PHYSICAL SCIENCE_PAPER 2
STOICHIOMETRY
GRADE 12
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 practices to ensure mastery of the topic 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 bookletpocussing on only 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 booklet 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.
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1. How to use this Study Guide

This book serves as a guide to understanding the material prescribed for the Grade 12 Physical Sciences subject. However, it does not 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 more on the areas that are seen as a challenge, given learners’ responses in the Grades 12 National examinations over the past few years.

This guide aims to help explain the usual concepts in your regular textbook. It also offers more exercises and examples that serve as building blocks to your own understanding of what is expected of you in this subject. The book draws on the basic knowledge obtained in lower grades and demonstrates how this knowledge fits in with the new material in Grade 12, starting from atomic theory. Figure 1 shows how studying chemistry is related to atomic theory. The atomic theory is a framework for the whole chemistry curriculum. It is therefore important to remember this theory, as well as the properties of the parts of an atom and the role they play in all 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 TO THE REST OF THE CHEMISTRY. Use information learnt in other sections to solve the problem 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 you understand the steps taken to solve the question. Then try the rest of the questions without looking at the solutions. After completing an exercise, check your solutions in the answer section. Move on to the rest of the questions and try to understand why you were wrong if you got the answer wrong. The solutions to all exercises are provided in the last section of this booklet.

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

2. Study and Examination Tips

Pay attention to your health. Here are some tips:
1. Eat wholesome food like grains, fruits and vegetables, and drink enough water.
2. Engage in light exercise like walking.
3. Get enough sleep and rest.
4. Develop a study timetable and try to be adhere to the time allocated to studying.

2.1 Subject specific

As you prepare to write the examination, it is important to carefully understand the rules that govern certain aspects of your work, i.e.: definitions, rules, laws and concepts. Understand these definitions/ rules/ laws/ concepts very well. Understand what they mean, where they apply and when they apply - and also when and where they do not apply. Also, 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 the question 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 to solve the question.
7. All questions have hints that point to the answer.
8. Check your work by going through these steps again.

2.2 Topic specific

Stoichiometry underlies all calculations in chemistry. Understanding stoichiometry will help you do calculations in any area of chemistry.

This topic begins with a short introduction on the mole concept. This is an important aspect of stoichiometry, since all calculations require knowledge of the molar mass, the mass and the number of moles. The activities in this booklet all have solutions. You are required to look how the first few solutions have been answered, and then do the rest first, without looking at the solutions, and only confirm the steps and the answers after trying them first. The exercises cover all the different topics, to demonstrate that stoichiometry will be used across an array of topics.

4. The Atomic Theory

The theory states that all matter is made up of tiny indivisible particles, called atoms. According to modern interpretations of the theory, the atoms of each element are effectively identical (except for isotopes), but differ from those of other elements, and combine to form compounds in fixed proportions.

The atomic theory underlies everything in science. The atomic theory was developed through the participation and cooperation of many scientists and a variety of experiments. While no-one has ever seen an atom, the experiments carried out over the years on the material world made scientists come to certain conclusions that led them to agree that everything in the world - including living and non-living matter - is made up of very small particles, called atoms. Atoms themselves are made up of sub-particles called: electrons, protons and neutrons. (Read the history of this development). In explaining the nature of the atom, the following is accepted as the nature of the atom:

The atom has a nucleus, made up of neutrons and protons. The electrons are negatively charged and circle the atom at a high speed. This motion never stops. According to a calculation done by some scientists, the speed is about 2,200 kilometres per second: so an electron can circle the Earth in just over 18 seconds!

Different substances are made up of different atoms. Atoms differ in terms of the following:

1. Number of protons in each atom.
2. Number of electrons in each atom.
3. Number of neutrons in each atom.
4. Number of energy levels on which electrons orbit in each atom.

All these differences give atoms different physical and chemical properties, and different intrinsic and extrinsic properties. The second section of Chemistry is concerned with products that are formed from the reaction of one set of atoms with another set of atoms. It is this formation of new products that is the subject of stoichiometry – specifically, knowing how much of these tiny particles are needed in order for a reaction to occur. Since atoms are extremely small, a way of measuring them was devised, i.e. the mole concept.
4.1 The Mole Concept

Mole: The amount of a substance that contains $6.02 \times 10^{23}$ particles. This number is also called Avogadro’s number.

One MOLE has $6.022 \times 10^{23}$ items. So, there are $6.022 \times 10^{23}$ items in a mole.

Quick Activity
1. How many atoms of potassium make up one Mole? ___________
2. How many atoms of potassium make up 2 Moles? ___________
3. How many formula units of salt make up 10 Moles? ___________
4. How many molecules of water make up 1 Mole? ___________
5. How many molecules of water make up 5 Moles? ___________
6. How many moles are $6.022 \times 10^{23}$ atoms of sodium? ___________
7. How many moles are $12.04 \times 10^{23}$ atoms of carbon? ___________
8. How many moles are $18.06 \times 10^{23}$ atoms of sodium? ___________
9. How many moles are $60.22 \times 10^{23}$ atoms of sodium? ___________
10. How many moles are $6.022 \times 10^{23}$ molecules of water? ___________
11. How many moles are $12.04 \times 10^{23}$ molecules of water? ___________
12. How many moles are $30.10 \times 10^{23}$ molecules of water? ___________

All chemistry calculations require an understanding of the mole concept. When we carry out calculations on reactions in chemistry, we always use the moles that react - NEVER the mass that reacts. If you are given the mass, you must always work out the number of moles contained in the mass.

In a chemical reaction, the reactants with a higher number of moles than is necessary will have some unreacted molecules(s) at the end of a reaction, the unreacted reactant(s) is/are called excess reactant(s). The reactant that is used up while there is still an excess is called a limiting reactant(s).

Calculate the mole of a given mass as follows:

4.2 Excess and Limiting Reagents: Concrete Models

If you have 6 car bodies and 48 tyres, how many cars with tyres can you produce?

See how many tyres you are able to use in this example. The remaining tyres are in excess. The cars are said to be limiting.
Which of the two reactants is in excess and which is the limiting reagent in the following diagrams?

Which of the two molecules, $\text{H}_2$, $\text{N}_2$, is a limiting reagent?

4.3 Worked Examples from Grade 10

Reminder:

At STP: 1 mole of any gas occupies 22.4 dm$^3$ at 0 °C (273 K) and 1 atmosphere.

1. Find the number of moles of ions in:
   (a) 2 moles of $\text{Fe}_2(\text{SO}_4)_3$
       From the formula, we know that 1 mole of $\text{Fe}_2(\text{SO}_4)_3$ contains 2 moles of Fe$^{3+}$ ions and 3 moles of SO$_4^{2-}$ ions.
       \[
       \text{No. of moles of Fe}^{3+} \text{ ions} = 2 \times 2 = 4 \\
       \text{No. of moles of SO}_4^{2-} \text{ ions} = 3 \times 2 = 6 \\
       \text{Total no. of moles of ions} = 4 + 6 = 10
       \]
   (b) 0.2 moles of $\text{Al(NO}_3)_3$
       \[
       \text{No. of moles of Al}^{3+} \text{ ions} = 0.2 \\
       \text{No. of moles of NO}_3^{-} \text{ions} = 0.2 \times 3 = 0.6 \\
       \text{Total no. of moles of ions} = 0.2 + 0.6 = 0.8
       \]

2. Given 1.6g of methane ($\text{CH}_4$), find:
   (a) The number of moles of $\text{CH}_4$
       Molar mass of $\text{CH}_4 = 12 + 1 \times 4 = 16 \text{ gmol}^{-1}$
       \[
       \text{Number of moles of } \text{CH}_4 = \frac{\text{mass of } \text{CH}_4}{\text{molar mass of } \text{CH}_4} = \frac{1.6}{16} = 0.10
       \]
   (b) The number of molecules of $\text{CH}_4$
       \[
       \text{Number of molecules of } \text{CH}_4 = \text{no. of moles} \times \text{Avogadro Number} \\
       = 0.10 \times 6.02 \times 10^{23} \approx 6.02 \times 10^{22}
       \]
(c) The number of H atoms.  
One CH$_4$ molecule contain 4 H atoms.  
\[
\text{number of H atoms} = 6.02 \times 10^{23} \times 4 = 2.41 \times 10^{23}
\]

