PHYSICAL SCIENCE_PAPER 2
STOICHIOMETRY
GRADE 12

Read to Lead

A Reading Nation is a Leading Nation

NDP 2030

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.

Contents

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 the 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!

Atom (carbon-12)
Nucleus
+ Proton
○ Neutron
● Electron

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.
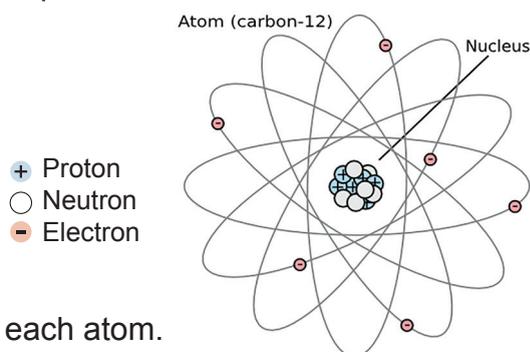
4.1 The Mole Concept

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 6.022×10^{23} items. So, there are 6.022×10^{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 6.022×10^{23} atoms of sodium? _____
7. How **many moles** are 12.04×10^{23} atoms of carbon? _____
8. How **many moles** are 18.06×10^{23} atoms of sodium? _____
9. How many moles are 60.22×10^{23} atoms of sodium? _____
10. How many moles are 6.022×10^{23} molecules of water? _____
11. How many moles are 12.04×10^{23} molecules of water? _____
12. How many moles are 30.10×10^{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?

Limiting Reactants

8 car bodies   48 tires   8 cars   plus 16 tires excess

$$CB \quad + \quad 4\,T \quad \longrightarrow \quad CT_4$$
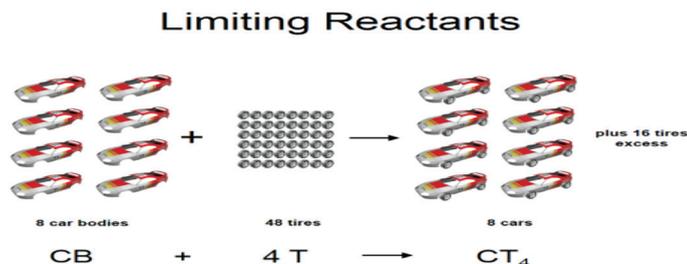
See how many tyres you are able to use in this example. The remaining tyres are in excess. The cars are said to be limiting.

Which of the two reactants is in excess and which is the limiting reagent in the following diagrams?

8 Carbon atoms   24 Hydrogen molecules   8 Methane molecules
C + $2 H_2$ $\longrightarrow$ CH_4

8 molecules of hydrogen are in excess

Before and After Reaction 2

$$N_2 + 3H_2 \longrightarrow 2 NH_3$$

excess   limiting

∞ = H_2
∞ = N_2
= NH_3

Before the reaction   After the reaction
LIMITING REACTANT DETERMINES AMOUNT OF PRODUCT
Zumdahl, Zumdahl, DeCoste, World of Chemistry 2002, page 270
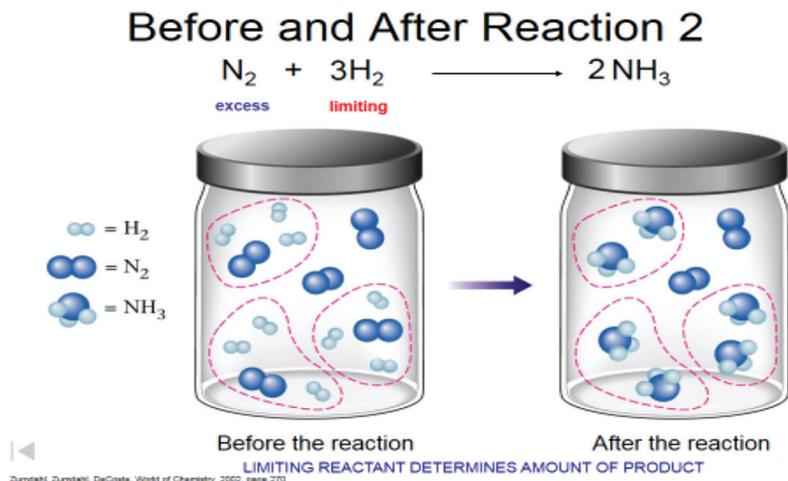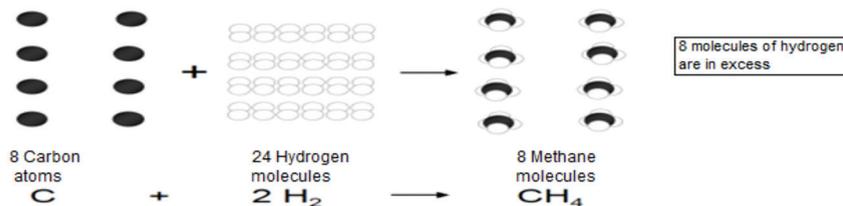
Which of the two molecules, H_2, N_2, is a limiting reagent?

