A Guide to Electrochemical Reactions

Teaching Approach

This chapter is allocated 8 hours in CAPS and covers electrochemical reactions, the galvanic cell, the electrolytic cell, and applications of electrochemistry. The first step should be to revise redox reactions, oxidation numbers, and oxidation and reduction before you continue with the chapter. Note that the sections are not ordered in this way in the CAPS document, but learners are more likely to struggle if you do not revise these concepts first. A good understanding of these topics is essential.

The chapter ends with applications of electrochemistry in industry, this is a good opportunity to make sure the learners understand the principles, but it is not necessary for them to learn the reactions by heart or study the industry itself. However they need to be able to identify the half-reactions in these cells.

The following topics are covered in this chapter.

• Revision of oxidation, reduction and redox reactions
  It is important to revise the concepts of oxidation number, oxidation and reduction reactions, as well as how to write and balance redox reactions and half-reactions. An understanding of these concepts is very important in every aspect of this chapter.

• Galvanic and electrolytic cells
  This section introduces the concept of an electrochemical reaction; it may be worth a brief revision of the zinc-copper experiment performed in Grade 11. Two types of electrochemical cell are covered: galvanic cells and electrolytic cells. Galvanic cells convert chemical potential energy to electrical potential energy, while electrolytic cells convert electrical potential energy to chemical potential energy. Learners should be able to draw and label a simple diagram of both of these types of cells and write the standard cell notation for the cell.

• Processes in electrochemical cells
  This section is a more in-depth look at the oxidation and reduction reactions that take place at the anode and cathode in electrochemical cells. Learners should be able to predict whether an electrode is the anode or cathode and that this depends on the half-reaction that take place there. They must also know which half-reaction will take place at the anode and which will take place at the cathode.

• Effects of current and potential on rate and equilibrium
  Learners should understand that when a cell is flat the reactions inside are in equilibrium and the concentration of the reactants and products are constant. There is no potential difference and so no current can flow. The more current there is, the faster the electrons will flow and the higher the rate of the reaction must be. None of the cells they deal with in this chapter are in equilibrium, because then there would be nothing to measure.
• Standard electrode potentials
A standard hydrogen electrode is used to help determine the EMF of cells. The learners should understand that the electrode potentials in the table of standard electrode potentials are relative electrode potentials, and allow us to compare different metals. They should realise that the table provided is for the reduction half-reactions and that the oxidation half-reaction table would have the reactions reversed and the $E^0$ value with the opposite sign. Take note at this point in this series Table 4 B was used in all situations.

They need to be able to use the table to:
- determine which metal is more likely to be reduced or oxidised when two metal are compared and determine whether a reaction is spontaneous or not.

• Applications of electrochemistry
The applications covered in this section are electroplating, the chloralkali industry, and the extraction of aluminium. The learners should understand the processes involved, but do not need to memorise the exact reactions, and need not study the industries themselves.

There are several experiments in this chapter. The learners will use dangerous chemicals and should be properly instructed on the correct use of safety equipment, such as safety goggles, gloves and protective clothing. They should also be reminded not to sniff any chemicals as the fumes can be dangerous as well.
Video Summaries

Some videos have a ‘PAUSE’ moment, at which point the teacher or learner can choose to pause the video and try to answer the question posed or calculate the answer to the problem under discussion. Once the video starts again, the answer to the question or the right answer to the calculation is given.

Mindset suggests a number of ways to use the video lessons. These include:

- Watch or show a lesson as an introduction to a lesson
- Watch or show a lesson after a lesson, as a summary or as a way of adding in some interesting real-life applications or practical aspects
- Design a worksheet or set of questions about one video lesson. Then ask learners to watch a video related to the lesson and to complete the worksheet or questions, either in groups or individually
- Worksheets and questions based on video lessons can be used as short assessments or exercises
- Ask learners to watch a particular video lesson for homework (in the school library or on the website, depending on how the material is available) as preparation for the next day’s lesson; if desired, learners can be given specific questions to answer in preparation for the next day’s lesson

1. Exploring Electrochemistry?
   The link between electricity and chemical reactions is discovered and discussed in this lesson.