3. Find the mass of:

<table>
<thead>
<tr>
<th>(a) 1 H$_2$O molecule</th>
<th>(b) 1 Cu atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>18 / 6.02 x 10$^{23}$</td>
<td>63.5 / 6.02 x 10$^{23}$</td>
</tr>
<tr>
<td>3.0 x 10$^{-23}$ g</td>
<td>1.05 x 10$^{-22}$ g</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>(c) 1 Na$^+$ ion</th>
<th>(d) 1 OH$^-$ ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>23 / 6.02 x 10$^{23}$</td>
<td>17 / 6.02 x 10$^{23}$</td>
</tr>
<tr>
<td>3.8 x 10$^{-23}$ g</td>
<td>2.8 x 10$^{-23}$ g</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>(e) 1 neutron</th>
<th>(f) 1 electron</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.0 / 6.02 x 10$^{23}$</td>
<td>0.00055 / 6.02 x 10$^{23}$</td>
</tr>
<tr>
<td>1.7 x 10$^{-24}$ g</td>
<td>9.1 x 10$^{-28}$ g</td>
</tr>
</tbody>
</table>

4. Find the mass of:

- **6.02 x 10$^{22}$ lead atoms**
  
  Number of moles of Pb atoms  
  = 6.02 x 10$^{22}$ / 6.02 x 10$^{23}$  
  = 0.100  
  
  Mass of Pb atoms  
  = 0.100 x 207  
  = 20.7 g

- **3.01 x 10$^{24}$ carbon dioxide molecules**
  
  Number of moles of CO$_2$ molecules  
  = 3.01 x 10$^{24}$ / 6.02 x 10$^{23}$  
  = 5.00  
  
  Molar mass of CO$_2$  
  = 12 + 16 x 2 = 44 g  
  
  Mass of CO$_2$ molecules  
  = 5.00 x 44  
  = 220 g

- **3.01 x 10$^{23}$ sulphate ions**
  
  Number of moles of SO$_4^{2-}$ ions  
  = 3.01 x 10$^{23}$ / 6.02 x 10$^{23}$  
  = 0.500  
  
  Molar mass of SO$_4^{2-}$  
  = 32 + 16 x 4 = 96 g  
  
  Mass of SO$_4^{2-}$ ions  
  = 0.500 x 96  
  = 48 g

5. (a) How many molecules are there in 3.00 moles of oxygen molecules?
  
  Number of oxygen molecules  
  = 3.00 x 6.02 x 10$^{23}$  
  = 1.806 x 10$^{24}$

(b) How many ions are there in 0.600 moles of potassium ions?
  
  Number of potassium ions  
  = 0.600 x 6.02 x 10$^{23}$  
  = 3.612 x 10$^{23}$

(Molar mass: O = 16.0, K = 39.0)
6. Calculate the number of moles of atoms in:
   (a) 127 g of copper
       Number of moles of copper
       = 127 / 63.5
       = 2 mol
   (b) 12.8 g of sulphur.
       Number of moles of sulphur
       = 12.8 / 32.0
       = 0.4 mol
       (Molar Mass of Cu = 63.5, S = 32.0)

7. How many atoms are there in:
   (a) 2.50 moles of oxygen atoms?
       Number of oxygen atoms
       = 2.50 x 6.02 x 10^{23}
       = 1.505 x 10^{24}
   (b) 6.00g of magnesium atoms?
       Number of moles of magnesium atoms
       = 6.00 / 24.0 = 0.25 mol
       Number of magnesium atoms
       = 0.25 x 6.02 x 10^{23}
       = 1.505 x 10^{23}

8. What is the mass of 2.50 moles of magnesium atoms?
   (Relative atomic mass: Mg = 24.0)
   Mass of magnesium atoms
   = 2.50 x 24.0
   = 60.0 g

More Worked examples:
1. What is the empirical formula of a compound containing: 40.0% C; 6.71% H; 53.28% O?
   Solution:

<table>
<thead>
<tr>
<th>Elements (atoms)</th>
<th>% Composition divided by molar mass</th>
<th>Divide ALL the answers in the left column by the lowest number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon (C)</td>
<td>$\frac{40.0}{12} = 3.33$</td>
<td>$\frac{3.33}{3.33} = 1$</td>
</tr>
<tr>
<td>Hydrogen (H)</td>
<td>$\frac{6.71}{1} = 6.71$</td>
<td>$\frac{6.71}{3.33} = 2$</td>
</tr>
<tr>
<td>Oxygen (O)</td>
<td>$\frac{53.28}{16} = 3.33$</td>
<td>$\frac{3.33}{3.33} = 1$</td>
</tr>
</tbody>
</table>

   C = 1; H = 2; O = 1
   The empirical formula = CH₂O
2. Determine the molecular formula of the compound in Question 1, if its molar mass is 180g mol\(^{-1}\).

\[ M_{\text{CH}_2\text{O}} = 30 \text{ g mol}^{-1} \]

Ratio \(\frac{180}{30} = 6\)

Molecular formula (multiply each of the number of atoms in the empirical formula by 6.)
Molecular formula = \(C_6H_{12}O_6\)

a. Nicotine is an alkaloid in the nightshade family of plants that is mainly responsible for the addictive nature of cigarettes. It contains: 74.02% C; 8.710% H; 17.27% N. If 40.57g of nicotine contains 0.25 mol nicotine, what is the molecular formula?

**Solution**

<table>
<thead>
<tr>
<th>Elements (atoms)</th>
<th>% Composition divided by molar mass</th>
<th>Divide ALL the answers in the left column by the lowest number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon (C)</td>
<td>(\frac{74,02}{12} = 6,168)</td>
<td>(\frac{6,168}{1,229} = 5)</td>
</tr>
<tr>
<td>Hydrogen (H)</td>
<td>(\frac{8,71}{1} = 8,71)</td>
<td>(\frac{8,71}{1,229} = 7)</td>
</tr>
<tr>
<td>Oxygen (O)</td>
<td>(\frac{17,2}{14} = 1,229)</td>
<td>(\frac{1,229}{1,229} = 1)</td>
</tr>
</tbody>
</table>

Empirical formula is \(C_5H_7N\)

\[ \frac{40,57}{M} = 0,25 \]

\[ 162,28 \text{ g mol}^{-1} = MNicotine \]

\[ M(C_5H_7N) = 81,14 \]

Ratio \(\frac{162,28}{81,14} = 2\)

Molecular formula: multiply each of the number of atoms in the empirical formula by 2.

Molecular formula = \(C_{10}H_{14}N_2\)

A pure sample of calcium chloride \(\text{CaCl}_2\) was found to contain 7.10 g of \(\text{Cl}^-\) ions. What mass of \(\text{Ca}^{2+}\) ions does the sample contain?

**Number of moles of \(\text{Cl}^-\) ions**

\[ = \frac{7.10}{35.5} = 0.200 \]

*The formula \(\text{CaCl}_2\) shows that the ratio of \(\text{Ca}^{2+}\) ions to \(\text{Cl}^-\) ions is 1 : 2.

Hence, the number of moles of \(\text{Ca}^{2+}\) ions \(= \frac{0.200}{2} = 0.100\)

Mass of \(\text{Ca}^{2+}\) ions in the sample \(= 0.100 \times 40 \]

\(= 4.0 \text{ g}\)
3. A metal (M) ionizes to give $M^{n+}$ ions. If the atomic mass of $M$ is 24 and 1.2g of $M$ ionizes to give $6.02 \times 10^{22}$ electrons, calculate $n$ (the charge on each ion of $M$).