4.3 Worked Examples from Grade 10

Reminder:

At STP: 1 mole of any gas occupies 22,4 dm³ at 0 °C (273 K) and 1 atmosphere.
1. Find the number of moles of ions in:
   (a) 2 moles of $Fe_2(SO_4)_3$
   From the formula, we know that 1 mole of $Fe_2(SO_4)_3$ contains 2 moles of Fe^{3+} ions and 3 moles of SO_4^{2-} ions.

No. of moles of Fe^{3+} ions	=	**2 x 2 = 4**
No. of moles of SO_4^{2-} ions	=	**3 x 2 = 6**
Total no. of moles of ions	=	**4 + 6 = 10**

   (b) 0.2 moles of $Al(NO_3)_3$

No. of moles of Al^{3+} ions	=	**0.2**
No. of moles of NO^{3-} ions	=	**0.2 x 3 = 0.6**
Total no. of moles of ions	=	**0.2 + 0.6 = 0.8**

2. Given 1.6g of methane (CH_4), find:
   (a) The number of moles of CH_4

Molar mass of CH_4	=	12 + 1 x 4 = 16 gmol⁻1
Number of moles of CH_4	=	**mass of CH_4 /molar mass of CH_4**
	=	**1.6/16 = 0.10**

   (b) The number of molecules of CH_4

Number of molecules of CH_4	=	no. of moles x Avogadro Number
	=	**0.10 x 6.02 x 10²³**
	=	**6.02 x 10²²**

(c) The number of H atoms.

One CH_4 molecule contain 4 H atoms.

number of H atoms = **6.02×10^{23} x 4**

= **2.41×10^{23}**

3. Find the mass of:

(a) 1 H_2O molecule **$18 / 6.02 \times 10^{23}$** **3.0×10^{-23} g**	(b) 1 Cu atom **$63.5 / 6.02 \times 10^{23}$** **$1.05 \times 10^{-22}$ g**
(c) 1 Na^+ ion **$23 / 6.02 \times 10^{23}$** **3.8×10^{-23} g**	(d) 1 OH^- ion **$17 / 6.02 \times 10^{23}$** **2.8×10^{-23} g**
(e) 1 neutron **$1.0 / 6.02 \times 10^{23}$** **$1.7 \times 10^{-24}$ g**	(f) 1 electron **$0.00055 / 6.02 \times 10^{23}$** **$9.1 \times 10^{-28}$ g**

4. Find the mass of:

6.02×10^{22} lead atoms **Number of moles of Pb atoms** **= $6.02 \times 10^{22} / 6.02 \times 10^{23}$** **= 0.100** **Mass of Pb atoms** **= 0.100 x 207** **= 20.7 g**	3.01×10^{24} carbon dioxide molecules **Number of moles of CO_2 molecules** **= $3.01 \times 10^{24} / 6.02 c 10^{23}$** **= 5.00** **Molar mass of CO_2** **= 12 + 16 x 2 = 44 g** **Mass of CO_2 molecules** **= 5.00 x 44** **= 220 g**	3.01×10^{23} sulphate ions **Number of moles of SO_4^{2-} ions** **= $3.01 \times 10^{23} / 6.02 \times 10^{23}$** **= 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

= $3.00 \times 6.02 \times 10^{23}$

= 1.806×10^{24}

(b) How many ions are there in 0.600 moles of potassium ions?

Number of potassium ions

= $0.600 \times 6.02 \times 10^{23}$

= 3.612×10^{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:

Elements (atoms)	% Composition divided by molar mass	Divide ALL the answers in the left column by the lowest number
Carbon (C)	$\dfrac{40,0}{12} = 3,33$	$\dfrac{3,33}{3,33} = 1$
Hydrogen (H)	$\dfrac{6,71}{1} = 6,71$	$\dfrac{6,71}{3,33} = 2$
Oxygen (O)	$\dfrac{53,28}{16} = 3,33$	$\dfrac{3,33}{3,33} = 1$

   C = 1; H = 2; O = 1
   The empirical formula = CH_2O

2. Determine the molecular formula of the compound in Question 1, if its molar mass is 180g. mol^{-1}

$M_{CH2}O = 30$ g.mol^{-1}

Ratio = $\dfrac{180}{30}$ = 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

Elements (atoms)	% Composition divided by molar mass	Divide ALL the answers in the left column by the lowest number
Carbon (C)	$\dfrac{74,02}{12}$ = 6,168	$\dfrac{6,168}{1,229}$ = 5
Hydrogen (H)	$\dfrac{8,71}{1}$ = 8,71	$\dfrac{8,71}{1,229}$ = 7
Oxygen (O)	$\dfrac{17,2}{14}$ = 1,229	$\dfrac{1,229}{1,229}$ = 1