2. Types of Electrochemical Cells
   The two types of electrochemical cells are shown in this lesson. Experiments clearly illustrate the difference between these two cells.

3. Redox and Half Reactions
   Before the working of the cells can be explained we first revise redox reactions, and learn how to use the table of standard reduction potentials.

4. How the Galvanic Cell Works
   The cell that converts chemical energy spontaneously to electrical energy. In this lesson we look at all the different components in the cell and how it works.

5. The Standard Electrode Potential
   We give the standard condition under which standard electrode potential are determined and then describe the standard hydrogen electrode and its role as reference electrode

6. Calculation of Cell Potential
   Calculations of cell potential are shown and the answer is interpreted.

7. Equilibrium, Current and Emf
   An explanation of the relationship between current and the rate of the reaction is given and we find out what a “flat” cell is.
8. Electrolytic Cells
In this type of cell electrical energy is converted to chemical energy through a non-spontaneous reaction. We look at the microscopic reaction in these type of cells.

9. The Chloralkali Industry
The production of chlorine, hydrogen and sodium hydroxide is shown in an experiment and then we discuss the chemical reactions of the chloroalkali-industry in the membrane cell.

10. Electro Refining and Extraction of Metals
Two simple example of electroplating, namely the refining of copper and silver plating, is shown and discussed and then the recovery of aluminium metal from bauxite is explained.
Resource Material

Resource materials are a list of links available to teachers and learners to enhance their experience of the subject matter. They are not necessarily CAPS aligned and need to be used with discretion.

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Task

Question 1
Choose the correct answer
1.1 Electrochemistry involve the transfer of (electrons/neutrons/protons)
1.2 Electrochemical reactions are (acid base/redox) reactions
1.3 What do we call a device that makes use of an electrochemical reaction? (closed circuit/electrochemical cell)

Question 2
Choose the correct answer
2.1 Reduction is the (gain/loss) of electrons.
2.2 Reduction occurs at the (anode/cathode)
2.3 Oxidation (gain/loss) of electrons
2.4 Oxidation occurs at the (anode/cathode)
2.5 Magnesium metal is written as [Mg(s)/Mg^{2+}(aq)]
2.6 Magnesium ions dissolved in water are written as [Mg(s)/Mg^{2+}(aq)]
2.7 Magnesium metal (Mg) has a (silver/colourless) colour
2.8 Magnesium (Mg^{2+}) has a (silver/colourless) colour

Question 3
Answer the following:
3.1 What is a galvanic cell?
3.2 Mention two alternative terms for a galvanic cell.
3.3 What energy conversion takes place in a galvanic cell?
3.4 Describe an electrolytic cell
3.5 Is an electrolytic cell spontaneous or non-spontaneous?

Question 4
In each of the following equations, say which atoms in the reactants are oxidised and which are reduced.
4.1 CuO(s) + H_{2}(g) → Cu(s) + H_{2}O(g)
4.2 2NO(g) + 2CO(g) → N_{2}(g) + 2CO_{2}(g)
4.3 Mg(s) + FeSO_{4}(aq) → MgSO_{4}(aq) + Fe(s)
4.4 Zn(s) + 2AgNO_{3}(aq) → 2Ag(s) + Zn(NO_{3})_{2}(aq)

Question 5
Write down the oxidation half-reaction, reduction half reaction and balanced net redox reaction of each of the following:
5.1 Copper reacts with a silver nitrate solution
5.2 Magnesium reacts with a nickel nitrate solution

Question 6
What is a spectator ion?
Question 7

The cell notation of a standard galvanic (voltaic) cell containing an unknown metal, Electrode X, is shown below:

\[ \text{X(s)} \mid \text{X}^{3+} (1 \text{ mol} \cdot \text{dm}^{-3}) \mid \text{II} \mid \text{Pb}^{2+} (1 \text{ mol} \cdot \text{dm}^{-3}) \mid \text{Pb (s)} \]

7.1 Name the component of the cell represented by the double vertical lines (II) in the above cell notation.

7.2 State the TWO standard conditions that are applicable to the cell.

7.3 Identify the oxidising agent in the above cell.

7.4 The initial reading on a voltmeter connected across the electrodes of the above cell is 1.53 V. Identify metal X by calculating the standard reduction potential of the unknown metal X.