**Number of moles for electrons given**

$$\text{Number of moles for electrons given} = \frac{\text{number of electrons}}{\text{Avogadro Number}}$$

$$= \frac{6.02 \times 10^{22}}{6.02 \times 10^{23}} = 0.100$$

**Number of moles of $M$ ionized**

$$= \frac{1.2}{24} = 0.050$$

0.050 moles of $M$ gives 0.100 moles of electrons on ionization.

Thus, 1 mole of $M$ gives $0.100 / 0.050 = 2$ moles of electrons.

Since the charge on an ion of $M$ is numerically equal to the number of moles of electrons given by 1 mole of $M$, each ion of $M$ carries 2 charges, i.e. $n = 2$.

5. Use the table below

<table>
<thead>
<tr>
<th>Substance</th>
<th>Molar mass of substance</th>
<th>Mass of substance present</th>
<th>Number of moles of substance present</th>
<th>Number of molecules/ formula units present</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sulphuric acid ($H_2SO_4$)</td>
<td>98g mol$^{-1}$</td>
<td>58.8g</td>
<td>0.6 mol</td>
<td>$3.612 \times 10^{23}$</td>
</tr>
<tr>
<td>Sodium hydroxide (NaOH)</td>
<td>40g mol$^{-1}$</td>
<td>2.0g</td>
<td>0.05 mol</td>
<td>$3.01 \times 10^{22}$</td>
</tr>
<tr>
<td>Potassium carbonate ($K_2CO_3$)</td>
<td>138g mol$^{-1}$</td>
<td>331.2g</td>
<td>2.4 mol</td>
<td>$1.44 \times 10^{24}$</td>
</tr>
</tbody>
</table>

(Relative atomic mass numbers: C = 12.0; H = 1.0; K = 39.0; Na = 23.0; O = 16.0; S = 32.0.)

6. If we breathed in $3913 \times 10^{18}$ molecules of air pollutant, nitrogen dioxide ($NO_2$), how many grams of NO$_2$ would we breathe in?

(Relative atomic mass numbers: N = 14.0; O = 16.0)

Number of moles of nitrogen dioxide:

$$= \frac{3913 \times 10^{18}}{6.02 \times 10^{23}}$$

$$= 6.5 \times 10^{-3} \text{mol}$$

Mass of nitrogen dioxide:

$$= 6.5 \times 10^{-3} \times (14.0 + 2 \times 16.0)$$

$$= 0.299g$$

7. (I) How many moles of calcium fluoride (CaF$_2$) are present in 16.5g of it?

**Formula mass of CaF$_2$**

$$= 40.0 + 19.0 \times 2$$

$$= 78.0 \text{ g}$$

Number of moles of CaF$_2$

$$= \frac{16.5}{78.0} = 0.21 \text{ mol}$$

(II) How many calcium and fluoride ions are present?

**Number of Ca$^{2+}$ ions**

$$= 0.21 \times 6.02 \times 10^{23} = 1.26 \times 10^{23}$$

**Number of F$^{-}$ ions**

$$= 0.42 \times 6.02 \times 10^{23} = 2.53 \times 10^{23}$$

(Relative atomic mass numbers: Ca = 40.0; F = 19.0)
8. What mass of water contains the same number of molecules as 2.20g of carbon dioxide?

(Number of moles of CO$_2$) = $\frac{2.20}{(12.0 + 2 \times 16.0)} = 0.05$ mol

Mass of water
= $0.05 \times (1.0 \times 2 + 16.0)$
= 0.9g

9. Identity the substance that contains the greater number of molecules from each set:

(a) 2 moles of carbon dioxide molecules (CO$_2$) or 8.40g of sulphuric acid (H$_2$SO$_4$)

Number of CO$_2$ molecules
= $2 \times 6.02 \times 10^{23}$
= $1.20 \times 10^{24}$

Number of moles of H$_2$SO$_4$
= $\frac{8.40}{(1.0 \times 2 + 32.0 + 16.0 \times 4)} = 0.086$ mol

Number of H$_2$SO$_4$ molecules
= $0.086 \times 6.02 \times 10^{23}$
= $5.16 \times 10^{22}$

(greater number of molecules)

(b) 88.0g of carbon dioxide (CO$_2$) or 84.0g of nitrogen (N$_2$)

Number of CO$_2$ molecules
= $\frac{88.0}{44.0} \times 6.02 \times 10^{23}$
= $1.20 \times 10^{24}$

Number of moles of N$_2$
= $\frac{84.0}{28.0} \times 6.02 \times 10^{23}$
= $1.81 \times 10^{24}$

(greater number of molecules)

(c) 5 x $10^{24}$ ammonia molecules (NH$_3$) or 8.00g of sulphuric acid (H$_2$SO$_4$)

Number of NH$_3$ molecules
= $5 \times 10^{24}$

(greater number of molecules)

Number of moles of H$_2$SO$_4$ molecules
= $\frac{8.00}{(1.0 \times 2 + 32.0 + 16.0 \times 4)} \times 6.02 \times 10^{23}$
= $4.91 \times 10^{22}$

(Stating atomic mass numbers: C = 12.0; H = 1.0; N = 14.0; O = 16.0; S = 32.0)

10. Calculate the molarity of each of the following solutions:

(a) 23.4g of sodium chloride (NaCl) solid in 1.0 dm$^3$ solution.

Number of moles of NaCl
= $\frac{23.4}{(23.0 + 35.5)}$
= 0.4 mol

Molarity of NaCl solution
= $0.4 / 1$
= 0.4 M

(b) 17.0 g of silver nitrate (AgNO$_3$) solid in 200.0 cm$^3$ solution.

Number of moles of AgNO$_3$
= $\frac{17.0}{(108.0 + 14.0 + 16.0 \times 3)}$
= 0.1 mol

Molarity of AgNO$_3$ solution
= $0.1 / (200.0 / 1000.0)$
= 0.5 M

(Stating atomic mass numbers: Ag = 108.0; Cl = 35.5; N = 14.0; Na = 23.0, O = 16.0)
17. Complete the table below:

<table>
<thead>
<tr>
<th>Substance</th>
<th>Molar mass (g mol⁻¹)</th>
<th>Concentration (g dm⁻³)</th>
<th>Molarity (M)</th>
<th>Mass of solute required to prepare 250.0 cm³ of solution (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>KOH</td>
<td>56.0 g mol⁻¹</td>
<td>5.60 g dm⁻³</td>
<td>0.10 M</td>
<td>5.60 g dm⁻³ × (250.0/1000) dm³ = 1.40 g</td>
</tr>
<tr>
<td>CuSO₄</td>
<td>159.5 g mol⁻¹</td>
<td>31.9 g dm⁻³</td>
<td></td>
<td>7.975 g</td>
</tr>
<tr>
<td>(COOH)₂H₂O</td>
<td>126.0 g mol⁻¹</td>
<td>30.24 g dm⁻³</td>
<td>0.24 M</td>
<td></td>
</tr>
</tbody>
</table>

(Relative atomic mass numbers: C = 12.0; Cu = 63.5; H = 1.0; N = 14.0; O = 16.0, S = 32.0)

18. Calculate the molarity of each of the following solutions:

(a) 500.0 cm³ of 0.10 M sodium hydrogencarbonate (NaHCO₃) solution.

Number of moles of NaHCO₃

= 0.10 x (500.0 / 1000.0) = 0.05 mol

Mass of NaHCO₃

= 0.05 x (23.0 + 1.0 + 12.0 + 16.0 x 3)
= 0.05 x 84
= 4.2 g

(b) 100.0 cm³ of 3.00 M potassium dichromate (K₂Cr₂O₇) solution.

Number of moles of K₂Cr₂O₇

= 3.00 x (100.0 / 1000.0) = 0.3 mol

Mass of K₂Cr₂O₇

= 0.3 x (39.0 x 2 + 52.0 x 2 + 16.0 x 7)
= 0.3 x 294
= 88.2 g

(Relative atomic mass numbers: C = 12.0; Cr = 52.0; H = 1.0; K = 39.0; Na = 23.0; O = 16.0)

19. 5.6g of a metal (M) combine with 2.4g of oxygen to form an oxide with the formula M₂O₃. What is the atomic mass of M?