Empirical formula is C_5H_7N

$\dfrac{m}{M}$ = n

$\dfrac{40,57}{M}$ = 0,25

162,28 g.mol^{-1} = MNicotine
$M(C_5H_7N) = 81,14$

Ratio = $\dfrac{162,28}{81,14}$ = 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 $CaCl_2$ was found to contain 7.10 g of Cl$^-$ ions.
What mass of Ca^{2+} ions does the sample contain?
Number of moles of Cl$^-$ ions   =   7.10 / 35.5
                                   =   0.200
The formula CaCl$_2$ shows that the ratio of Ca^{2+} ions to Cl$^-$ ions is 1 : 2.
   Hence, the number of moles of Ca^{2+} ions   =   0.200 / 2   =   0.100
          Mass of Ca^{2+} ions in the sample   =   0.100 x 40
                                         =   4.0 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 6.02×10^{22} electrons, calculate n (the charge on each ion of M).

Number of moles for electrons given
= **number of electrons/ Avogadro Number**
= **6.02×10^{22} / 6.02×10^{23} = 0.100**
Number of moles of M ionized
= **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

Substance	Molar mass of substance	Mass of substance present	Number of moles of substance present	Number of molecules/ formula units present
Sulphuric acid (H_2SO_4)	**98g mol-1**	**58.8g**	**0.6 mol**	**3.612×10^{23}**
Sodium hydroxide (NaOH)	**40g mol-1**	**2.0g**	**0.05 mol**	**3.01×10^{22}**
Potassium carbonate (K_2CO_3)	**138g mol-1**	**331.2g**	**2.4 mol**	**1.44×10^{24}**

(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 3913×10^{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:
= **3913×10^{18} / 6.02×10^{23}**
= **6.5×10^{-3} mol**
Mass of nitrogen dioxide
= **6.5×10^{-3} x (14.0 + 2 x 16.0)**
= **0.299g**

7. (I) How many moles of calcium fluoride (CaF2) are present in 16.5g of it?
Formula mass of CaF_2
= 40.0 + 19.0 x 2
= **78.0 g**
Number of moles of CaF_2
= **16.5 / 78.0** **=** **0.21 mol**
(II) How many calcium and fluoride ions are present?
Number of Ca^{2+} ions
= 0.21 x 6.02×10^{23} **=** **1.26×10^{23}**
Number of F^- ions
= 0.42 x 6.02×10^{23} **=** **2.53×10^{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? (Relative atomic mass numbers: C = 12.0; H = 1.0; O = 16.0)

Number of moles of CO_2 = 2.20 / (12.0 + 2 x 16.0) = 0.05 mol

Mass of water

= **0.05 x (1.0 x 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_2SO_4)

        Number of CO_2 molecules

        = **$2 \times 6.02 \times 10^{23}$**

        = **1.20×10^{24}**   **(greater number of molecules)**

        Number of moles of H_2SO_4

        = **8.40 / (1.0 x 2 + 32.0 + 16.0 x 4) = 0.086 mol**

        Number of H2SO4 molecules

        = **$0.086 \times 6.02 \times 10^{23}$**

        = **5.16×10^{22}**

    (b) 88.0g of carbon dioxide (CO_2) or 84.0g of nitrogen (N_2)

        Number of CO_2 molecules

        = **$(88.0 / 44.0) \times 6.02 \times 10^{23}$**

        = **1.20×10^{24}**

        Number of moles of N_2

        = **$(84.0 / 28.0) \times 6.02 \times 10^{23}$**

        = **1.81×10^{24}**   **(greater number of molecules)**

    (c) 5×10^{24} ammonia molecules (NH_3) or 8.00g of sulphuric acid (H_2SO_4)

        Number of NH_3 molecules

        = **5×10^{24}**   **(greater number of molecules)**

        Number of moles of H_2SO_4 molecules

        = **$[8.00 / (1.0 \times 2 + 32.0 + 16.0 \times 4)] \times 6.02 \times 10^{23}$**

        = **4.91×10^{22}**

(Relative 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

            = **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 (AgNO3) solid in 200.0 cm3 solution.

        Number of moles of $AgNO_3$

            = **17.0 / (108.0 + 14.0 + 16.0 x 3)**

            = **0.1 mol**

        Molarity of $AgNO_3$ solution

            = **0.1 / (200.0 / 1000.0)**

        = **0.5 M**

(Relative atomic mass numbers: Ag = 108.0; Cl = 35.5; N = 14.0; Na = 23.0, O = 16.0)

17. Complete the table below:

Substance	Molar mass (g mol-1)	Concentration (g dm-3)	Molarity (M)	Mass of solute required to prepare 250.0 cm3 of solution (g)
KOH	56.0g mol-1	5.60 g dm-3	0.10 M	5.60 g dm-3 × (250.0/1000) dm3 = 1.40g
CuSO4	159.5g mol-1	31.9 g dm^{-3}		7.975g
(COOH) 2˙2H2O	126.0g mol-1	30.24 g dm^{-3}	0.24 M	

(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 cm^3 of 0.10 M sodium hydrogencarbonate ($NaHCO_3$) solution.