7.5 Write down the balanced equation for the net (overall) reaction taking place in this cell. Omit the spectator ions.

7.6 How will the initial voltmeter reading be affected if the concentration of the electrolyte in the X(s) I X^{3+} (aq) half – cell is increased? Write down only INCREASES, DECREASES or REMAINS THE SAME.

7.7 Write down the value of the reading on the voltmeter when the cell reaction has reached equilibrium.

Question 8

The diagram below shows an electrochemical cell set up under standard conditions using aluminum (Al) and nickel (Ni) electrodes. AlCl₃(aq) and NiCl₂(aq) are used as the electrolytes, and a solution of sodium nitrate (NaNO₃(aq)) is used in the salt bridge.

8.1 The diagram indicates that electrons flow from metal X to metal Y. Identify metal X and electrolyte B.

8.2

8.2.1 What is the concentration of electrolyte B?

8.2.2 Write down the FORMULA of the substance that moves towards metal Y in the salt bridge.

8.3 Write down the half reaction that occurs at the cathode of this cell.

8.4 Calculate the reading on the voltmeter at a temperature of 289 K.

8.5 State what happens to the concentration of metal ions in the solution containing electrolyte A as time goes by? (Just write down INCREASES, DECREASES, and REMAINS THE SAME.)
8.6 What effect does the change of concentration mentioned in 8.5 have on the voltmeter reading. Briefly explain your answer.

**Question 9**
The diagram below represents a cell that can be used to electroplate a tin medal with a thin layer of silver to improve its appearance.

![Diagram of electroplating cell]

9.1 Which one of P or the Medal is the anode in this cell?
9.2 Write down the following:
   9.2.1 SYMBOL of element of which electrode P is composed
   9.2.2 NAME of the electrolyte that has to be used to achieve the desired results
9.3 Switch S is now closed. Write down the visible changes that will occur at the following:
   9.3.1 Electrode P
   9.3.2 The medal
9.4 Write down the equation for the half – reaction to the answer to 9.3.2.
9.5 How will the concentration of the electrolyte change during the electroplating process? Write down only INCREASES, DECREASES or REMAINS THE SAME.
9.6 You want to coat the medal with copper instead of silver. State TWO changes that you will make to the above cell to obtain a medal coated with copper.

**Question 10**

A hydrogen half-cell is connected to an unknown half cell
The voltmeter reading in the external circuit is: 0.80 V.
10.1 Identify the unknown cell showing all your calculations.
10.2 Write down the cell notation of the cell.
Question 11
Aluminium is one of the most abundant metals on earth, yet it is expensive – largely because of the amount of electricity needed to extract it. Aluminium ore is called bauxite. The bauxite is purified to yield a white powder, aluminium oxide, from which aluminium can be extracted. The diagram below shows an electrolytic cell used for the extraction of aluminium at temperatures as high as 1 000 °C.

11.1 State the energy conversion that takes place in this electrolytic cell.
11.2 Is aluminium formed at the positive or negative electrode? Write down POSITIVE or NEGATIVE only.
11.3 Use the Table of Standard Reduction Potentials (Table 4B) to write down the half-reaction for the formation of aluminium.
11.4 Explain why carbon dioxide gas is formed at one of the electrodes.
11.5 Why should the carbon electrodes be replaced regularly?
11.6 Write down TWO negative effects that the extraction of aluminium can have on the environment.

QUESTION 12
The diagram below shows a type of membrane cell used in the chloroalkali industry.
12.1 Name the gases A and B.
12.2 Why is the membrane called a cationic membrane?
12.3 Write down the half-reaction that takes place at electrode N.
12.4 Apart from its use in household products, name ONE industrial use of chlorine.
12.5 Explain why this electrolytic process cannot be done in one large container without a membrane.
Task Answers

Question 1
1.1 Electrochemistry involve the transfer of (electrons/neutrons/protons)
1.2 Electrochemical reactions are (acid base/redox) reactions
1.3 What do we call a device that makes use of an electrochemical reaction? (closed circuit/electrochemical cell)