Let the atomic mass of M be Ar.

\[
\frac{\text{Mass of M in compound}}{\text{Mass of oxygen in compound}} = \frac{2Ar}{16 \times 3} = \frac{5.6}{2.4} = 56
\]

Hence, the atomic mass of the metal M is 56.

20. A crystalline salt with the formula M₂S₂O₃ · 5H₂O is found to contain 36.3% by mass of water of crystallization. Calculate:

(a) the formula mass of the hydrated salt
(b) the atomic mass of the metal M.

(a) \[ \frac{5(1 \times 2 + 16)}{\text{formula mass of hydrated salt}} = \frac{36.3}{100} = 248 \]

Formula mass of hydrated salt = 248.

(b) Let the atomic mass of M be A.

Formula mass of hydrated salt

\[ = 2Ar₂ + 32 \times 2 + 16 \times 3 + 5(1 \times 2 + 16) = 248 \]

\[ Ar = \frac{248}{5} = 23 \]

Therefore, the atomic mass of M is 23.
21. 10.0g of hydrated iron (II) sulphate \( (\text{Fe}_2\text{SO}_4, n\text{H}_2\text{O}) \) gave 4.53g of water after strong heating. Find the value of \( n \).

<table>
<thead>
<tr>
<th></th>
<th>\text{FeSO}_4</th>
<th>\text{H}_2\text{O}</th>
</tr>
</thead>
<tbody>
<tr>
<td>Relative mass (g)</td>
<td>5.47</td>
<td>4.53</td>
</tr>
<tr>
<td>Relative number of mass</td>
<td>( 5.47 / 152 = 0.036 )</td>
<td>( 4.53 / 18 = 0.252 )</td>
</tr>
<tr>
<td></td>
<td>0.036 / 0.036 = 1</td>
<td>0.252 / 0.036 = 7</td>
</tr>
</tbody>
</table>

Therefore, \( n = 7 \)

22. A metal (\( M \)) forms two chlorides (A and B), which contain 55.9% and 65.5% by mass of chloride, respectively. The empirical formula of A is found to be \( \text{MCl}_2 \). Determine the empirical formula of B (without having to find the atomic mass of \( M \)).

If the mass of \( M \) in chloride B is the same as that in chloride A (i.e. 44.1g), the mass of chlorine in chloride B will be equal (by proportion) to
\[
65.5 \times 44.1 / 34.5 = 83.7g
\]

Therefore, for the same mass (hence the same number of moles) of \( M \):
- the mass of chlorine in chloride B / the mass of chlorine in chloride A
- = number of moles of chlorine in chloride B / number of moles of chlorine in chloride A
- = 83.7 / 55.9 = 1.5

Since the formula of chloride A is given to be \( \text{MCl}_2 \), the formula of chloride B should be \( \text{MCl}_3 \).
More activities and solutions

Exercises 1

Hydrogen peroxide decomposes at room temperature according to the following balanced chemical equation:

\[ 2\text{H}_2\text{O}_2(aq) \rightarrow \text{O}_2(g) + 2\text{H}_2\text{O}(l) \]

1.1 What does (aq) represent in the equation above? (1)

1.2 Identify the type of reaction above. Choose between PRECIPITATION and REDOX. Give a reason for your answer. (2)

1.3 Is the reaction an example of a physical change or a chemical change? (1)

1.4 Define the term one mole of a substance. (2)

1.5 If 4 moles of hydrogen peroxide decompose, calculate the volume of gas formed at STP. (4)

1.6 Calculate the number of oxygen atoms in H\textsubscript{2}O\textsubscript{2}, if 17g of H\textsubscript{2}O2 decomposes. (4)

Solutions

1.1 A solution in which the solvent is water. (1)

1.2 Redox. There is a change in oxidation number of atoms involved this reaction (i.e. hydrogen and oxygen). (2)

1.3 Chemical change. (1)

1.4 One mole is the amount of substance that has the same number of particles as there are atoms in 12g carbon-12. (2)

1.5 \( n(\text{H}_2\text{O}_2) : n(\text{O}_2) = 2:1 \)
   Therefore: \( n(\text{O}_2) = 2 \text{ mol} \)
   \( n(\text{O}_2) = \frac{2}{2} = 1 \)
   \( V = 44.8 \text{ dm}^3 \) (4)

1.6 \( n(\text{H}_2\text{O}_2) = \frac{17g}{34g} = 0.5 \text{ mol} \)
   \( n(\text{O}_2) = (0.5)(2) = 1 \text{ mol} \)
   \( N = \text{ atoms} \) (4)
\[
n(\text{H}_2\text{O}_2) : n(\text{O}_2) = 2:1
\]

Therefore: \(n(\text{O}_2) = 2\) mol

\[
n(\text{O}_2) = \frac{2}{2} = 1
\]

\[
V = 44.8\, \text{dm}^3
\]
Exercise 2

2.1 The empirical formula of a certain compound is to be determined. On analysis of a sample of the compound it was found to contain: 40% C; 6.6% H; 53.3% O.

2.1.1 Define the term empirical formula. (2)

2.1.2 Determine the empirical formula of the compound. Show ALL calculations. (5)

2.1.3 If the molecular mass of the compound is 60 g·mol⁻¹, calculate the molecular formula of the compound. (3)

2.2 The molar mass of hydrated sodium carbonate is found to be 268 g·mol⁻¹. The formula of the hydrated sodium carbonate is Na₂CO₃·xH₂O.

Calculate the number of moles of water of crystallisation (x) in the compound. (4)

Solutions

2.1.1 The empirical formula is the simplest whole number ratio of atoms in a compound. (2)

2.1.2 If 100g of the compound is available, then:

<table>
<thead>
<tr>
<th>Elements (atoms)</th>
<th>Mass divided by molar mass</th>
<th>Divide ALL the answers in the middle column by the lowest molar mass of all</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon (C)</td>
<td>40,0 (\div) 12 = 3,3 [3,3]</td>
<td>3,3 [3,3] = 1</td>
</tr>
<tr>
<td>Hydrogen (H)</td>
<td>6,6 (\div) 1 = 6,6</td>
<td>6,6 [3,3] = 2</td>
</tr>
<tr>
<td>Oxygen (O)</td>
<td>53,3 (\div) 16 = 3,3 [3,3]</td>
<td>3,3 [3,3] = 1</td>
</tr>
</tbody>
</table>

C = 1; H = 2; O = 1

The empirical formula = CH₂O (5)

2.1.3 \(M\) (CH₂O) = 30 g·mol⁻¹

\[\frac{M}{(CH₂O)} = 2\]

Molecular formula: multiply each of the number of atoms in the empirical formula by 2.

Molecular formula = C₂H₄O₂ (3)

2.2 \(M\) (Na₂CO₃) = 106 g·mol⁻¹

\(M\) (xH₂O) = 268 – 106 = 162g·mol⁻¹

\(n\) (H₂O) = \(n\) (H₂O) = 9 mol (4)
Exercise 3

Calcium carbonate (CaCO$_3$) reacts with dilute hydrochloric acid (HCl) according to the following balanced equation:

$$\text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)$$

3.1 The above reaction is an example of an acid-base reaction. Define the term acid-base reaction. (2)

The graph below shows the relationship between the volume of carbon dioxide gas, CO$_2$(g) formed and the mass of PURE calcium carbonate.