Number of moles of $NaHCO_3$

   =   **0.10 x (500.0 / 1000.0)**   =   **0.05 mol**

Mass of $NaHCO_3$

     =   **0.05 x (23.0 + 1.0 + 12.0 + 16.0 x 3)**

     =   **0.05 x 84**

     =   **4.2g**

      (b) 100.0 cm^3 of 3.00 M potassium dichromate ($K_2Cr_2O_7$) solution.

Number of moles of $K_2Cr_2O_7$

   =   **3.00 x (100.0 / 1000.0)**   =   **0.3 mol**

Mass of $K_2Cr_2O_7$

     =   **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 M_2O_3. What is the atomic mass of M?

      Let the atomic mass of M be Ar.

      Mass of M in compound / Mass of oxygen in compound

     =   **2Ar / 16 x 3**

     =   **5.6 / 2.4**

     =   **56**

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

20. A crystalline salt with the formula $M_2S_2O_3$. $5H_2O$ 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)   **5(1 x 2 + 16) / formula mass of hydrated salt**

   =   **36.3 / 100**

   =   **248**

      Formula mass of hydrated salt = 248.

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

      Formula mass of hydrated salt

          =   **$2Ar_2$ + 32 x 2 + 16 x 3 + 5(1 x 2 + 16) = 248**

          Ar   =   **23**

          Therefore, the atomic mass of M is 23.

21. 10.0g of hydrated iron (II) sulphate ($Fe_2SO_4 \cdot nH_2O$) gave 4.53g of water after strong heating. Find the value of n.

	$FeSO_4$	H_2O
Relative mass (g)	5.47	4.53
Relative number of mass	5.47 / 152 = 0.036	4.53 / 18 = 0.252
	0.036 / 0.036 = 1	0.252 / 0.036 = 7

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 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 x 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 MCl_2, the formula of chloride B should be MCl_3.

More activities and solutions

Exercises 1

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

$$2H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(\ell)$$
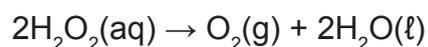
1.1 What does (aq) represent in the equation above? (1)

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_2O_2, if 17g of H_2O2 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 $n(H_2O_2) : n(O_2)$
   $\qquad$ 2:1
   Therefore: $n(O_2)$ = 2 mol
   $n(O_2)$ =
   2 =
   V = 44,8 dm3
   (4)

1.6 $n(H_2O_2)$ =
   $n(H_2O_2)$ =
   $n(H_2O_2)$ = 0,5 mol

   $n(O_2)$ =
   (0,5)(2) =
   N = atoms
   (4)

$n(H_2O_2) : n(O_2)$

      2:1

Therefore: $n(O_2)$ = 2 mol

$n (O_2)$ =

2 =

V = 44,8 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 $Na_2CO_3.xH_2O$.

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

   [14]

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:

Elements (atoms)	Mass divided by molar mass	Divide ALL the answers in the middle column by the lowest molar mass of all
Carbon (C)	$\dfrac{40,0}{12} = 33,3$	$\dfrac{3,3}{3,3} = 1$
Hydrogen (H)	$\dfrac{6,6}{1} = 6,6$	$\dfrac{6,6}{3,3} = 2$
Oxygen (O)	$\dfrac{53,3}{16} = 3,3$	$\dfrac{3,3}{3,3} = 1$

   C = 1; H = 2; O = 1
   The empirical formula = CH_2O (5)

2.1.3 M_{CH_2O} = 30g.mol-1
   Ratio = = 2
   Molecular formula: multiply each of the number of atoms in the empirical formula by 2.
   Molecular formula = $C_2H_4O_2$ (3)

2.2 M (Na_2CO_3) = 106g·mol-1
   M (x H_2O) = 268 – 106 = 162g·mol-1
   n(H_2O) =
   n(H_2O) = 9 mol (4)

**Exercise 3**

Calcium carbonate ($CaCO_3$) reacts with dilute hydrochloric acid ($HC\ell$) according to the following balanced equation:

