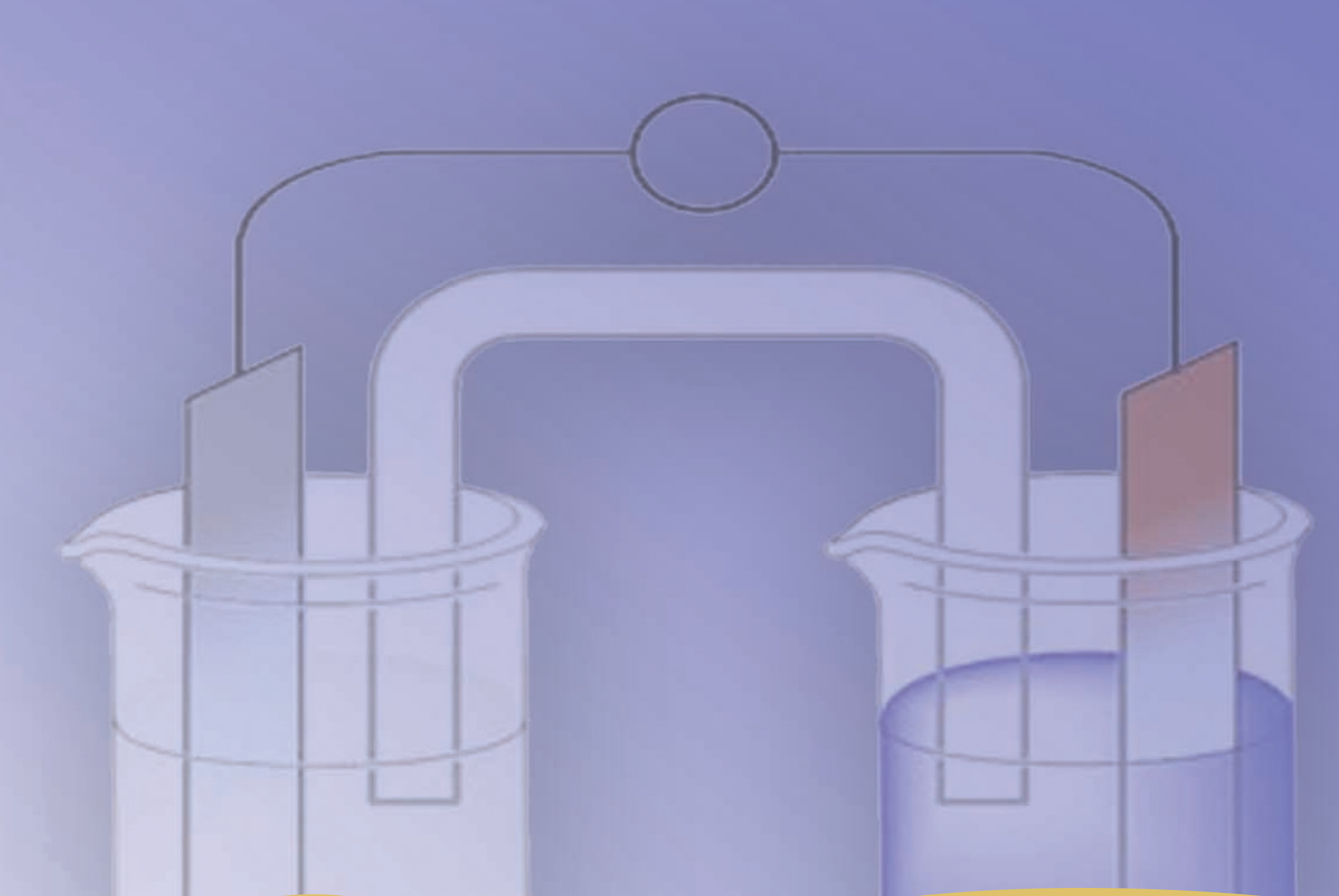


PHYSICAL SCIENCE

ELECTROCHEMISTRY : PAPER 2

GRADE 12

basic education
Department:
Basic Education
REPUBLIC OF SOUTH AFRICA

Read to Lead
A Reading Nation is a Leading Nation

2030
NDP

Foreword

In order to improve learning outcomes the Department of Basic Education conducted research to determine the specific areas that learners struggle with in Grade 12 examinations. The research included a trend analysis by subject experts of learner performance over a period of five years as well as learner examination scripts in order to diagnose deficiencies or misconceptions in particular content areas. In addition, expert teachers were interviewed to determine the best practicesto ensure mastery of thetopic by learners and improve outcomes in terms of quality and quantity.

The results of the research formed the foundation and guiding principles for the development of the booklets. In each identified subject, key content areas were identified for the development of material that will significantly improve learner's conceptual understanding whilst leading to improved performance in the subject.

The booklets are developed as part of a series of booklets, with each bookletfocussing onlyon one specific challenging topic. The selected content is explained in detail and include relevant concepts from Grades 10 - 12 to ensure conceptual understanding.

The main purpose of these booklets is to assist learners to master the content starting from a basic conceptual level of understanding to the more advanced level. The content in each booklets is presented in an easy to understand manner including the use of mind maps, summaries and exercises to support understanding and conceptual progression. These booklets should ideally be used as part of a focussed revision or enrichment program by learners after the topics have been taught in class. The booklets encourage learners to take ownership of their own learning and focus on developing and mastery critical content and skills such as reading and higher order thinking skills.

Teachers are also encouraged to infuse the content into existing lesson preparation to ensure in-depth curriculum coverage of a particular topic. Due to the nature of the booklets covering only one topic, teachers are encouraged to ensure learners access to the booklets in either print or digital form if a particular topic is taught.

Table of Contents

2. HOW TO USE THIS BOOKLET

This booklet serves as a guide to understanding Grade 12 Physical Sciences. It does not, however, replace your textbook. The authors have used their experience and focused their attention on the areas that learners seem to struggle with a lot. The book focuses mostly on the challenges that have been observed from learners' responses in the Grade 12 national examinations over the past few years.

This guide aims to help explain the concepts in your regular textbook further. It also offers more exercises that serve as building blocks to your own understanding of what is expected of you in these topics. The book draws on the basic knowledge acquired in the lower grades and demonstrates how this knowledge fits in with the new chapters encountered in Grade 12, starting from atomic theory. Figure 1 on section 4 of this book, shows how your study of chemistry is related to the atomic theory. It is therefore important to remember this theory, the properties of the parts of an atom and the role that they play throughout chemistry.

The chemistry in your textbook is presented in chapters and sections, but remember to integrate these chapters when you answer questions, because everything is interrelated. Science, as an expression of nature, draws from nature; therefore, do not think of concepts in isolation from what you already know about chemistry or physics! Use the information learnt in other sections to solve the problems at hand.

All the questions in this booklet have solutions. Some questions have been sourced from past papers, while the rest were sourced elsewhere. Study each question carefully and make sure that you understand the steps taken to solve the question. Try some of the questions without looking at the solutions. After completing an exercise, check your solutions against the answers. Move on to the rest of the questions and try to understand why you were wrong in instances when you were actually wrong.

Do not hesitate to ask your teacher if you struggle with any of the exercises.

3. STUDY AND EXAMINATION TIPS

As you prepare to write your examination, it is important to understand the rules, laws and concepts that govern certain aspects of your work. Ensure that you understand these rules/ laws/ definitions and concepts very well. Understand what they mean, **where they apply, when they apply,** and also **when and where they do not apply.**

Pay attention to your health:

1. Eat wholesome food: grains, fruit and vegetables; and drink enough water.
2. Engage in light exercise like walking.
3. Get enough sleep and rest.

Develop a study timetable and try to be loyal to the time allocated to studying.

3.1 Subject-specific tips

Always:

1. Start with the questions that you know you are able to answer.
2. Read the question that you are working on carefully.
3. Understand what it says and what is required of you.
4. Write down the information that you have.
5. Write down the information that you do not have.
6. Use existing information to derive what you need in order to solve the question.
7. All questions have hints that point to the answer.
8. Check your work by going through these steps again.