Question 2
2.1 Reduction is the (gain/loss) of electrons.
2.2 Reduction occurs at the (anode/cathode)
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2.4 Oxidation occurs at the (anode/cathode)
2.5 Magnesium metal is written as [Mg(s)/Mg$^{2+}$(aq)]
2.6 Magnesium ions dissolved in water are written as [Mg/Mg$^{2+}$(aq)]
2.7 Magnesium metal (Mg) has a (silver/colourless) colour
2.8 Magnesium (Mg$^{2+}$) has a (silver/colourless) colour

Question 3
3.1 A galvanic cell is an electrochemical cell that converts chemical energy to electrical energy through a spontaneous reaction.
3.2 Voltaic cell or Daniell cell
3.3 Chemical energy to electrical energy
3.4 An electrolytic cell is a cell that converts electrical energy to chemical energy
3.5 Non spontaneous

Question 4
The oxidation numbers are given above the atom.
4.1 \[\begin{array}{cccc}
+2 & -2 & 0 & 0 \\
CuO(s) + H_2(g) & \rightarrow & Cu(s) & + H_2O(g) \\
\end{array}\]
\[\begin{array}{c}
\text{Loss} \\
\text{Gain} \\
\end{array}\]
H$_2$ is oxidised
Cu in CuO is reduced to Cu

4.2 \[\begin{array}{cccc}
+2 & -2 & +2 & -2 \\
2NO(g) + 2CO(g) & \rightarrow & N_2(g) & + 2CO_2(g) \\
\end{array}\]
\[\begin{array}{c}
\text{Loss} \\
\text{Gain} \\
\end{array}\]
C in CO is oxidised
N in NO is reduced

4.3 \[\begin{array}{cccc}
0 & +2 & +2 & 0 \\
Mg(s) + FeSO_4(aq) & \rightarrow & MgSO_4(aq) & + Fe(s) \\
\end{array}\]
\[\begin{array}{c}
\text{Gain} \\
\text{Loss} \\
\end{array}\]
Mg(s) is oxidised
Fe in FeSO₄ is reduced

\[
\begin{align*}
\text{Gain} & \quad \text{Loss} \\
\text{Zn(s) + 2AgNO}_3(aq) & \rightarrow 2\text{Ag(s)} + \text{Zn(NO}_3)_2(aq) \\
\end{align*}
\]

Zn(s) is oxidised
Ag in AgNO₃ is reduced to Ag(s)

**Question 5**

5.1 Copper reacts with a silver nitrate solution
\[
\text{Cu(s)} + \text{AgNO}_3(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + \text{Ag(s)}
\]

Oxidation half reaction: \( \text{Cu(s)} \rightarrow \text{Cu}^{2+}(aq) + 2e^- \)

Reduction half reaction: \( \text{Ag}^+(aq) + e^- \rightarrow \text{Ag(s)} \) (x2)

Net reaction: \( 2\text{Ag}^+(aq) + \text{Cu(s)} \rightarrow \text{Cu}^{2+}(aq) + 2\text{Ag(s)} \)

5.2 Magnesium reacts with a nickel nitrate solution
\[
\text{Mg(s)} + \text{Ni(NO}_3)_2(aq) \rightarrow \text{Mg(NO}_3)_2(aq) + \text{Ni(s)}
\]

Oxidation half reaction: \( \text{Mg(s)} \rightarrow \text{Mg}^{2+}(aq) + 2e^- \)

Reduction half reaction: \( \text{Ni}^{2+}(aq) + 2e^- \rightarrow \text{Ni(s)} \) (x2)

Net reaction: \( \text{Ni}^{2+}(aq) + \text{Mg(s)} \rightarrow \text{Mg}^{2+}(aq) + \text{Ni(s)} \)

**Question 6**

A spectator ion is a chemical substance in the reaction that does not take part in the redox reaction, and that is chemically unchanged.

**Question 7**

7.1 Salt bridge

7.2 Concentration of the electrolyte = 1 mol·dm⁻³

Temperature = 25°C/298 K

7.3 Pb²⁺/lead (II) ions (are reduced from Pb²⁺ to Pb⁰)

7.4 7.4 \( E^0_{\text{cell}} = E^0_{\text{cathode}} - E^0_{\text{anode}} \)

\[
1.53 = (-0.13) - E^0_{\text{anode}}
\]

\[
E^0_{\text{anode}} = -1.66 \text{ V}
\]

Thus unknown metal \( X \) is Al.