$$CaCO_3(s) + 2HC\ell(aq) \rightarrow CaC\ell_2(aq) + CO_2(g) + H_2O(\ell)$$

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.
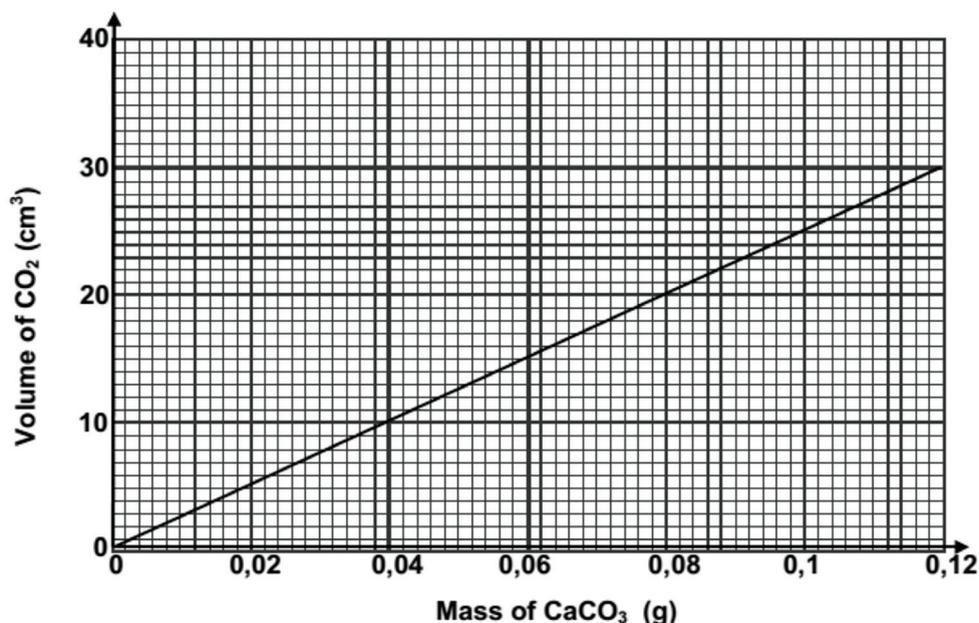
3.2 From the graph, determine the volume of $CO_2(g)$ produced when 0,072g of PURE $CaCO_3(s)$ reacts. (1)

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

[13]

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

3.2 18 cm³ (1)

3.3.1 The number of moles of solute per cubic decimetre/ litre of solution. (2)

3.3.2 %CaCO₃ = x 100
     %CaCO₃ = 40% (3)

3.3.3 n(CaCO₃) =
      n(CaCO₃) =
      n(CaCO₃) = 0,001
   n(HCℓ) = 2 x n(CaCO₃) = 2 x 10⁻³ mol

      Volume of acid
      c =
      0,1 =
      V = 0,02 dm3 (5)

[13]

Exercise 4

4.1 Study the balanced chemical equation of the reaction between sodium carbonate (Na_2CO_3) and hydrochloric acid ($HC\ell$) and answer the questions that follow.

$$Na_2CO_3 + 2HC\ell \rightarrow 2NaC\ell + CO_2 + H_2O$$
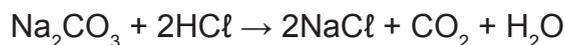
Identify the type of reaction above. Choose between REDOX and GAS FORMING. (1)

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. (2)

    4.2.2 Calculate the initial number of moles of sodium carbonate. (2)

    4.2.3 Calculate the mass of CO_2 produced during this reaction. (4)

    4.2.4 Calculate the mass of sodium chloride produced if 4,87 dm³ of carbon dioxide is produced at STP. (6)

4.3 14,2g of a sample of hydrated sodium carbonate ($Na_2CO_3 \cdot xH_2O$) was heated until no further change in mass was recorded. On heating, all the water of crystallisation evaporated as follows:

$$Na_2CO_3 \cdot xH_2O \rightarrow Na_2CO_3 + x\, H_2O$$
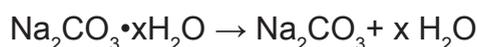
If 4 moles of hydrogen peroxide decompose, calculate the volume of gas formed at STP.

Calculate the number of oxygen atoms in H_2O if 17g of H_2O_2 decomposes. (5)

[20]

4.1 Gas forming (1)

    4.2.1 $M(Na_2CO_3) = 2(23) + 12 + 3(16)$
        $M(Na_2CO_3) = 106$g.mol-1 (2)

    4.2.2 $n(Na_2CO_3) =$
        $n(Na_2CO_3) =$
        $n(Na_2CO_3) = 0,1$ mol (2)

    4.2.3 $n(Na_2CO_3) : n(CO_2)$
              1 : 1
          Thus: $n(CO_2) = 0,1$mol
        $n(CO_2) =$
        $0,1 =$
        $m = 4,4$ g (4)

    4.2.4 $n\,(CO_2) =$

n (CO_2) =
n (CO_2) = 0,217 mol

n(CO_2) : n(NaCl)
        1 : 2
    n(NaCl) = 0,434 mol
n(NaCl) =
0,434 =
m = 25,16 g   (6)

4.3 Mass of H_2O = 14.2 – 5,3
            = 8,9 g

n(Na_2CO_3) =   n(H_2O) =
n(Na_2CO_3) =   n(H_2O) =
n(Na_2CO_3) = 0,05 mol   n(H_2O) = 0,5 mol
Na_2CO_3 : H_2O
0,05 : 0,5
0,05 : 0,05
 1 : 10
Thus, X = 10   (5)