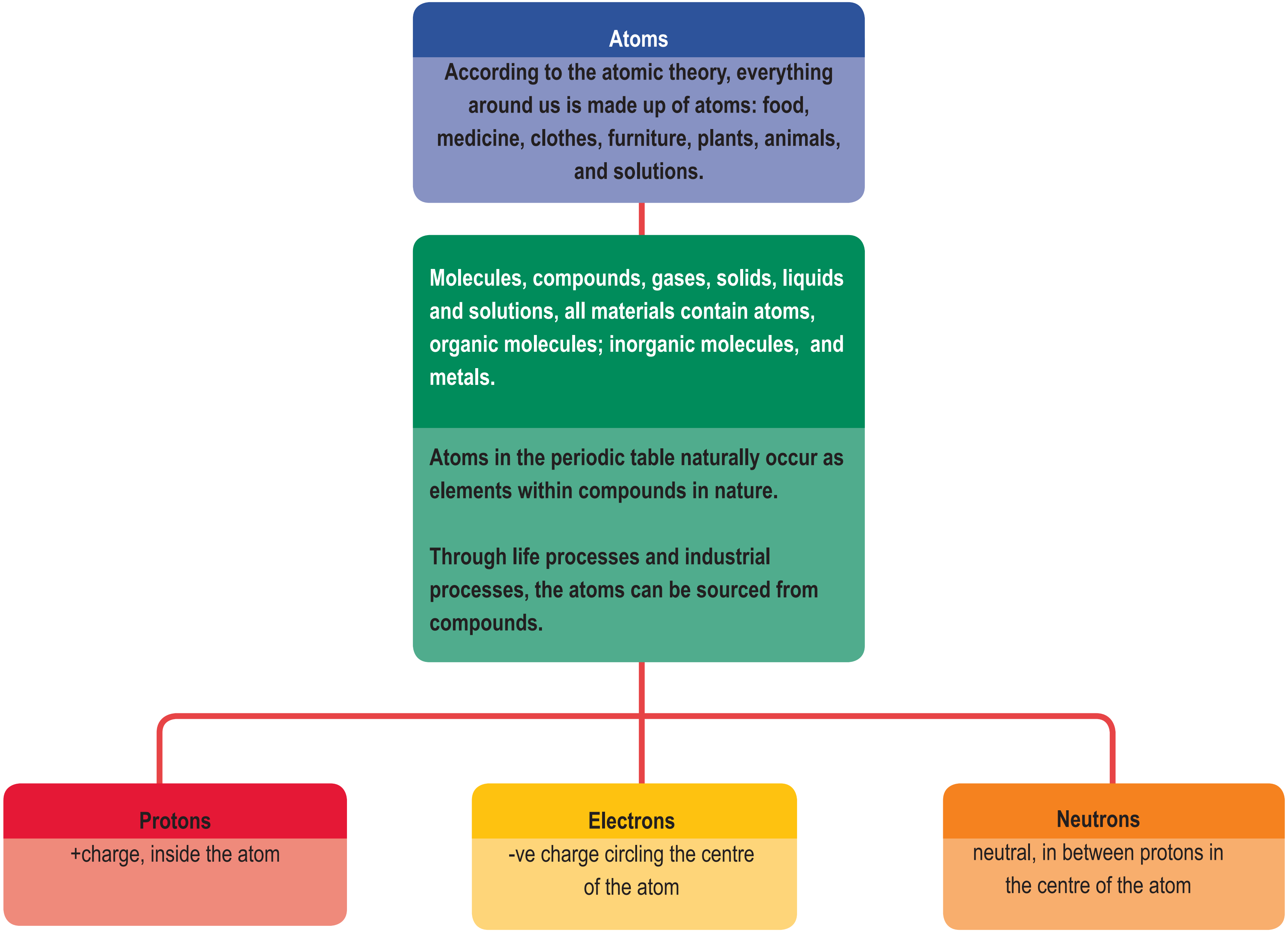
3.2 Topic-specific tips

About a third of the Grade 12 Physical Sciences Paper 2 examination comprises of Electrochemistry questions. This booklet will give you a few tips that will help you to achieve marks in organic chemistry. Electrochemistry is at the centre of synthesis in chemistry. It encompasses concepts like oxidation, reduction, electronegativity, potential difference and electricity. Electrochemistry is a special case of redox reactions. Redox reactions are responsible for almost all synthesis (formation of new products) in nature and in industry. Remember that when we add an electron or remove an electron from an atom or compound, we change its chemical and physical properties.

3.3 Overview of electrochemistry

Electrochemistry, as studied in this grade, does more than just show how electrons gained or lost electrons and form new products using redox reactions. It shows how electrons, if provided with a pathway can produce light. It also show how electricity can cause a chemical change by donating electrons to ions and form compounds or elements from ions in solution. An atom in a different oxidation state displays characteristics that are different from those of the atom before it changed its oxidation state. The difference could be in the state, colour, reactivity, etc.

4. THE ATOMIC THEORY: ITS PLACE IN CHEMISTRY



Atoms

According to the atomic theory, everything around us is made up of atoms: food, medicine, clothes, furniture, plants, animals, and solutions.

Molecules, compounds, gases, solids, liquids and solutions, all materials contain atoms, organic molecules; inorganic molecules, and metals.

Atoms in the periodic table naturally occur as elements within compounds in nature.

Through life processes and industrial processes, the atoms can be sourced from compounds.

Protons

+charge, inside the atom

Electrons

-ve charge circling the centre of the atom

Neutrons

neutral, in between protons in the centre of the atom

This sections deals mainly with electrons, i.e.:

1. How they can attach themselves to other atoms, leaving their atoms, as well as the conditions that allow for this transfer of electrons.
2. How they move through a conductor.
3. How this movement of electrons can be directed to be useful to humanity.
4. How to predict which atoms will release their electrons and which will accept the electrons.
5. The industrial benefits of electrons.

4.1 Electrochemical reactions

(CAPS, p. 134–137)

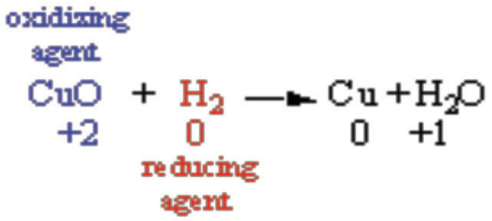
	 This is the name of the process whereby electron(s) are removed from an atom. Reduction happens when an atom gains electron(s). Chemical and physical properties of atoms are unique to atoms. An atom has the same number of protons and electrons. When oxidation or reduction takes place and one or more electrons are removed or added to the atom in question, the atom or compound acquires a charge as well as different physical and chemical properties to those of the original atom before the removal or addition of electrons. Most noticeable with some atoms is the colour change. Note that charged ions will usually exist in solution, e.g. H^+ and OH^- (water).
Oxidation	Oxidation brings about charge as well a physical and chemical change. Changes from Fe^{2+} Very pale green to Fe^{3+} Very pale violet/ brown
Reduction	$\mathbf{Cl_2 + H_2 \rightarrow HCl\ (H^+\ Cl^-)}$ **0 0 1 - 1** **Oxidation states:** In the above reaction, hydrogen gas is added to chlorine gas to form hydrogen chloride. Thus, chlorine gas is being reduced (from an oxidation state of 0 to -1) to form hydrochloric acid.

Both oxidation and reduction take place. An electron that is given away is readily accepted in the same reaction. When an electron leaves an atom, it is accepted by another atom in the same reaction. Hence, the term redox.

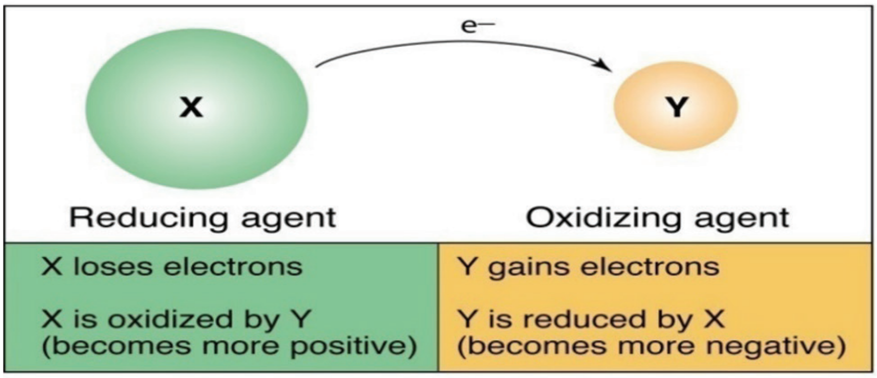
TERMINOLOGY

Pay extra attention to the following terminology. Most learners find it very confusing.

Oxidising agent and **reducing agent**



The oxidation state for copper in CuO in this reaction is 2^+ while the oxidation state of oxygen is 2^-. The product of this reaction is Cu^0 and H_2O. The oxidation state of Cu is now 0 and not -2 anymore. CuO is an oxidising agent, while H_2 is a reducing agent.



Reducing agent	Oxidizing agent
X loses electrons X is oxidized by Y (becomes more positive)	Y gains electrons Y is reduced by X (becomes more negative)

A substance that is oxidised, loses electrons. It is called a reducing agent because the electrons it loses reduce another atom or compound.