![Graph showing the relationship between volume of CO$_2$ and mass of CaCO$_3$](image)

3.2 From the graph, determine the volume of CO$_2$(g) produced when 0.072g of PURE CaCO$_3$(s) reacts. (1)

3.3 A certain antacid tablet, with a mass of 0.25g, contains mainly calcium carbonate, which reacts with dilute hydrochloric acid in the stomach to produce carbon dioxide gas. The concentration of hydrochloric acid in the stomach is 0.1 mol·dm$^{-3}$

3.3.1 Define the term concentration of a solution. (2)

3.3.2 It is found that 25 cm$^3$ of CO$_2$(g) is formed when one antacid tablet completely reacts. Use the information in the graph to calculate the percentage CaCO$_3$(s) in one antacid tablet. (3)

3.3.3 Calculate the volume of hydrochloric acid that will be neutralised by ONE antacid tablet. (5)

Solutions

3.1 A reaction in which a proton/ hydrogen ion/ H$^+$ is transferred from one reactant to another. (2)
3.2 18 cm³

3.3.1 The number of moles of solute per cubic decimetre/litre of solution.

3.3.2 \(\%\text{CaCO}_3 = x \times 100\)
\(\%\text{CaCO}_3 = 40\%\)

3.3.3 \(n(\text{CaCO}_3) =\)
\(n(\text{CaCO}_3) =\)
\(n(\text{CaCO}_3) = 0.001\)
\(n(\text{HCl}) = 2 \times n(\text{CaCO}_3) = 2 \times 10^{-3} \text{ mol}\)

Volume of acid
\(c = 0.1 = V = 0.02 \text{ dm}^3\)
Exercise 4

4.1 Study the balanced chemical equation of the reaction between sodium carbonate \((\text{Na}_2\text{CO}_3)\) and hydrochloric acid \((\text{HCl})\) and answer the questions that follow.

\[
\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}
\]

Identify the type of reaction above. Choose between REDOX and GAS FORMING. \(\text{(1)}\)

4.2 In a reaction, 10.6g of sodium carbonate reacts completely with excess hydrochloric acid.

4.2.1 Calculate the molar mass of sodium carbonate. \(\text{(2)}\)

4.2.2 Calculate the initial number of moles of sodium carbonate. \(\text{(2)}\)

4.2.3 Calculate the mass of \(\text{CO}_2\) produced during this reaction. \(\text{(4)}\)

4.2.4 Calculate the mass of sodium chloride produced if 4.87 dm\(^3\) of carbon dioxide is produced at STP. \(\text{(6)}\)

4.3 14.2g of a sample of hydrated sodium carbonate \((\text{Na}_2\text{CO}_3\cdot\text{xH}_2\text{O})\) was heated until no further change in mass was recorded. On heating, all the water of crystallisation evaporated as follows:

\[
\text{Na}_2\text{CO}_3\cdot\text{xH}_2\text{O} \rightarrow \text{Na}_2\text{CO}_3 + \text{xH}_2\text{O}
\]

If 4 moles of hydrogen peroxide decompose, calculate the volume of gas formed at STP.

Calculate the number of oxygen atoms in \(\text{H}_2\text{O}\) if 17g of \(\text{H}_2\text{O}_2\) decomposes. \(\text{(5)}\)

[20]

4.1 Gas forming \(\text{(1)}\)

4.2.1 \(\text{M(}\text{Na}_2\text{CO}_3) = 2(23) + 12 + 3(16)\)

\[
\text{M(}\text{Na}_2\text{CO}_3) = 106\text{g.mol}^{-1}
\]

4.2.2 \(\text{n(Na}_2\text{CO}_3) = \text{n(Na}_2\text{CO}_3) = \text{n(Na}_2\text{CO}_3) = 0.1\text{ mol}\)

4.2.3 \(\text{n(Na}_2\text{CO}_3) : \text{n(CO}_2) = 1 : 1\)

Thus: \(\text{n(CO}_2) = 0.1\text{mol}\)

\[
\text{n(CO}_2) = 0.1 = m = 4.4\text{ g}
\]

4.2.4 \(\text{n(CO}_2) = \)
\( n(\text{CO}_2) = n(\text{CO}_3) = 0,217 \text{ mol} \)

\[
n(\text{CO}_2) : n(\text{NaCl}) = 1 : 2 \\
n(\text{NaCl}) = 0,434 \text{ mol} \\
m = 25,16 \text{ g} \quad (6)
\]

4.3 Mass of \( \text{H}_2\text{O} = 14,2 - 5,3 \)
\[= 8,9 \text{ g} \]

\[
n(\text{Na}_2\text{CO}_3) = n(\text{H}_2\text{O}) = n(\text{H}_2\text{O}) = 0,05 \text{ mol} \\
\text{Na}_2\text{CO}_3 : \text{H}_2\text{O} = 0,05 : 0,5 \\
0,05 : 0,05 \\
1 : 10
\]
Thus, \( X = 10 \quad (5) \)
Exercise 5

The reaction between sodium and water is represented by the following balanced chemical equation:

\[ 2\text{Na}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{NaOH}(aq) + \text{H}_2(g) \]

5.1 Write down the value of the temperature and the pressure at STP. (2)

5.2 Calculate the following:

5.2.1 Mass (in grams) of hydrogen gas produced. (5)

5.2.2 Volume (in dm\(^3\)) of hydrogen gas produced at STP. (3)

5.2.3 Mass (in grams) of NaOH produced. (4)

5.2.4 Concentration of the sodium hydroxide solution. (3)

5.1 Temperature : 0°C or/and 273 K
Pressure : 101,3 kPa or 1 atm (2)

5.2.1 \( n(\text{Na}) = \frac{m}{M} \)
\( n(\text{Na}) = \frac{10}{23} \)
\( n(\text{Na}) = 0,43 \text{ mol} \)
Na : H\(_2\)
2 : 1
Thus, 0,22 mol H\(_2\) produced

5.2.2 \( n(\text{H}_2) = \frac{V}{V_m} \)
\( 0,22 = \frac{V}{22,4} \)
\( V = 4,93 \text{ dm}^3 \) (3)

5.2.3 \( n(\text{Na}) : n(\text{NaOH}) \)
2 : 2
Thus, mol NaOH = 0,43 mol

\( n(\text{NaOH}) = \frac{m}{M} \)
\( 0,43 = \frac{m}{40} \)
\( m = 17,2 \text{ g} \) (4)
5.2.4 \( c = \frac{n}{V} \)

\[
c = \frac{0.43}{2}
\]

\[c = 0.22 \text{ mol.dm}^{-3}\]  

(3)  

[17]
Exercise 6

Hydrogen H2(g) and nitrogen N2(g) react to form ammonia NH3(g). The reaction that takes place is represented by the following equation:

\[ \text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \]

6.1.1 Define the term 1 mole. (2)

6.1.2 How many moles of ammonia will be produced from 1 mole of hydrogen gas? (1)

6.1.3 Initially 10 cm³ of nitrogen and 24 cm³ of hydrogen are mixed in a container. The temperature and pressure remain constant.

Calculate the volume of gas that will remain in the container after the reaction is completed. (4)

In another experiment, 80g of hydrogen gas reacts with nitrogen gas to form ammonia.