[20]

Exercise 5

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

$$2Na(s) + 2H_2O(\ell) \rightarrow 2NaOH(aq) + H_2(g)$$
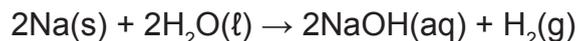
5.1 Write down the value of the temperature and the pressure at STP. (2)

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)

[17]

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

    5.2.1 $n(Na) = \dfrac{m}{M}$

         $n(Na) = \dfrac{10}{23}$
         $n(Na) = 0,43$ mol
         $Na : H_2$
         $2 : 1$
         Thus, 0,22 mol H_2 produced

         $n(H2) = \dfrac{m}{M}$

         $0,22 = \dfrac{m}{23}$
         $m = 0,44$ g H_2 produced (5)

    5.2.2 $n(H2) = \dfrac{V}{V_m}$

         $0,22 = \dfrac{V}{22,4}$
         $V = 4,93$ dm^3 (3)

    5.2.3 $n(Na) : n(NaOH)$
         $2 : 2$
         Thus, mol NaOH = 0,43 mol

         $n(NaOH) = \dfrac{m}{M}$

         $0,43 = \dfrac{m}{40}$
         $m = 17,2$ g (4)

5.2.4 $c = \dfrac{n}{V}$

$c = \dfrac{0,43}{2}$

$c = 0,22$ mol.dm^{-3} (3)

[17]

(3)

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:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$
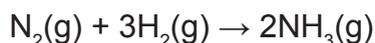
6.1.1 Define the term 1 mole. (2)

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)

[16]

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 (1)

6.1.3 Vol(N_2) : V(H_2) : V(NH $_3$) = 1 : 3 : 2
Vol(N_2 reacted = V(H_2)
Vol(N_2 reacted = (24)
Vol(N_2 reacted = 8 dm³
Volume N_2 remains = 10 – 8 = 2 dm³
The volume of gas that remains = 2 + 16
The volume of gas that remains = 18 dm³ gas (4)

6.1.4 n $= \dfrac{m}{M}$

n $= \dfrac{80}{2}$

n = 40 mol (2)

6.1.5 $Vol(N_2) = \dfrac{1}{3}$ (40) x 22,4

$Vol(N_2) = 2\,98,67\ dm^3$

6.2 $m(O_2) = 239 - 207 = 32\ g$

$n_{Pb} = \dfrac{m}{M}$

$n_{Pb} = \dfrac{207}{207}$

$n_{Pb} = 1\ mol$

$nO_2 = \dfrac{m}{M}$

$nO_2 = \dfrac{32}{M}$

$n_O = 2\ mol$

$n(Pb) : n(O)\ 1 : 2$

PbO_2 (5)

[16]

Exercise 7

Nitric acid *(HNO₃)* 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 cm³ 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.
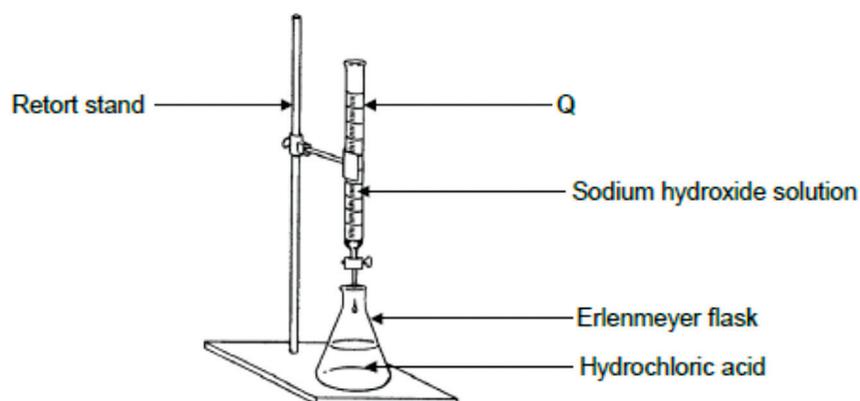
7.3.2 Write down the name of the apparatus Q in the above diagram. (1)

7.3.3 The following indicators are available for the titration:

INDICATOR	pH RANGE
A	3,1 – 4,4
B	6,0 – 7,6
C	8,3 – 10,0

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 cm³ 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:

$$MgO_{(s)} + 2HC\ell_{(aq)} \rightarrow MgC\ell_2(aq) + 2H_2O(\ell)$$

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. (1)

8.2 NO_3^- / Nitrate ion

8.3 pH= -log[H_3O^+]

$$= - \log (0,3)$$

$$=0,52 \tag{3}$$

8.2

1.2.1 c= n/V

2= n/0,1

n(HCℓ)=0,2 mol (3)

8.2.2 Burette (1)

8.2.3 B

Titration of a strong acid and a strong base. (3)