X is a reducing agent, as it causes a reducing reaction in Y; Y is an oxidising reagent as it causes an oxidising reaction in X.

WORKED EXAMPLES

Activity

1. Calculate the oxidation number for carbon in each of the following compounds:
   (a) diamond, C; (b) dry ice, CO_2; (c) marble, $CaCO_3$;(d) baking soda, $NaHCO_3$

Solution

(Recall the rules for assigning oxidation states)

Solution: Calculate oxidation numbers for carbon. Begin by recalling that uncombined elements, as well as compounds, are electrically neutral. Thus, free elements and compounds have no charge (a) In diamonds, the oxidation number of carbon is zero. (b) With dry ice, oxygen is assigned an oxidation number of –2. Determine the oxidation number of carbon in CO_2 (c) With marble, assign calcium ion an oxidation number of +2, and oxygen a value of –2. Determine the value of carbon in $CaCO_3$ (d) With baking soda, assign sodium ion an oxidation number of +1, hydrogen a value of +1, and oxygen a value of –2. Determine the oxidation number of carbon in $NaHCO_3$

Begin by recalling that the charge on an ion corresponds to the sum of the oxidation numbers. For example:

(a) In S^{2-}, the oxidation number of sulphur is 2^-.

(b) In $SO_3{}^{2-}$, the polyatomic anion has a charge of 2^-. Assign oxygen an oxidation number of –2 and write the equation.

(c) In $SO_4{}^{2-}$, the polyatomic anion has a charge of 2^-. Assign oxygen an oxidation number of –2 and write the equation

(d) In $S_2O_3{}^{2-}$, the polyatomic anion has a charge of 2^-. Assign oxygen an oxidation number of –2 and write the equation.

Corrosion is an example of a redox reaction. Corrosion is a process through which metals in manufactured states return to their natural **oxidation** states. This process is a **reduction-oxidation reaction** in which the metal is being **oxidised** by its surroundings, often by the oxygen in the air.

4.2 Electrochemistry - a special case of redox reactions

In electrochemistry, the movement of electrons between the metals and ions from species is directed by using a conducting wire. These electrons then produce energy in the form of light, heat or power to drive electronic apparatus and machinery. There are many factors that determine if electrons will readily leave their atom or readily accept electrons. The standard reduction potential table reflects that potential.

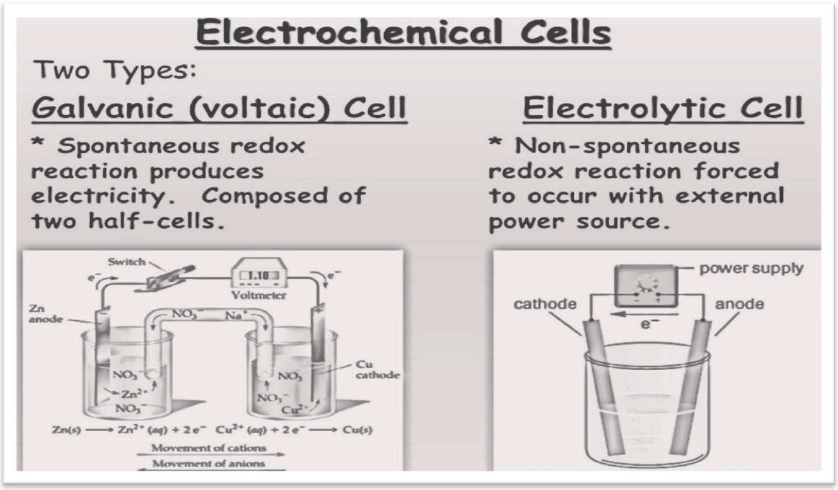
Some reactions are spontaneous while others are non-spontaneous. In a non-spontaneous reaction, electrons in the form of electricity are used to break compounds. These reactions happen inside of what is called an electrochemical cell. In a spontaneous reaction, the reaction happens without needing electricity.

Electrochemical cells either use chemical reactions to generate electricity (spontaneous) or, conversely, use electricity to provide energy for useful chemical reactions (non-spontaneous).

4.2.1 CONSTRUCTING AN ELECTROCHEMICAL CELL

There are only two types of electrochemical cells

The two main types of electrochemical cells are: **electrolytic cells** and **galvanic cells** (also called **voltaic cells**). There are, however, many examples that are either galvanic (producing electricity) or electrolytic, using electricity.



Electrochemical Cells

Two Types:

Galvanic (voltaic) Cell

* Spontaneous redox reaction produces electricity. Composed of two half-cells.

Electrolytic Cell

* Non-spontaneous redox reaction forced to occur with external power source.

While the galvanic cell **produces** electricity, the electrolytic cell **needs** electricity.

Differences between a galvanic cell and an electrolytic cell

Galvanic cell	Electrolytic cell
1. In a galvanic cell, electrical energy is produced.	1. In electrolytic cell, electrical energy is consumed.
2. In a galvanic cell, the reaction taking place is spontaneous.	2. In an electrolytic cell, the reaction that takes place is non-spontaneous - hence the need for electricity.
3. The two half cells are set up in different containers and are connected through a salt bridge or porous partition.	3. Both electrodes are placed in the solution or molten electrolyte in the same container.
4. In a galvanic cell, electrons migrate to an anode and migrate from a cathode.	4. In an electrolytic cell, the electrons migrate from the anode and migrate to the cathode.
5. The electrons move from the anode to cathode in an external circuit.	5. Electrons are supplied by the external source. They enter through the cathode and come out through the anode.

4.2.2.1 ELECTROLYTIC CELLS DRIVE CHEMICAL REACTIONS WHEN ELECTRICAL ENERGY IS APPLIED

The most familiar example of electrolysis is the decomposition (breakdown) of water into hydrogen and oxygen by means of an electric current. This gives rise to what is generally termed a hydrogen fuel cell.

Did you know?

The elements, as they appear on the periodic table, are not necessarily found as elements in nature. They are found in compounds. Nature and human beings have developed methods of isolating (decomposition/ breakdown) these elements from their compounds into individual elements, as represented in the periodic table.

Electrolysis is one of the methods used to break down compounds into the desired elements.

Electrolysis is the process that can be used to decompose compounds other than water. An example of electrolysis is the isolation of sodium chloride (table salt) into sodium (Na) and chlorine (Cl) from sodium chloride. In isolating Na and Cl, an electrochemical cell is set up by immersing two electrodes in molten NaCl. Chemical reactions occur at each electrode as electrons are pumped by the energy source from one electrode to the other, causing the decomposition of NaCl.

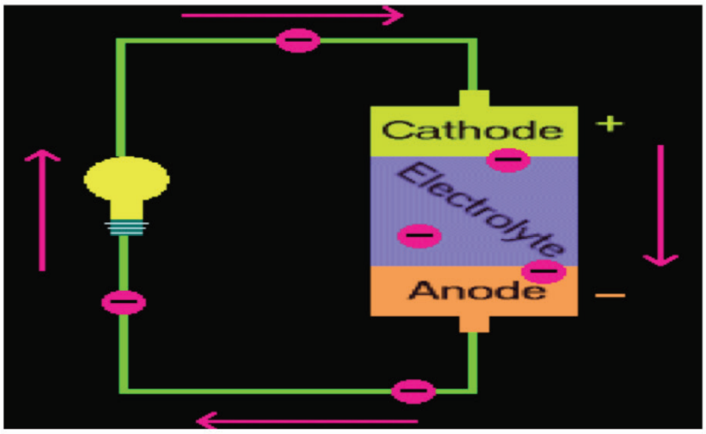
Magnesium and aluminium are elements produced commercially by electrolysis from their compounds.

4.2.2.2 GALVANIC CELLS USE CHEMICAL REACTIONS TO GENERATE ELECTRICAL ENERGY

Remember that all atoms are surrounded by electrons. These electrons, when directed, go into what seems like a flow (and generally called a flow). This flow gives us electricity. Some chemical reactions are therefore used to generate electricity, e.g. batteries (a chemical reaction that produces electrical light).

How does a battery work?

Batteries have three parts, an anode (-), a cathode (+), and the electrolyte. The cathode and anode (the positive and negative sides at either end of a traditional battery) are hooked up to an electrical circuit. The chemical reactions in the battery cause a build-up of electrons at the anode. This results in an electrical difference between the anode and the cathode. The electrons want to rearrange themselves to eliminate this difference, since electrons repel each other and try to move to a place with fewer electrons. In a battery, such as the one in the picture, the only place to go is to the cathode. But the electrolyte between the anode and the cathode keeps the electrons from going straight from the anode to the cathode in the battery. When the circuit is closed (using a wire that connects the cathode and the anode), the electrons will be able to get to the cathode. In the picture, the electrons move along the wire, lighting the light bulb along the way. This is one way of describing how electrical potential causes electrons to flow through the circuit.