Calculate the:

6.1.4 The number of moles of hydrogen gas that reacted. (2)

6.1.5 The volume of nitrogen gas used at STP. (2)

6.2 When 207g of lead, Pb, combines with oxygen, 239g of a certain oxide of lead is formed. Use a calculation to determine the formula of this oxide of lead. (5)

6.1.1 One mole is the amount of substance that has the same number of particles as there are atoms in 12g carbon-12. (2)

6.1.2 0,67 mol mol

6.1.3 \( \text{Vol}(\text{N}_2) : \text{V(H}_2) : \text{V(NH}_3) = 1 : 3 : 2 \)

\( \text{Vol(} \text{N}_2 \text{ reacted} = \text{ V(H}_2) \)

\( \text{Vol(} \text{N}_2 \text{ reacted} = (24) \)

\( \text{Vol(} \text{N}_2 \text{ reacted} = 8 \text{ dm}^3 \)

Volume \( \text{N}_2 \) remains = 10 – 8 = 2 dm³

The volume of gas that remains = 2 + 16

The volume of gas that remains = 18 dm³ gas

6.1.4 \( n = \frac{m}{M} \)

\( n = \frac{80}{2} \)

\( n = 40 \text{ mol} \)
6.1.5 \( \text{Vol}(N_2) = \frac{1}{3} \times (40) \times 22.4 \)
\[ \text{Vol}(N_2) = 2\,986.7\,\text{dm}^3 \]

6.2 \( m(O_2) = 239 - 207 = 32\,\text{g} \)

\[ n_{\text{Pb}} = \frac{m}{M} \]

\[ n_{\text{Pb}} = \frac{207}{207} \]
\[ n_{\text{Pb}} = 1\,\text{mol} \]

\[ n_{O_2} = \frac{m}{M} \]

\[ n_{O_2} = \frac{32}{M} \]
\[ n_o = 2\,\text{mol} \]

\( n(\text{Pb}) : n(O) = 1 : 2 \)
\[ \text{PbO}_2 \] (5)

[16]
Exercise 7

Nitric acid \((HNO_3)\) is a strong acid and is an important acid used in industry.

7.1 Give a reason why nitric acid is classified as a strong acid. (1)

7.2 Write down the NAME or FORMULA of the conjugate base of nitric acid. (1)

7.3 Calculate the pH of a 0,3 mol·dm\(^{-3}\) nitric acid solution. (3)

A laboratory technician wants to determine the percentage purity of magnesium oxide. He dissolves a 4,5 g sample of the magnesium oxide in 100 cm\(^3\) hydrochloric acid of concentration 2 mol·dm\(^{-3}\).

7.3.1 Calculate the number of moles of hydrochloric acid added to the magnesium oxide. (3)

He then uses the apparatus below to titrate the EXCESS hydrochloric acid in the above solution against a hydroxide solution.

![Diagram](image)

7.3.2 Write down the name of the apparatus Q in the above diagram. (1)

7.3.3 The following indicators are available for the titration:

<table>
<thead>
<tr>
<th>INDICATOR</th>
<th>pH RANGE</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>3,1 – 4,4</td>
</tr>
<tr>
<td>B</td>
<td>6,0 – 7,6</td>
</tr>
<tr>
<td>C</td>
<td>8,3 – 10,0</td>
</tr>
</tbody>
</table>

Which ONE of the above indicators (A, B or C) is most suitable to indicate the exact endpoint in this titration? Give a reason for your answer. (3)

7.3.4 During titration, the technician uses distilled water to wash any sodium hydroxide that was spilt against the side of the Erlenmeyer flask into the solution. Give a reason why the addition of distilled water to the Erlenmeyer flask will not influence the results. (1)

7.3.5 At the endpoint of the titration, the technician finds that 21 cm\(^3\) of a 0,2 mol·dm\(^{-3}\) sodium hydroxide solution has neutralised the EXCESS hydrochloric acid. Calculate the number of moles of hydrochloric acid that are in excess. (3)
7.3.6 The balanced equation for the reaction between hydrochloric acid and magnesium oxide is:

\[ \text{MgO} \text{(s)} + 2\text{HCl} \text{(aq)} \rightarrow \text{MgCl}_2 \text{(aq)} + 2\text{H}_2 \text{O(ℓ)} \]

Calculate the percentage purity of the magnesium oxide. Assume that only the magnesium oxide in the 4.5 g sample reacted with the acid. (5)
Exercise 8

8.1 Ionises / dissociates completely in water.  

8.2 $\text{NO}_3^-$ / Nitrate ion  

8.3 $\text{pH} = -\log[H_3O^+] = - \log (0,3)$  
\[= 0,52\]

8.2.1 $c = \frac{n}{V}$  
\[2 = \frac{n}{0,1}\]  
$n(\text{HCl}) = 0,2 \text{ mol}$

8.2.2 Burette

8.2.3 B  
Titration of a strong acid and a strong base.

8.2.4 The number of moles of acid in the flask remains constant.

8.2.5 $c = \frac{n}{V}$  
\[0,2 = \frac{n}{0,021}\]  
$n = 4,2 \times 10^{-3} \text{ mol}$  

$n(\text{HCl in excess}) = n(\text{NaOH})$  
\[= 4,2 \times 10^{-3} \text{ mol}\]

$n(\text{HCl in excess})$:  
$0,2 - 4,2 \times 10^{-3} = 0,196 \text{ mol}$

$n(\text{MgO reacted})$:  
$= (1/2) n(\text{HCl}) = n(\text{MgO})$  
$= (1/2)(0,196)$  
$= 9,8 \times 10^{-2} \text{ mol}$

$m(\text{MgO}) = nM$  
$= (0,098)(40)$  
$= 3,92 \text{ g}$

\[\% \text{ purity} = \frac{3,92}{4,5} \times 100 = 87,11\%\]
Exercise 9

9.1 Sulphuric acid is a diprotic acid.
   9.1.1 Define an acid in terms of the Lowry-Brønsted theory. (2)
   9.1.2 Give a reason why sulphuric acid is referred to as a diprotic acid. (1)

9.2 The hydrogen carbonate ion can act as an acid and a base. It reacts with water according to the following balanced equation:

\[ \text{HCO}_3^{-} + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3 + \text{OH}^- \]

9.2.1 Write down ONE word for the underlined phrase. (1)

9.2.2 \text{HCO}_3^{-} acts as a base in the above reaction. Write down the formula of the conjugate acid of \text{HCO}_3^{-}. (1)

9.3 A learner accidentally spills some sulphuric acid of concentration 6 mol·dm\(^{-3}\) from a flask on the laboratory bench. Her teacher tells her to neutralise the spilled acid by sprinkling sodium hydrogen carbonate powder onto it. The reaction that takes place is indicated below. (Assume that the \text{H}_2\text{SO}_4 ionises completely.)

\[ \text{H}_2\text{SO}_4 + 2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O} + 2\text{CO}_2 \]

The fizzing, due to the formation of carbon dioxide, stops after the learner has added 27 g sodium hydrogen carbonate to the spilled acid.

9.3.1 Calculate the volume of sulphuric acid that was spilt. Assume that all the sodium hydrogen carbonate reacts with all the spilt acid. (6)

The learner now dilutes some of the 6 mol·dm\(^{-3}\) sulphuric acid solution in the flask to 0.1 mol·dm\(^{-3}\).

9.3.2 Calculate the volume of the 6 mol·dm\(^{-3}\) sulphuric acid solution needed to prepare 1 dm\(^{3}\) of the dilute acid. (2)

During titration, 25 cm\(^{3}\) of the 0.1 mol·dm\(^{-3}\) sulphuric acid solution is added to an Erlenmeyer flask and titrated with a 0.1 mol·dm\(^{-3}\) sodium hydroxide solution.

9.3.3 The learner uses bromothymol blue as an indicator. What is the purpose of this indicator? (1)

9.3.4 Calculate the pH of the solution in the flask after the addition of 30 cm\(^{3}\) of sodium hydroxide. At this point, the endpoint of the titration has not yet been reached. (8)
Solutions

9.1
9.1.1 An acid is a proton (H+ ion) donor. (2)

9.1.2 It ionises to form 2 protons/2 moles of H+ ions.
   OR
   It donates 2 H+ ions per H2SO4 molecule. (1)

9.2
9.2.1 Amphiprotic (substance)/ Ampholyte (1)

9.2.2 H2SO4(aq) (1)

9.3
9.3.1 \( n(\text{NaHCO}_3) = \frac{m}{M} = \frac{27}{84} = \frac{0.321485714 \text{ mol}}{0.32 \text{ mol}} \)

\( n(\text{H}_2\text{SO}_4) = n(\text{NaHCO}_3) = \frac{1}{2} (0.32) = 0.16 \text{ mol} \)

\( c = \frac{n}{V} = 0.16/V \)

\( V = 0.027 \text{ dm}^3 \) (6)

9.3.2 \( n_a (\text{initial}) = n_a (\text{final}) \)

