrode **A**. (2)

6.4 Due to small amounts of zinc impurities in the impure copper, the electrolyte becomes contaminated with Zn^{2+} ions.

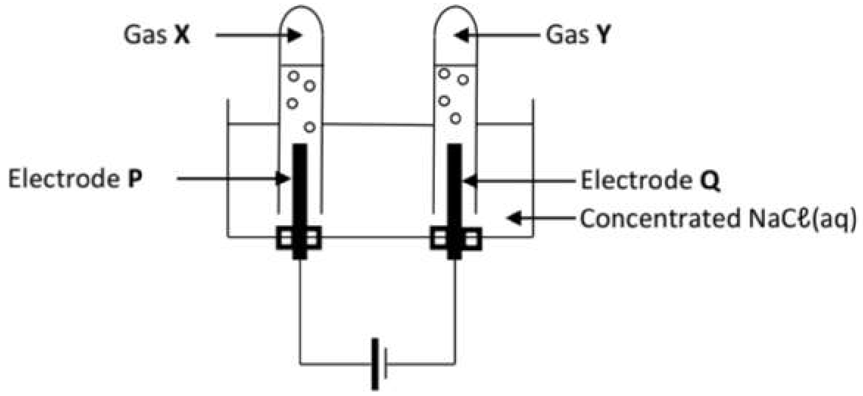
Refer to the attached Table of Standard Reduction Potentials to explain why the Zn^{2+} ions will not influence the purity of the copper obtained during this process. (3)

6.5 After the purification of the impure copper was completed, it was found that $2,85 \times 10^{-2}$ moles of copper were formed.

The initial mass of electrode **B** was 2,0 g. Calculate the percentage of copper that was initially present in electrode **B**. (4)

QUESTION 7

In the electrochemical cell below, carbon electrodes are used during the electrolysis of a concentrated sodium chloride solution.



The balanced equation for the net (overall) cell reaction is:

$$2H_2O(\ell) + 2C\ell^-(aq) \rightarrow C\ell_2(g) + H_2(g) + 2OH^-(aq)$$

7.1 Is the reaction EXOTHERMIC or ENDOTHERMIC? (1)

7.2 Is electrode **P** the ANODE or the CATHODE? Give a reason for the answer. (2)

7.3 Write down the:

7.3.1 NAME or FORMULA of gas **X** (1)

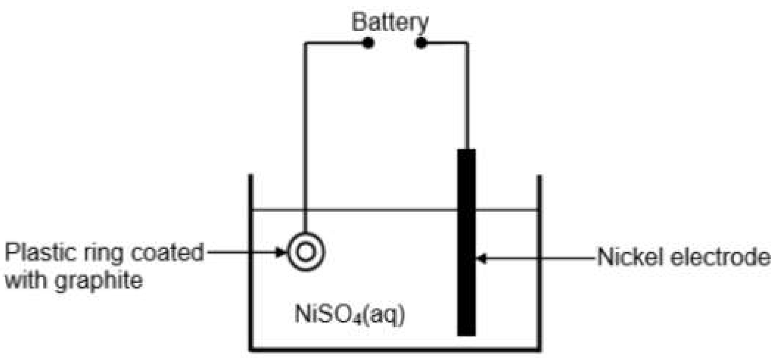
7.3.2 NAME or FORMULA of gas **Y** (1)

7.3.3 Reduction half-reaction (2)

7.4 Is the solution in the cell ACIDIC or ALKALINE (BASIC) after completion of the reaction? Give a reason for the answer. (2)

QUESTION 8

The diagram below shows a simplified electrolytic cell that can be used to electroplate a plastic ring with nickel. Prior to electroplating the ring is covered with a graphite layer.



8.1 Define the term electrolyte. (2)

8.2 Give ONE reason why the plastic ring must be coated with graphite prior to electroplating. (1)

8.3 Write down the:

8.3.1 Half-reaction that occurs at the plastic ring (2)

8.3.2 NAME or FORMULA of the reducing agent in the cell. Give a reason for the answer. (2)

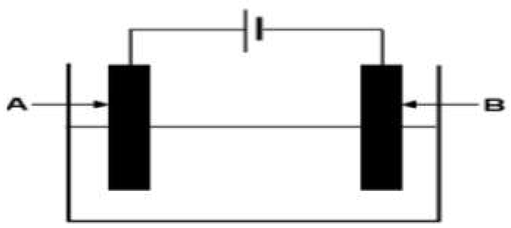
8.4 Which electrode, the RING or NICKEL, is the cathode? Give a reason for the answer.

The nickel electrode is now replaced with a carbon rod.

8.5 How will the concentration of the electrolyte change during electroplating? Write down only INCREASES, DECREASES or NO CHANGE. Give a reason for the answer. (2)

QUESTION 9

The diagram below shows an electrochemical cell used to purify copper. A solution that conducts electricity is used in the cell.



9.1 Write down:

9.1.1 ONE word for the underlined phrase above the diagram (1)

9.1.2 The type of electrochemical cell illustrated above (1)

9.2 In which direction (from A to B or from B to A) will electrons flow in the external circuit? (1)

9.3 Which electrode (A or B) is the:

9.3.1 Cathode (1)

9.3.2 Impure copper (1)

9.4 How will the mass of electrode A change as the reaction proceeds? Choose from INCREASES, DECREASES or REMAINS THE SAME.

Give a reason for the answer. (2)