An **electrode** is a conductor that is used to make contact with a non-metallic part of a circuit. This could be a gas or an aqueous solution.

An electrode is classified as either a **cathode** or an **anode** depending on whether current is flowing into or out of the electrode. A specific metal in its solution can either be a cathode or an anode, depending on what it is connected to. The table of reduction potentials shows which will be an anode and which will be a cathode. The one with high potential is the cathode, and the anode has low potential.

Anodes and cathodes

An electrode at which the oxidation reaction occurs is called the anode. Electrons from the **anode** migrate to the **positive** terminal of the battery, and electrons from the battery migrate to the cathode. An electrode at which the reduction reaction occurs is called the **cathode.** The cathode supplies electrons to positively **charged** cations which migrate to it from the electrolyte. The one with the **highest reduction potential** will be what you want to select as the reduction half reaction and therefore want to be your cathode. The one with the **lowest reduction potential** will be what you want to select as the oxidation half reaction and therefore want to be your anode. Note, for a good electrochemical cell, you want a cathode with a large positive reduction potential and an anode with a large (in magnitude) negative number reduction potential. Both reduction potentials can be positive or negative, as long as one is larger than the other in terms of their values.

The **cathode** supplies electrons to the positively **charged** cations, which flow to it from the electrolyte (even if the cell is galvanic, i.e. when the cathode is positive and therefore would be expected to repel the positively **charged** cations. This is due to electrode potential relative to the electrolyte solution being.

4.2.2.3 CELL NOTATION

Cell notation in chemistry is a shorthand way of expressing a certain reaction in an electrochemical cell. The cell anode and cathode (half cells) are separated by two bars (II) or slashes that represent a salt bridge, with the **anode on the left** and the **cathode on the right.** Individual solid, liquid or aqueous **phases** within each half cell are separated by a single bar (I). Concentrations of dissolved species in each phase are written in **parentheses** () and the state of each phase (usually s (solid), l (liquid), g (gas) or aq. (aqueous solution)) is included in a subscript after the species name. Note the oxidation states of aqueous solutions.

Worked activity

Describe, in shorthand notation, a galvanic cell for which the cell reaction is

$Cu(s) + 2Fe^{3}+(aq) \rightarrow Cu^{2}+(aq) + 2Fe^{2}+(aq)$

Solution

First divide the cell reaction into half-equations

- oxidation: $Cu(s) \rightarrow Cu^{2}+(aq) + 2e$
- reduction: $Fe^{3}+ + e– \rightarrow Fe^{2+}$

Then write the oxidation as the left-hand electrode, with the reduction on the right.

$Cu|Cu2+||Fe^{2+},Fe^{3+}|Pt$

Since both Fe^{2+} and Fe^{3+} are aqueous solutions, a Pt electrode is used. Fe^{0} is the only oxidation state that is not aqueous for this metal.

4.2.2.4 TABLE OF REDUCTION POTENTIALS

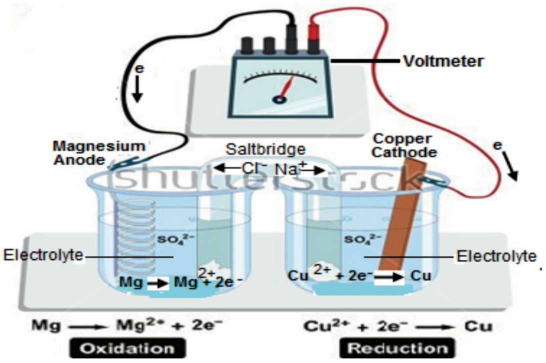
The standard reduction potentials are all based on the standard hydrogen electrode. By using the table of reduction potentials, we can predict if the cell is spontaneous or not. If it is **not spontaneous, it needs electricity** to react and it is an electrolytic cell.

Predicting spontaneous reactions

Refer to the table of reduction potentials to predict whether or not the reaction is spontaneous. Ni^{2+} (aq) is a weaker oxidising agent than Sn^{2+} (aq). Moreover, Sn(s) is a weaker reducing agent than Ni(s). Because the reactants are the weaker pair of oxidising and reducing agents, the reaction is non-spontaneous. Conversely, the reverse reaction is spontaneous because the products are the stronger oxidising and reducing agents.

4.3 Examples of cells

4.3.1 Galvanic/ voltaic cell; Mg and Cu



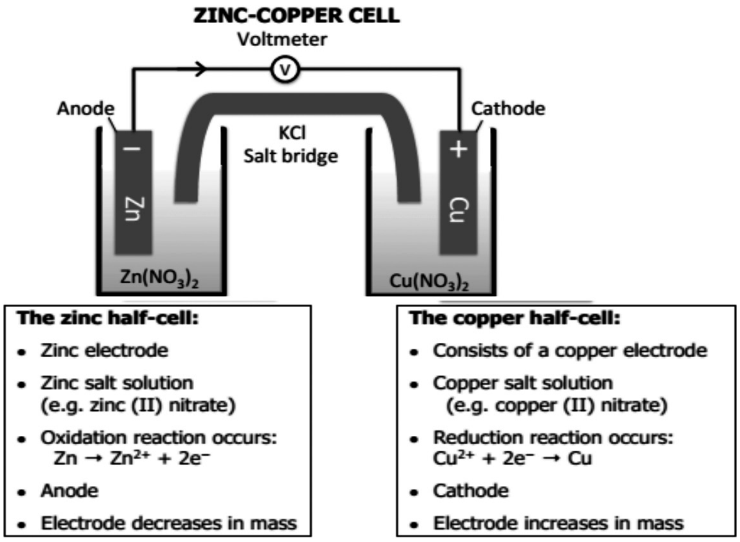
- A galvanic cell reaction is always a spontaneous, exothermic reaction, during which chemical energy is converted to electrical energy.
- A positive value of the standard emf is an indication that the reaction is spontaneous under standard conditions.
- Standard conditions refer to: temperature - 25^0C/298K
  concentration - 1mol.dm^{-3}
  pressure - 1,013 x 10^5 Pa (for gaseous reactants or products)
- Two half cells are placed in separate containers.
- The anode and cathode are connected using an external circuit, which allows for current (electrons) to flow from the anode to the cathode.
- In the internal circuit, ions flow to complete the circuit.
- **The standard hydrogen electrode (SHE)**: The hydrogen half-cell is allocated a reference potential of 0,00 V.

  All other half-cells will have standard potential, which is either higher or lower than this reference, i.e. their potential difference is measured against the SHE.
- When the circuit is complete, the current will begin to flow.

  The current and potential difference of the cell is related to the rate of the reaction and the extent to which the reaction in the cell has reached equilibrium.
- As the chemical reaction proceeds, the rate of the forward reaction will decrease; therefore, the rate of transfer of electrons will also decrease, which results in the E^{θ}_{cell} value decreasing.

  When the cell potential decreases, the current in the circuit will also decrease.
- The cell potential will continue to decrease gradually, until equilibrium is reached, at which point the cell potential will be zero and the battery is “flat”.

4.3.2 FOR THE ZINC-COPPER CELL:



The zinc half-cell:

- Zinc electrode
- Zinc salt solution (e.g. zinc (II) nitrate)
- Oxidation reaction occurs: $Zn \rightarrow Zn^{2+} + 2e^-$
- Anode
- Electrode decreases in mass

The copper half-cell:

- Consists of a copper electrode
- Copper salt solution (e.g. copper (II) nitrate)
- Reduction reaction occurs: $Cu^{2+} + 2e^- \rightarrow Cu$
- Cathode
- Electrode increases in mass

For the zinc-copper cell:

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

Reduction half-cell : $Cu^{2+}\ (aq) + 2e^- \rightarrow Cu\ (s)$

Nett cell : $Zn\ (s) + Cu^{2+}(aq) \rightarrow Zn^{2+}\ (aq) + Cu\ (s)$

Calculations:

$E^{\theta}_{cell} = E^{\theta}_{cathode} - E^{\theta}_{anode}$

$E^{\theta}_{cell} = 0{,}34 - (-0{,}76)$

$E^{\theta}_{cell} = +1{,}1\ V$

Cell notation is written as follows:

Anode **Salt bridge** **Cathode**

↓ ↓ ↓

$Zn(s)\ /Zn^{2+}(aq)\ (1\ mol{\cdot}dm^{-3})\ //\ Cu^{2+}(aq)\ (1\ mol{\cdot}dm^{-3})\ /\ Cu(s)$

Please note:

In the cell notation:

When the half reactions do not include conductors/ metals (i.e. when the half-cell involves gases/ ions), unreactive/ inert electrodes are used, e.g. carbon or platinum.