8.2.4 The number of moles of acid in the flask remains constant. (1)

8.2.5 c= n/V

0,2= n/0,021

n=4,2 × 10^{-3} mol

n(HCl in excess)=n(NaOH)

=4,2 × 10^{-3} mol (3)

n(HCl in excess):

0,2 - 4,2 × 10^{-3} =0,196 mol

n(MgO reacted):

= (1/2) n(HCl) = n(MgO)

= (1/2)(0,196)

= 9,8 × 10^{-2} mol

m(MgO)=nM

= (0,098)(40)

= 3,92 g

% purity = 3,92/4,5 ×100 = 87,11% (5)

[21]

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:

$$HCO^-_{3(aq)} \; + \; H_2O_{(\ell)} \quad H_2CO_{3\,(aq)} \; + \; OH^-_{(aq)}$$

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

    9.2.2 $HCO^-_{3(aq)}$ acts as a base in the above reaction. Write down the formula of the conjugate acid of $HCO^-_{3(aq)}$.   (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 H_2SO_4 ionises completely.)

$$H_2SO_{4(aq)} \; + \; 2NaHCO_3(s) \; \rightarrow \; Na_2SO_{4(aq)} \; + \; 2H_2O(\ell) \; + \; 2CO_{2(g)}$$

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)

[22]

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 H_2SO_4 molecule.   (1)

9.2
    9.2.1 Amphiprotic (substance)/ Ampholyte   (1)

    9.2.2 $H_2SO_{3(aq)}$ (1)

9.3
    9.3.1 $n(NaHCO_3) = m/M$
    $= 27/84$
    $= (0{,}321485714 \text{ mol})$
               $= 0{,}32 \text{ mol}$

      $n(H_2SO_4) = n(NaHCO_3) = (1/2)(0{,}32)$
    $= 0{,}16 \text{ mol} \ (0{,}1607142857 \text{ mol})$

    $c = n/V$
    $6 = 0{,}16/V$
    $V = 0{,}027 \text{ dm}^3$ (6)

    9.3.2 $n_a \text{ (initial)} = n_a \text{ (final)}$
    $c_a V_a \text{ (initial)} = c_a V_a \text{ (final)}$
    $(6) V_a = (01)(1)$
    $V_a = 0{,}017 \text{ dm}^3$ (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

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

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

$$n\_a/n\_b = 1/2$$
$$n_a \text{ (neutralised)} = (1/2) n_b = (1/2)(3 \times 10^{-3})$$
$$= 1,5 \times 10^{-3} \text{ mol}$$

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

$$c_a = n/V$$
$$= (1 \times 10^{-3})(25 \times 10^{-3}) + (30 \times 10^{-3})$$
$$= 0,018 \text{ mol·dm}^{-3}$$

$$pH = -\log [H\_3 O^+]$$
$$= -\log (2 \times 0,018)$$
$$= 1,44$$

(8)

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.

EXPERIMENT	DILUTE ACID
1	100 cm³ of 0.1 mol·dm⁻³ H_2SO_4
2	50 cm³ of 0.2 mol·dm⁻³ H2SO₄
3	100 cm³ of 0.1 mol·dm⁻³ HCℓ

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.
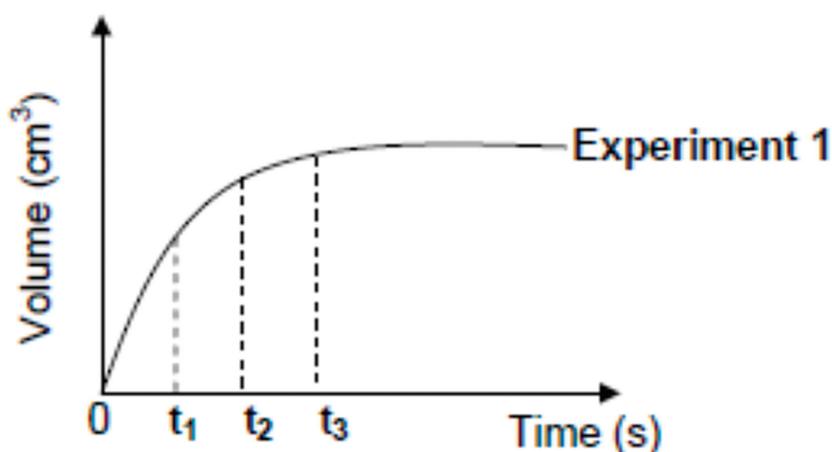
Use this graph and answer the questions that follow.