\( c_a V_a (\text{initial}) = c_a V_a (\text{final}) \)

\( (6) \frac{8}{V} = (01)(1) \)

\( V_a = 0.017 \text{ dm}^3 \) (2)

9.3.3 It shows the endpoint of titration. It shows when neutralisation occurs. (1)
9.3.4 - Marking criteria:
- Substitute initial [acid] and volume.
- Substitute reacted [base] and volume.
- Use the ratio 1:2
- Initial mole acid – mole acid
- Substitute volume acid + volume base
- pH formula
- Substitute 2 × c_a in pH formula
- Final answer: 1,44

\[ n_a \text{ (initial)} = c_a V_a \]
\[ = (0,1)(25 \times 10^{-3}) \]
\[ = 2,5 \times 10^{-3} \text{ mol} \]

\[ n_b \text{ (reacted)} = c_b V_b \]
\[ = (0,1)(30 \times 10^{-3}) \]
\[ = 3 \times 10^{-3} \text{ mol} \]

\[ \frac{n_a}{n_b} = 1/2 \]
\[ n_a \text{ (neutralised)} = \frac{1}{2} n_b = \frac{1}{2} (3 \times 10^{-3}) \]
\[ = 1,5 \times 10^{-3} \text{ mol} \]

\[ n_a \text{ (left)} = n_a \text{ (initial)} - n_a \text{ (neutralised)} \]
\[ = 2,5 \times 10^{-3} - 1,5 \times 10^{-3} \]
\[ = 1 \times 10^{-3} \text{ mol} \]

\[ c_a = \frac{n}{V} \]
\[ = (1 \times 10^{-3})(25 \times 10^{-3}) + (30 \times 10^{-3}) \]
\[ = 0,018 \text{ mol.dm}^{-3} \]

\[ \text{pH} = -\log [H_3 O^+] \]
\[ = -\log (2 \times 0,018) \]
\[ = 1,44 \]
Exercise 10

Dilute acids are indicated in the table below. These react with EXCESS zinc in each of the three experiments to produce hydrogen gas. The zinc is completely covered with the acid in each experiment.

<table>
<thead>
<tr>
<th>EXPERIMENT</th>
<th>DILUTE ACID</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>100 cm³ of 0.1 mol·dm⁻³ H₂SO₄</td>
</tr>
<tr>
<td>2</td>
<td>50 cm³ of 0.2 mol·dm⁻³ H₂SO₄</td>
</tr>
<tr>
<td>3</td>
<td>100 cm³ of 0.1 mol·dm⁻³ HCl</td>
</tr>
</tbody>
</table>

The volume of hydrogen gas produced is measured in each experiment.

10.1 Name TWO of the essential apparatus pieces needed to determine the rate of hydrogen production.  
(2)

The graph below was obtained for Experiment 1.

Use this graph and answer the questions that follow.

10.1 At which time (t₁, t₂ or t₃) is:
   10.1.1 The reaction rate the highest?  
   10.1.2 The mass of the zinc present in the flask the smallest?  
(1)(1)

10.2 At which time interval does the largest volume of hydrogen gas form per second?  
Choose from: between t₁ and t₂ OR between t₂ and t₃.  
(1)

10.3 Re-draw the graph for Experiment 1 in your ANSWER BOOK.  
On the same set of axes, sketch the graphs that will be obtained for 
Experiment 2 and Experiment 3.  
Clearly label the three graphs as EXPERIMENT 1, EXPERIMENT 2 and 
EXPERIMENT 3.  
(4)
10.4 The initial mass of zinc used in each experiment is 0.8 g. The balanced equation for the reaction in Experiment 3 is:

\[ \text{Zn}^{(s)} + 2\text{HCl}^{(aq)} \rightarrow \text{ZnCl}_2^{(aq)} + \text{H}_2^{(g)} \]

10.4.1 Calculate the mass of zinc used present in the flask after completion of the reaction in Experiment 3. (5)

10.5 How will the mass of the zinc present in the flask after completion of the reaction in Experiment 2 compare to the answer to QUESTION 5.5.1? Write down one of the following only: LARGER THAN; SMALLER THAN; EQUAL TO. (1)
Solutions

10.1. Time: Stopwatch
Volume: Gas syringe / Burette / Measuring cylinder / Chemical balance / Erlenmeyer flask / Graduated flask

10.2.
10.2.1. \( t_1 \)
10.2.2. \( t_3 \)

10.3. Between \( t_1 \) and \( t_2 \)

10.4.

![Graph showing volume over time for Exps. 1, 2, and 3.]

Marking criteria

<table>
<thead>
<tr>
<th>Exp. 2</th>
<th>Exp. 3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial gradient higher than that of Exp. 1</td>
<td></td>
</tr>
<tr>
<td>Curve reaches same constant volume as for Exp. 1, but earlier.</td>
<td></td>
</tr>
<tr>
<td>Initial gradient lower than that of Exp. 1.</td>
<td></td>
</tr>
<tr>
<td>Curve reaches a smaller constant volume, as for Exp. 1, at a later stage.</td>
<td></td>
</tr>
</tbody>
</table>

\begin{align*}
10.5.1 \quad c(HC) &= \frac{n}{V} \\
&= \frac{n}{100 \times 10^{-3}} \\
&= 0.01 \text{ mol} \\

n(Zn) &= \frac{1}{2} n(HC) = 0.005 \text{ mol} \\

n(Zn) &= \frac{m(Zn_{\text{reacted}})}{M} \\
&= 0.005 \times \frac{m(Zn_{\text{reacted}})}{65} \\
m(Zn_{\text{reacted}}) &= 0.325 \text{ g} \\
m(Zn_{1}) &= 0.8 - 0.325 = 0.475 \text{ g}
\end{align*}

10.5.2 Smaller than

[15]
Exercise 11

An unknown gas, \( X_{2(g)} \), is sealed in a container and allowed to form \( X_{3(g)} \) at 300°C. The reaction reaches equilibrium according to the following balanced equation:

\[
3X_2 \rightleftharpoons 2X_3
\]

11.1. How will the rate of formation of \( X_{3(g)} \) compare to the rate of formation of \( X_{2(g)} \) at equilibrium? Write down only one of the following: HIGHER THAN; LOWER THAN; EQUAL TO. (1)

The reaction mixture is analysed at regular time intervals. The results obtained are shown in the table below.

<table>
<thead>
<tr>
<th>TIME (s)</th>
<th>([ X_2 ]) (mol ∙ dm(^{-3}))</th>
<th>([ X_3 ]) (mol ∙ dm(^{-3}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>0,4</td>
<td>0</td>
</tr>
<tr>
<td>2</td>
<td>0,22</td>
<td>0,120</td>
</tr>
<tr>
<td>4</td>
<td>0,08</td>
<td>0,213</td>
</tr>
<tr>
<td>6</td>
<td>0,06</td>
<td>0,226</td>
</tr>
<tr>
<td>8</td>
<td>0,06</td>
<td>0,226</td>
</tr>
<tr>
<td>10</td>
<td>0,06</td>
<td>0,226</td>
</tr>
</tbody>
</table>

11.2. Calculate the equilibrium constant, \( K_c \), for this reaction at 300°C. (4)

11.3. More \( X_{3(g)} \) is now added to the container.

11.3.1. How will this change affect the amount of \( X_{2(g)} \)? Write down only one of the following: INCREASES; DECREASES; REMAINS THE SAME. (1)

11.3.2. Use Le Chatelier’s principle to explain your answer to QUESTION 5.3.1. (2)

The curves on the set of axes below (not drawn to scale) were obtained from the results in the table above.

11.4. How does the rate of the forward reaction compare to that of the reverse reaction at \( t_1 \)? Write down only one of the following: HIGHER THAN; LOWER THAN; EQUAL TO. (1)

The reaction is now repeated at a temperature of 400°C. The curves indicated by the dotted lines were obtained at this temperature.
11.5. Is the forward reaction EXOTHERMIC or ENDOTHERMIC? Explain in full how you arrived at your answer.