The anode (oxidation half-cell) is always written on the left and the cathode (reduction half-cell) is on the right.

/ represents electrode-electrolyte interphase. (What is interphase? Explain).

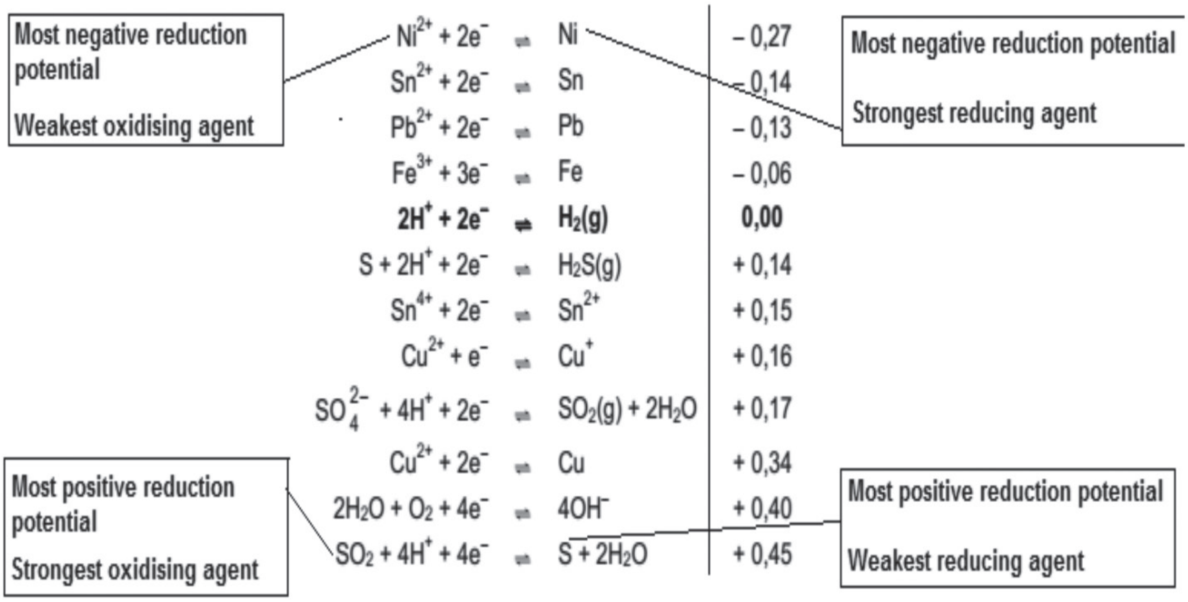
// represents a salt bridge. (Give examples of salts.)

Look at the following cell notations and insert commas in the correct place:

1. **Pt(s)/H$_2$(g)(1 atm)/H$^+$(aq)(1 mol·dm^{-3})//Br$_2$(g)(1 atm)/Br$^-$(aq)(1 mol·dm^{-3})/Pt(s)**
2. **Ca(s)/Ca^{2+}(aq)(1 mol·dm^{-3})//Fe^{3+}(aq)(1 mol·dm^{-3})/Fe^{2+}(aq)(1 mol·dm^{-3})/C(s)**

4.4 Tips on how to use a standard reduction potential table

Setion of the TABLE OF STANDARD REDUCTION POTENTIALS (4B)



$Ni^{2+} + 2e^-$	$\rightleftharpoons$	Ni	− 0,27
$Sn^{2+} + 2e^-$	$\rightleftharpoons$	Sn	− 0,14
$Pb^{2+} + 2e^-$	$\rightleftharpoons$	Pb	− 0,13
$Fe^{3+} + 3e^-$	$\rightleftharpoons$	Fe	− 0,06
$2H^+ + 2e^-$	**$\rightleftharpoons$**	**$H_2(g)$**	**0,00**
$S + 2H^+ + 2e^-$	$\rightleftharpoons$	$H_2S(g)$	+ 0,14
$Sn^{4+} + 2e^-$	$\rightleftharpoons$	Sn^{2+}	+ 0,15
$Cu^{2+} + e^-$	$\rightleftharpoons$	Cu^+	+ 0,16
$SO_4^{2-} + 4H^+ + 2e^-$	$\rightleftharpoons$	$SO_2(g) + 2H_2O$	+ 0,17
$Cu^{2+} + 2e^-$	$\rightleftharpoons$	Cu	+ 0,34
$2H_2O + O_2 + 4e^-$	$\rightleftharpoons$	$4OH^-$	+ 0,40
$SO_2 + 4H^+ + 4e^-$	$\rightleftharpoons$	$S + 2H_2O$	+ 0,45

Activity

Build two spontaneous cells and two non-spontaneous cells using the table above.

5. ACTIVITIES AND SOLUTIONS

Exercise 1

1.1 Consider the electrochemical cell represented by the cell notation below, where X is an unknown metal:

Pt(s) | Fe^{2}+(aq), Fe^{3}+(aq) || X^{+}(aq) | X(s)

The cell potential of this cell was found to be 0,03 V.

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

1.1.2 What does the single line (|) in the above cell notation represent? (1)

1.1.3 Write down the half-reaction that takes place at the anode in the above cell. (2)

1.1.4 Identify X with the aid of a calculation. (5)

1.2 A Pt(s) | Fe^{2}+(aq), Fe^{3+}(aq) half-cell is connected to a Cu(s) | Cu^{2}+(aq) half-cell.

Write down:

1.2.1 The chemical symbol for the electrode in the cathode half-cell. (1)

1.2.2 The NAME of the oxidising agent. (1)

1.2.3 The overall balanced cell reaction that takes place in this cell. (3)

Solutions **[14]**

1.1.1 Galvanic cell ✓ (1)

1.1.2 Indicates phase boundary/ interphase / phase separator. ✓ (1)

1.1.3 $Fe^{2+} \rightarrow Fe^{3+} + e^{-}$ ✓✓ (2)

1.1.4 $E^{\ominus}_{cell} = E^{\ominus}_{reduction} - E^{\ominus}_{oxidation}$ ✓

0,03 ✓ = $E^{\ominus}_{reduction}$ - (0,77) ✓

$E^{\ominus}_{reduction}$ = 0,80 V ✓ = $\boldsymbol{E^{\ominus}_{X/X^{+}}}$

∴ X = Silver/Ag✓ (5)

1.2.1 Pt ✓ (1)

1.2.2 Iron(III) (ions)/ ferric ions ✓ (1)

1.2.3 $2Fe^{3+} + Cu$ ✓ $\rightarrow 2Fe^{2+} + Cu^{2+}$✓ balancing ✓ (3)

[14]

Exercise 2

2.1 A group of learners use the redox reaction below to construct an electrochemical cell:

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

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

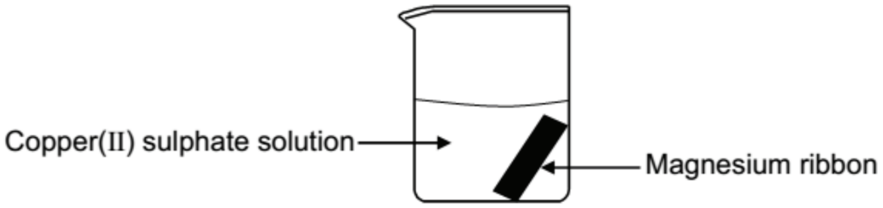
2.1.2 Name a substance that should be used as an electrode in the anode half-cell. (1)

2.1.3 Write down the **name** or **formula** of the reducing agent. (1)

2.1.3 Write down the cell notation of the cell. (3)

2.1.4 Calculate the initial emf of this cell under standard conditions. (4)

2.2 In a separate experiment, the learners place a magnesium ribbon in a beaker containing a blue solution of copper (II) sulphate. After a while, the solution becomes colourless.