7.5 7.5 \( 2 \text{Al}^3+ + 3 \text{Pb}^{2+} \rightarrow 2 \text{Al}^{3+} + 3 \text{Pb} \)

7.6 Decreases (As the reaction proceeds, the metal \( X \) dissolves creating more ions in solution. As the surface area of \( X \) gets less, there is less area for the reaction to occur and thus the PD drops)

7.7 0 V

**Question 8**

8.1 Metal \( X: \text{Al}^3+ \)

Electrolyte B: NiCl₂

8.2
8.2.1 1 mol.dm⁻³
8.2.2 NO₃⁻
8.3 Ni²⁺ + 2e⁻ → Ni
8.4 \( E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} \)
    = 0,25 V - (-1,66 V)
    = 1,41 V
8.5 Increase
8.6 The voltmeter reading decreases.
As the [A³⁺] increases, the reverse reaction in the A│A³⁺ half cell: A ⇌ A³⁺ + 3e⁻ is favored.
This lowers the electrode potential of this half-cell, resulting in a lower cell potential.

**Question 9**
9.1 P OR silver
9.2
9.2.1 Ag
9.2.2 Silver nitrate or silver ethanoate
9.3
9.3.1 Silver /metal bar becomes eroded / pitted / smaller / thinner/ eaten away / mass decreases / reduces (Dissolved not accepted)
9.3.2 A (silver) layer forms on the medal. / Medal becomes coated with silver / shiny silver layer /
    Medal’s mass increases
9.4 Ag⁺ + e⁻ → Ag
9.5 Remains the same.
9.6 Replace the silver solution with a copper solution of a soluble copper salt.
    Replace the silver bar/electrode P/anode with a copper bar.

**Soluble Copper Salts:**
Copper (II) nitrate CuSO₄  (CH₃COO)₂Cu  CuCl₂  CuBr₂  CuI₂

**Question 10**
10.1 \( E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} \)
    0,80 = \( E^{\circ}_{\text{cathode}} - (0,00) \)
    \( E^{\circ}_{\text{cathode}} = 0,80 \) V
Search the Table of Standard Reduction Potentials for a value of 0,80 V : Ag.
10.2 Write down the cell notation of the cell:
    Pt(s) / H₂(g )/(101,3kPa)/H⁺(aq)(1 mol.dm⁻³) // Ag⁺(aq) (1 mol.dm⁻³) / Ag(s)

**Question 11**
11.1 Electrical energy to chemical energy
11.2 negative
11.3 Al³⁺ (aq) + 3e⁻ → Al(s)
11.4 Carbon will burn in / react with O₂ because of the high temperature to form CO₂
    C(s) + O₂(g) → CO₂(g)
11.5 Carbon burns away / used up/oxidised / loses e⁻ (and needs to be replenished)
11.6 Any two:
    Ecological Impact
    • Loss of landscape due to the size of the chemical plant needed
    • Disposal of red mud (iron(III) oxide formed during extraction of aluminium oxide from bauxite)
        into lagoons causing them to become unsightly
    Environmental Impact
    • Carbon dioxide from the burning of the anodes contributes to the (enhanced) greenhouse
effect (air pollution /global warming)
- Carbon monoxide is poisonous
- Fluorine (and fluorine compounds) lost from the cryolite during the electrolysis process is poisonous
- Alkali of red mud dams can drain into soil and contaminate groundwater
- Pollution caused by power generation (for electrolytic process) using coal-fired plants leads to acid rain / enhanced greenhouse effect
- Noise pollution

**Question 12**

12.1 A: Chlorine
   B: Hydrogen

12.2 Allows only the cations (positive ions) to pass through it.

12.3 \[ 2H_2O + 2e^- \rightarrow H_2 + 2OH^- \]

12.4 Any one
   - Manufacture of PVC, paper, drugs etc.
   - Disinfectant for water

12.5 In a single pot the chlorine will react with water to form chlorine water / OR the chlorine will react with the OH^- ions to form bleach. OR Products formed will be contaminated.
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