The Maxwell-Boltzmann distribution curve below represents the number of particles against kinetic energy at 300°C.

11.6. Re-draw this curve in your ANSWER BOOK. On the same set of axes, sketch the curve that will be obtained at 400°C. Clearly label the curves as 300°C and 400°C, respectively.
Exercise 12

12.1 Equal to

\[ K_c = \frac{[X_3]^2}{[X_2]^2} \]

(1)

12.2

\[ K_c = \frac{(0.226)^2}{(0.06)^2} \]

= 206.46

(4)

12.3

6.3.1 Increases

- The increase in [X3] is opposed. The change is opposed.
- The reverse reaction is favoured. X3 is used. [X3] decreases

(2)

12.4 Higher than

(1)

12.5 Exothermic

- [X3] decreases and [X2] increases.
- Kc decreases if temperature increases.
- The decrease in temperature favoured the forward reaction.

(4)

12.6

Marking criteria

The peak of the curve at 400°C is lower than at 300°C and it shifted to the right.
The curve at 400°C has a larger area at the higher E_k.

(2)
Exercise 13

Initially, 2.2 g of pure CO\(_2\)(g) is sealed in an empty 5 dm\(^3\) container at 900°C.

13.1 Calculate the initial concentration of CO\(_2\)(g). (4)

13.2 Give a reason why equilibrium will not be established. (1)

CaCO\(_3\)(g) is now added to the 2.2 g CO\(_2\)(g) in the container. After a while, equilibrium is established at 900°C according to the following balanced equation:

\[
\text{CaCO}_3(s) \rightleftharpoons \text{CaO}(s) + \text{CO}_2(g)
\]

The equilibrium constant for this reaction at 900°C is 0.0108.

13.3 Give a reason why this reaction will only reach equilibrium in a SEALED container. (1)

13.4 Calculate the minimum mass of CaCO\(_3\)(s) that must be added to the container to achieve equilibrium. (7)

13.5 How will EACH of the following changes affect the amount of CO\(_2\)(g)?
Write down only one of the following: INCREASES; DECREASES; REMAINS THE SAME.
6.5.1 More CaCO\(_3\)(s) is added at 900°C. (1)
6.5.2 The pressure is increased. (1)

13.6 It is found that the equilibrium constant (K\(_{eq}\)) for this reaction is \(2.6 \times 10^{-6}\) at 727°C. Is the reaction EXOTHERMIC or ENDOTHERMIC? Explain in full how you arrived at the answer. (4)
Solutions

6.1 \[ M (\text{CO}_2) = 12 + 2(16) = 44 \text{ g mol}^{-1} \]
\[
 c = \frac{m}{MV} = \frac{2,2}{(44)(5)} \\
= 0,01 \text{ mol dm}^{-3} \tag{4}
\]

6.2 If only \text{CO}_2 is present, the reverse reaction cannot take place. \tag{1}

6.3 \text{CO}_2 is a gas that will escape if the container is not sealed. \tag{1}

6.3 \[ K_c = [\text{CO}_2] \]
\[ = 0,0108 \]
\[ [\text{CO}_2] = 0,0108 \text{ mol dm}^{-3} \]

<table>
<thead>
<tr>
<th></th>
<th>CaCO$_3$</th>
<th>CaO</th>
<th>CO$_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ratio</td>
<td>1</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>Initial quantity (mol)</td>
<td>0,004</td>
<td>0</td>
<td>0,05</td>
</tr>
<tr>
<td>Change (mol)</td>
<td>Used</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>0,004</td>
<td></td>
</tr>
<tr>
<td>Form</td>
<td></td>
<td>0,004</td>
<td></td>
</tr>
<tr>
<td>Quantity at equilibrium (mol)</td>
<td></td>
<td>0,054</td>
<td></td>
</tr>
<tr>
<td>Equilibrium concentration (mol dm$^{-3}$)</td>
<td></td>
<td>0,0108</td>
<td></td>
</tr>
</tbody>
</table>

\[
m(\text{CaCO}_3) = nM \\
= (0,004)(100) \\
= 0,4 \text{ g} \tag{7}
\]

6.5

8.5.1 Remains the same. \tag{1}
8.5.2 Decreases. \tag{1}

6.6 Endothermic
- \( K_c \) increases with an increase in temperature.
- The increase in temperature favours the forward reaction.
- The increase in temperature favours the endothermic reaction. \tag{4}
Exercise 14

Methanol and hydrochloric acid react according to the following balanced equation:

\[ \text{CH}_3\text{OH}^{(aq)} + \text{HCl}^{(aq)} \rightarrow \text{CH}_3\text{Cl}^{(aq)} + \text{H}_2\text{O}(l) \]

14.1 State TWO factors that can INCREASE the rate of this reaction. (2)

14.2 Define the term reaction rate. (2)

14.3 The rate of the reaction between methanol and hydrochloric acid is investigated. The concentration of HCl\(^{(aq)}\) was measured at different time intervals. The following results were obtained:

<table>
<thead>
<tr>
<th>TIME (MINUTES)</th>
<th>HCl CONCENTRATION (mol·dm(^{-3}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>1,90</td>
</tr>
<tr>
<td>15</td>
<td>1,45</td>
</tr>
<tr>
<td>55</td>
<td>1,10</td>
</tr>
<tr>
<td>100</td>
<td>0,85</td>
</tr>
<tr>
<td>215</td>
<td>0,60</td>
</tr>
</tbody>
</table>

14.3.1 Calculate the average rate, in (mol·dm\(^{-3}\))·min\(^{-1}\) during the first 15 minutes. (3)

14.3.2 Use the data in the table to draw a graph of concentration versus time. NOTE: The graph is not a straight line. (3)

14.3.2 From the graph, determine the concentration of HCl\(^{(aq)}\) at the 40th minute. (1)

14.3.2 Use the collision theory to explain why the reaction rate decreases with time. Assume that the temperature remains constant. (3)

14.3.2 Calculate the mass of CH\(_3\)Cl\(^{(aq)}\) in the flask at the 215\(^{th}\) minute. The volume of the reagents remains 60 cm\(^3\) during the reaction. (5)
Exercise 15

15.1 Increase temperature.
   Increase the concentration of acid.
   Add a catalyst.  
\[ \text{Increase temperature.} \] 
\[ \text{Increase concentration of acid.} \] 
\[ \text{Add catalyst.} \] 
\[ \text{(2)} \]

15.2 Change in concentration of products / reactants per unit time.
   OR
   Rate of change in concentration. 
\[ \text{(2)} \]

15.2.3 \[ \text{Average rate}=\frac{-\Delta c}{\Delta t} \]
\[ =\frac{(1.45 - 1.9)}{(15 - 0)} \]
\[ =0.03 \text{ (mol} \cdot \text{dm}^{-3}) \cdot \text{min}^{-1} \] 
\[ \text{(3)} \]

Marking criteria

| Four points correctly plotted. |
| Curve drawn as shown. |

15.2.3 1,2 mol· dm\(^{-3}\)  
\[ \text{(1)} \]

15.2.4
- The concentration of reactants decreases.
- Less particles per unit volume.
- Less effective collisions per unit time.  
\[ \text{(3)} \]
15.2.5 $\Delta c(\text{HC}\ell) = 0.6-1.9$

\[ c(\text{HC}\ell) = n/V \]

\[ 1.3 = \frac{n}{60 \times 10^{-3}} \]

\[ n(\text{HC}\ell) = 0.078 \text{ mol} \]

\[ n(\text{CH}_3\text{Cl}) = n(\text{HC}\ell) = 0.078 \text{ mol} \]

\[ n(\text{CH}_3\text{Cl}) = \frac{m}{M} \]

\[ 0.078 = \frac{m}{(50.5)} \]

\[ m(\text{CH}_3\text{Cl}) = 3.94 \text{ g} \]