2.2.1 Besides the colour change of the solution, state ONE observable change in the beaker with regard to the magnesium ribbon. (1)

2.2.2 Refer to the relative strengths of oxidising agents or reducing agents to explain why the solution in the beaker changes colour from blue to colourless. (3)

Solutions **[15]**

2.1.1 A substance that loses/ donates electrons. ✓✓ (2)

2.1.2 Platinum/Pt✓ (1)

2.1.3 Sn^{2+}✓(aq)/tin(II) ions✓ (1)

2.1.3 ✓$Pt \mid Sn^{2+}(aq)$, $Sn^{4+}(aq) \parallel$ ✓$Ag^{+}(aq) \mid Ag(s)$ ✓

OR

$Pt \mid Sn^{2+}(1\ mol{\cdot}dm^{-3})$,Sn^{4+} $(1\ mol{\cdot}dm^{-3}) \parallel Ag^{+}$ $(1\ mol{\cdot}dm^{-3}) \mid Ag(s)$ (3)

2.1.4 $E^{\ominus}cell = E^{\ominus}_{reduction} - E^{\ominus}_{oxidation}$ ✓

$E^{\ominus}cell = +0,80$ ✓ $- (0,15)$ ✓

$E^{\ominus}cell = 0,65\ V$ ✓ (4)

2.2.1 Magnesium becomes smaller. ✓/ A brown solid forms. /

Mg disappears/ is eaten away/ Mg changes colour. (1)

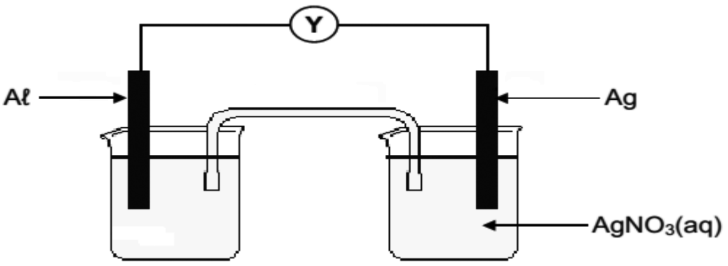
2.2.2 Cu^{2+} is a stronger oxidising agent (than Mg^{2+})✓ and will be reduced✓ to Cu. ✓ **OR**

Mg is a stronger reducing agent (than Cu) ✓ and will reduce Cu^{2+}✓ to Cu. ✓ (3)

[15]

Exercise 3

3.1 Learners set up a galvanic cell and measure its emf under standard conditions.



3.1.1 Write down the name of component Y. (1)

3.1.2 Is Aℓ the **anode** or the **cathode?** (1)

3.1.3 Write down the overall (net) cell reaction that takes place in this cell when it is working. (3)

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

3.2 Consider the half-cells, **P, Q** and **R,** represented in the table below:

HALF-CELL		
P	**Q**	**R**
$Zn \mid Zn^{2+}(aq)$	$C\ell_2 \mid C\ell^{-}(aq)$	$Cu \mid Cu^{2+}(aq)$

Different combinations of the half-cells above are compared to determine the highest emf produced under standard conditions.

3.2.1 Write down the NAME of a suitable electrode for the half-cell Q. (1)

3.2.2 State the standard conditions under which the half-cells should operate to ensure a fair comparison. (2)

3.2.3 Write down the **name** or **formula** of the strongest reducing agent in the half-cells above. (1)

3.2.4 Which combination of half-cells will produce the highest emf? Choose from **PR, PQ** or **QR.** (NO calculation is required.) (1)

[14]

Solutions

3.1.1 Voltmeter/ multimeter ✓ (1)

3.1.2 ANODE ✓ (1)

3.1.3 $3Ag^{+}(aq) + A\ell(s)$ ✓ $\qquad 3Ag(s) + A\ell^{3+}(aq)$. ✓ balancing✓ (3)

3.1.4 $E^{\ominus}cell = E^{\ominus}_{reduction} - E^{\ominus}_{oxidation}$ ✓

$E^{\ominus}cell = +0,80$ ✓ $- (- 1,66)$ ✓

$E^{\ominus}cell = 2,46$ V ✓ (4)

3.2.1 Platinum/ Pt/ Carbon/ C. (1)

3.2.2 **Any two:**

Concentration: 1 $mol{\cdot}dm^{-3}$ ✓

Temperature: 25 °C/298 K ✓

Pressure: 101,3 kPa/1,01 x 105 Pa/1 atm (2)

3.2.3 Zinc/ Zn ✓ (1)

3.2.4 PQ ✓ (1)

[14]

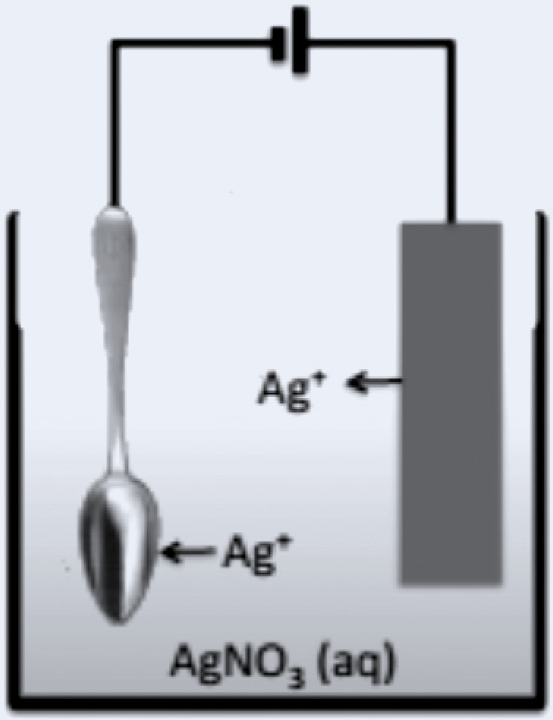
ELECTROLYTIC CELLS

- An electrolytic cell reaction is always a non-spontaneous, endothermic reaction, which requires a battery (energy/ power supply).
- The electrical energy is converted to chemical energy.
- Two electrodes are in the same container.
- Anode - where oxidation takes place – positive electrode.
  Cathode - where reduction takes place – negative electrode.
- The anode and cathode are connected to an external circuit, which is connected to a DC power source.

- Application of electrolytic cells are:
  - Electroplating
  - Metal refining/ Cu refining.
  - Extraction of aluminium.

- ***Electroplating:***
- Electroplating is used to protect metals that oxidise easily, by covering them with a thin layer of a metal that does not oxidise easily, e.g. chromium, silver or gold.
- A relatively cheap metal is covered by an expensive metal.
- Silver is used to cover cutlery, as it is too expensive to make a spoon of pure silver, and it is too weak to use. Chromium can be used to cover car parts like bumpers.

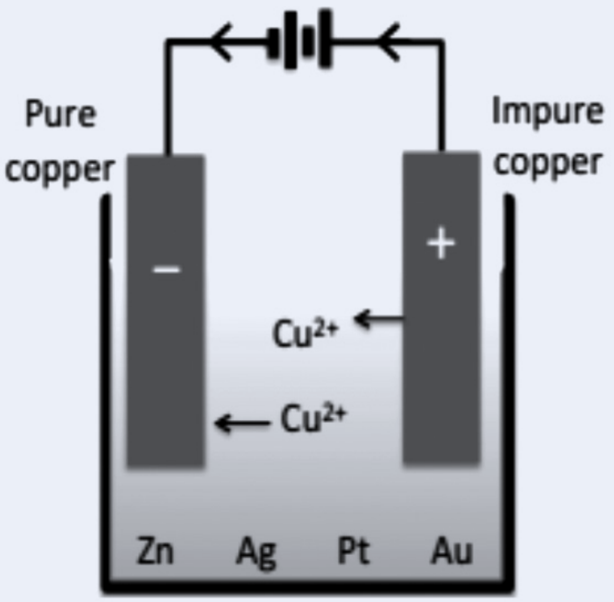
- For example: electroplating a spoon with silver. When the half reactions do not include conductors/ metals (i.e. when the half-cell involves gases/ ions), unreactive/ inert electrodes are used, e.g. carbon or platinum.
- The anode is silver. It will be oxidised to Ag+ ions. The mass of the silver electrode decreases.
- The cathode is the object (spoon) to be plated. The Ag^+ ions from the electrolyte will be reduced to form silver metal, which plates the spoon.
- The anode and electrolyte always contain the plating metal.



- ✓ The concentration of ions in the solution remains constant, as they are produced at the same rate as they are used. Rate of oxidation = rate of reduction.
- ✓ Oxidation half-reaction (anode): $Ag(s) \rightarrow Ag^+(aq) + e^-$
- ✓ Reduction half-reaction (cathode): $Ag^+(s) + e^- \rightarrow Ag(s)$
- ✓ Nett cell reaction : $Ag^+(aq) + Ag(s) \rightarrow Ag^+(aq) + Ag(s)$

✓ ***Refining of copper:***

- When copper is purified, the process is similar to that of electroplating.
- Impure copper is used, as the anode and the cathode is pure copper.
- At the anode, the copper is oxidised to produce Cu^{2+} ions in the electrolyte.
- The mass of the impure copper anode decreases.
- At the cathode, the Cu^{2+} ions in the electrolyte are reduced to form a pure copper layer on the cathode.
- The mass of the cathode increases.



✓ The other elements and compounds found in the impure copper anode are precipitated to the bottom of the reaction vessel.

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

✓ Reduction half-reaction (cathode): $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

✓ Nett cell reaction: $Cu^{2+}(aq) + Cu(s) \rightarrow Cu(s) + Cu^{2+}(aq)$

✓ ***Electrolysis of copper (II) chloride***

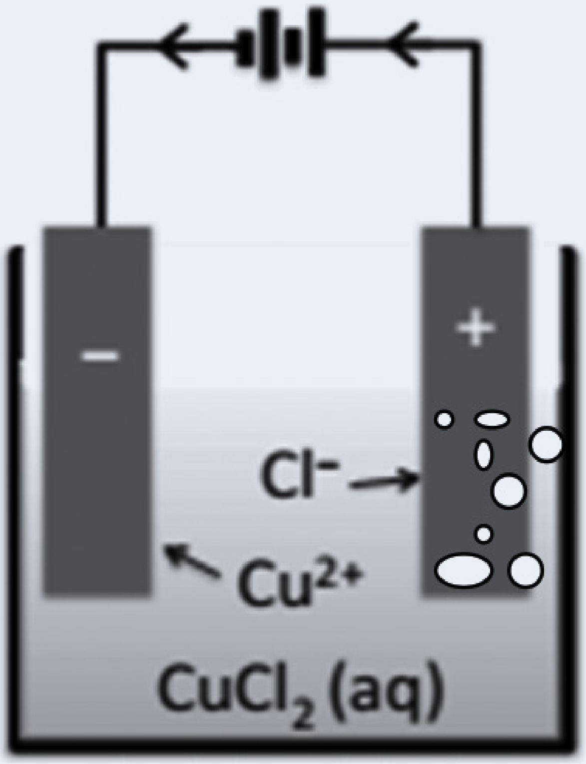
- Oxidation half-reaction (anode):
  $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$
- Reduction half-reaction (cathode):
  $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
- Nett cell:
  $Cu^{2+}(aq) + 2Cl^-(aq) \rightarrow Cl_2(g) + Cu(s)$

✓ Chlorine gas is produced at the anode, while copper metal is produced at the cathode.

✓ Observations:

✓ Bubbles of chlorine gas will be observed at the anode.

✓ The cathode will grow thicker with time because copper metal will deposit at the cathode.



Extraction of aluminium

- Aluminium is found in the mineral known as bauxite, which contains primarily aluminium oxide (Al_2O_3) in an impure form.
- Bauxite is not found in South Africa, so it is imported from Australia for refining.

STEP 1

Converting impure Al_2O_3 to pure Al_2O_3 .

Bauxite treated with NaOH – impure Al_2O_3 becomes $Al(OH)_3$.

$Al(OH_{)3}$ is heated (T > 1000 °C).

$Al(OH)_3$ becomes pure Al_2O_3, alumina.

STEP 2

Melting Al_2O_3 Alumina is dissolved in cryolite (sodium aluminium hexafluoride – Na_3AlF_6).

The melting point of $A\ell_2O_3$ is higher than 2 000°C. Cryolite is added to the ore before it is heated.

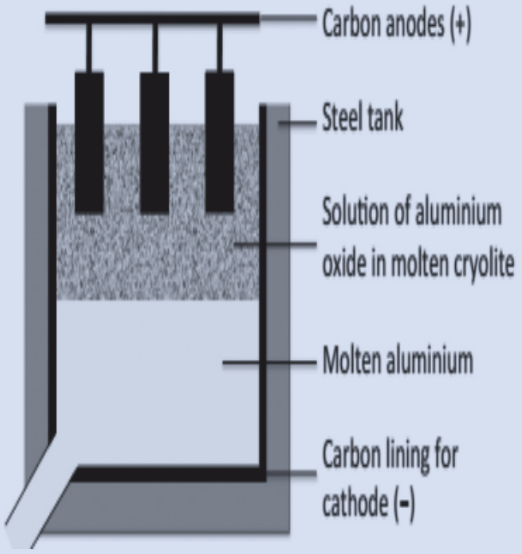
Cryolite does not lower the melting point of $A\ell_2O_3$ -it dissolves it.

$A\ell_2O_3$ dissolved in molten cryolite can be electrolysed easily.

⇒Less electricity is required to extract the aluminium, thereby decreasing the extraction cost.

STEP 3

Molten alumina (Al_2O_3) – cryolite mixture placed in reaction vessel.



- ➢ Owing to the high temperature of the reaction, the oxygen produced reacts with the carbon electrodes to produce carbon dioxide gas.

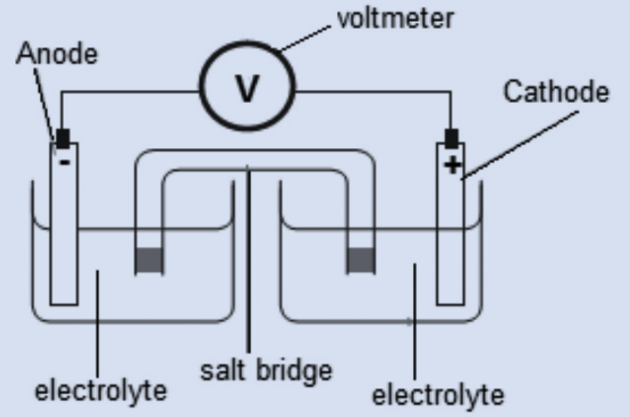
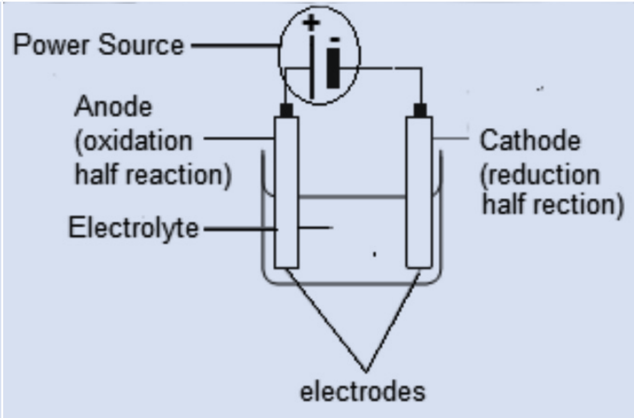
  The carbon electrodes therefore need to be replaced regularly.

  $C + O_2 \rightarrow CO_2$
- ➢ Aluminium extraction uses a large amount of electrical energy ∴ the cost of aluminium extraction is very high.

- ✓ Anodes (+) are carbon rods in a mixture.
- ✓ Cathode (-) is the carbon lining of the tank.
- ✓ At cathode Al^{3+} ions are reduced to Al metal.
- ✓ Oxidation half-reaction (anode): $2O^{2-}(aq) \rightarrow O_2(g) + 4e^-$
- ✓ Reduction half-reaction (cathode): $Al^{3+}(aq) + 3e^- \rightarrow Al(l)$
- ✓ Nett cell reaction : $2Al_2O_3(aq) \rightarrow 4Al(l) + 3O_2(g)$

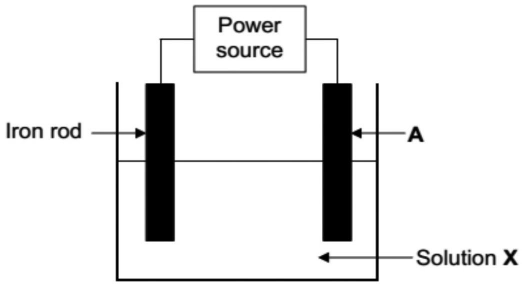
ENVIRONMENTAL IMPACT OF THE EXTRACTION OF ALUMINIUM:

- The ecological impact of Aℓ extraction includes: 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 rivers, lagoons and ground water.
- The environmental impact of Aℓ extraction include the following: carbon dioxide from the burning of the anodes contributes to the greenhouse effect, which causes global warming. Fluorine and its compounds lost are from the cryolite during the electrolysis process and are poisonous. Chemicals in the red mud dams drain into the soil and contaminate ground water.
- Pollution caused by power generation (for electrolytic process) using coal-fired plants leads to acid rain and adds to the greenhouse effect.
- There is noise pollution from the extraction plant.

GALVANIC CELL	ELECTROLYTIC CELL
	
✓ A cell in which chemical energy is converted into electrical energy. ✓ Spontaneous: has self-sustaining electrode reactions. ✓ Exothermic reactions. ✓ Salt bridge ✓ Positive electrode: **C**athode: Redu**C**tion ✓ Negative electrode: An**O**de: Oxidation ✓ Two half-cells are placed in separate containers. ✓ Emf is always positive.	✓ A cell in which electrical energy is converted into chemical energy. ✓ Non-spontaneous: needs a continuous supply of (electrical) energy. ✓ Endothermic reactions. ✓ No salt bridge. ✓ Positive electrode: An**O**de: Oxidation ✓ Negative electrode: **C**athode: Redu**C**tion ✓ Two electrodes are in the same container. ✓ Emf is always negative.

Exercise 4

The diagram below shows an electrolytic cell that is used to electroplate an iron rod with **copper.** Solution **X** is made up of an unknown **nitrate.**



4.1 Solutions such as solution X are always used in electrochemical cells.

4.1.1 Write down the general term used to describe these solutions. (1)

4.1.2 What is the function of these solutions in electrochemical cells? (1)

4.2 Write down the **formula** of solution **X.** (1)

4.3 Which electrode (**A** or **iron rod**) is the negative electrode? Give a reason for your answer. (2)

4.4 Write down the half reaction that takes place at electrode **A.** (2)

4.5 Electrode **A** is now replaced by a silver rod without making any other changes to the cell. After a while, **two** metallic ions are found to be present in the solution.

4.5.1 Name the **two** metallic ions present in the solution. (2)

4.5.2 Refer to the relative strength of oxidising agents to explain which **one** of the two ions will preferably be involved in the plating process. (2)

[11]

Solutions

4.1.1 Electrolyte✓ (1)

4.1.2 Conduct electricity/ carry charges. ✓ (1)

4.2 $Cu(NO_3)_2$ ✓ (1)

4.3 Iron rod ✓

Reduction takes place (on the iron rod) ✓ (2)

4.4 $Cu \rightarrow Cu^2 + + 2e-$ ✓✓ (2)

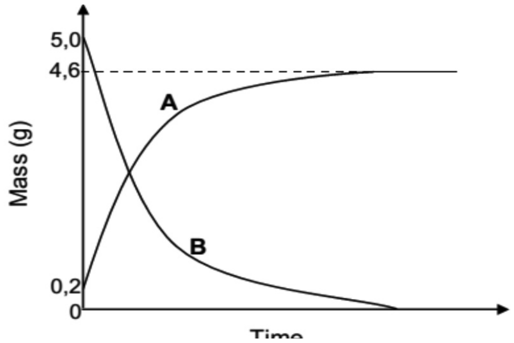
4.5.1 Copper(II) (ions)/Cu^{2+} ✓ and silver (ion)/Ag^+.✓ (2)

4.5.2 Ag+/silver (I) ions is a stronger oxidising agent than $Cu^{2+/}$copper (II) ions✓ and will be reduced more readily to form silver/Ag on the iron rod. ✓ (2)

[11]

Exercise 5

The graph below represents the changes in mass that occur at electrode **A** and electrode **B** in an electrolytic cell during purification of copper.



5.1 Define *electrolysis.* (2)

5.2 Which graph, **A** or **B**, represents the change in mass of the anode during electrolysis? Motivate your answer. (2)

5.3 Write down the equation of the half-reaction that takes place at the cathode of this cell. (2)

5.4 Use the information in the graph to calculate the percentage purity of the impure copper. (4)

[10]

Solutions

5.1 The chemical process by which electrical energy is converted to chemical energy.
OR
The use of electrical energy to produce a chemical change. (2)

5.2 **B.**
The anode where oxidation takes place will be eroded and becomes small in size. (2)

5.3 $Cu^{2+} + 2e^{-} \rightarrow Cu$ (2)

5.4 $\% \text{ Purity} = \dfrac{m(Cu)}{m(Cu)_{\text{impure}}} \times 100$

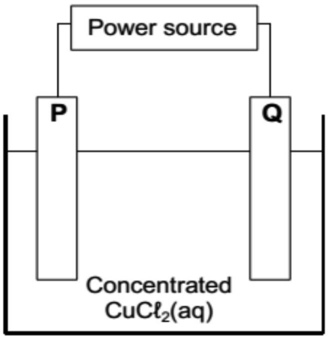
$\% \text{ Purity} = \dfrac{4,4✓}{5✓} \times 100✓$

$\% \text{ Purity} = 88\%✓$ (4)

[10]

Exercise 6

The simplified diagram below represents an electrochemical cell used to refine copper. One of the electrodes consists of impure copper.



6.1 What type of power source, AC or DC, is used to drive the reaction in this cell? (1)

6.2 When an electric current passes through the $CuC\ell_2(aq)$, the mass of electrode **P** increases.

Is electrode P the cathode or the **anode?** Write down the relevant half-reaction to support your answer. (3)

6.3 The impure copper contains zinc impurities that are oxidised to zinc ions.

Refer to the relative strengths of oxidising agents to explain why zinc ions will not influence the quality of the pure copper produced in this cell. (3)

6.4 Electrodes **P** and **Q** are now replaced by carbon electrodes.

6.4.1 What will be observed at electrode **Q**? (2)

6.4.2 How will the concentration of the electrolyte change as the reaction proceeds? Choose from i**ncreases, decreases** or **remains the same**. (2)

[11]

Solutions

6.1 DC ✓ (1)

6.2 Cathode ✓ ✓

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

6.3 Cu^{2+} is a stronger oxidising agent ✓ than Zn^{2}+ ions ✓; therefore Zn^{2+} ions will not be reduced (to Zn). ✓ (3)

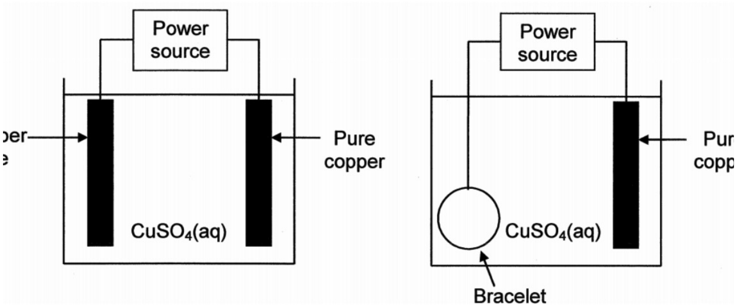
6.3.1 (Chlorine) gas/ bubbles is/ are formed. ✓✓ (2)

6.3.2 Decreases ✓✓ (2)

[11]

Exercise 7

Below are simplified diagrams of two electrochemical cells, **A** and **B**, which are used in **industry.** Cell **A** is used in the purification of copper ore containing silver and platinum impurities. Cell **B** is used to electroplate a bracelet with a layer of copper.



7.1 Write down the name of the type of electrochemical cell (**electrolytic** or **galvanic**) of which the above two cells are examples. (1)

7.2 Pure copper is used as one of the electrodes in each of the cells above. In which cell (A or B) is the pure copper the:

7.2.1 cathode? (1)

7.2.2 anode? (1)

7.3 Consider cell B. Initially, $CuSO_4(aq)$ has a blue colour.

7.3.1 How will the intensity of the blue colour change whilst the cell is functioning? Write down **increases, decreases** or **remains the same.** Give a reason for your answer. (3)

7.3.2 Write down the reaction that takes place at the pure copper electrode. (2)

7.4 Consider cell A:

7.4.1 Give a reason why the sludge formed in this cell is of economic importance. (1)

7.4.2 Name **one** negative impact that the energy usage in this process has on the environment. (2)

Solutions **[11]**

7.1 Electrolytic✓ (1)

7.2.1 A ✓ (1)

7.2.2 B ✓ (1)

7.3.1 Remains the same. ✓ The rate of oxidation of copper at the anode is equal to the rate of reduction of copper (II) ions at the cathode. ✓✓ (3)

7.3.2 $Cu \rightarrow Cu^{2+} + 2e^{-}$ ✓✓ (2)

7.4.1 It contains precious metals/ valuable/ expensive metals. ✓ (1)

7.4.2 Consumes a large amount of electricity/ energy. ✓
Depletes coal resources. ✓ **OR** Contributes to global warming. **OR** Habitats are destroyed when mining coal. **OR** Contributes to acid rain. (2)

[11]

6. MESSAGE TO GRADE 12 LEARNERS FROM THE WRITERS

The author would like to wish you well in your examinations at the end of the year. Make use of this material and your textbook to prepare for the finals.

Remember that you need to start now and not on the eve of the examination. It is never too early to start.

7. THANK YOU AND ACKNOWLEDGEMENTS

A big thank you to all who contributed to the writing of this booklet. It is our hope that it supports your learning and enhances your understanding of the topic discussed here.

PHYSICAL SCIENCE

ELECTROCHEMISTRY : PAPER 2

GRADE 12

Department of Basic Education
222 Struben Street, Pretoria, 0001
Private Bag X895, Pretoria, 0001, South Africa
Tel: (012) 357 3000 Fax: (012) 323 0601

Hotline: 0800 202 933

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ISBN NO. 978-1-4315-3299-5