10.1 At which time (t_1, t_2 or t_3) is:
   10.1.1 The reaction rate the highest? (1)
   10.1.2 The mass of the zinc present in the flask the smallest? (1)

10.2 At which time interval does the largest volume of hydrogen gas form per second? Choose from: **between t_1 and t_2** OR **between t_2 and t_3.** (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:

$$Zn_{(s)} + 2HC\ell_{(aq)} \rightarrow ZnC\ell_{2(aq)} + H_{2(g)}$$

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)

[15]

Solutions

10.1. Time: Stopwatch
    Volume: Gas syringe / Burette / Measuring cylinder / Chemical balance /
    Erlenmmeyer flask / Graduated flask   (2)

10.2.
    10.2.1. t_1 (1)
    10.2.2. t_3 (1)

10.3. Between t_1 and t_2 (1)

10.4.
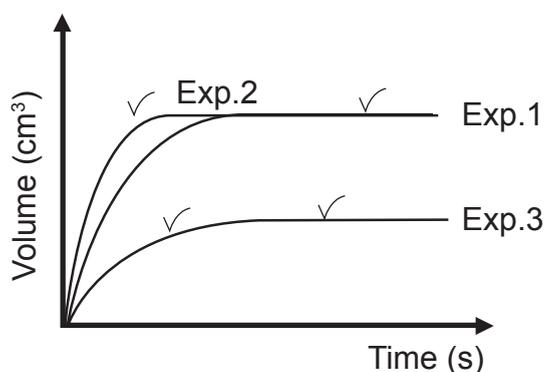
Marking criteria	
Exp. 2	Initial gradient higher than that of Exp. 1
	Curve reaches same constant volume as for Exp. 1, but earlier.
Exp. 3	Initial gradient lower than that of Exp. 1.
	Curve reaches a smaller constant volume, as for Exp. 1, at a later stage.

(4)

10.5.1 $c(HC\ell) = n/V$
    $0,1 = n/100 \times 10^{-3}$
    $n(HC) = 0,01$ mol

    $n(Zn) = \square (1/2) \, n(HC) = 0.005$ mol

    $n(Zn) = m(Zn_{reacted})/M$
    $0.005 = m(Zn_{reacted})/(65)$
    $m(Zn_{reacted}) = 0.325$ g
 $m(Zn_f) = 0.8 - 0.325 = 0.475$ g   (5)

10.5.2 Smaller than   (1)

[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:

$$3X_2 \rightleftharpoons 2X_{3(g)}$$
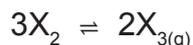
11.1. How will the rate of formation of $X_{3(g)}$ compare to the rate of formation of $X_{2(g)}$ at equilibrium? Write down only one of the following: HIGHER THAN; LOWER THAN; EQUAL TO. (1)

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

TIME (s)	[X_2] (mol · dm^{-3})	[X_3] (mol · dm^{-3})
0	0,4	0
2	0,22	0,120
4	0,08	0,213
6	0,06	0,226
8	0,06	0,226
10	0,06	0,226

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.
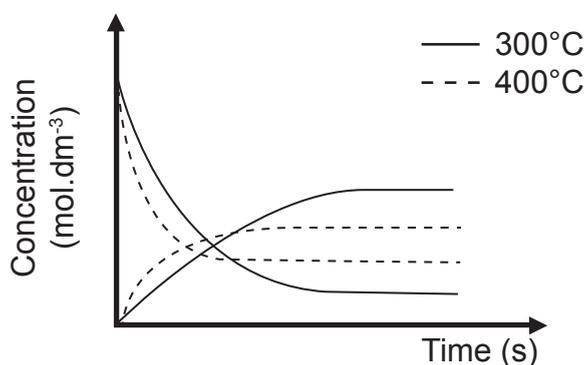
11.5. Is the forward reaction EXOTHERMIC or ENDOTHERMIC? Explain in full how you arrived at your answer.

The Maxwell-Boltzmann distribution curve below represents the number of particles against kinetic energy at 300°C. (4)
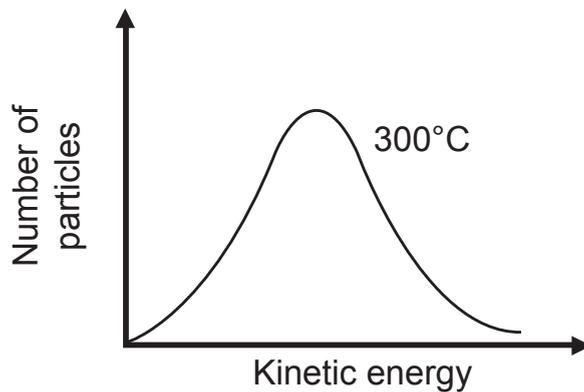
11.6. Re-draw this curve in your ANSWER BOOK. On the same set of axes, sketch the curve that will be obtained at 400°C. Clearly label the curves as 300°C and 400°C, respectively. (2)

[15]

Exercise 12

12.1 Equal to (1)

12.2 $K_c = \dfrac{[X_3]^2}{[X2]^2}$

$$= \dfrac{(0,226)^2}{(0,06)^2}$$

$$= 236,46$$

(4)

12.3
  6.3.1 Increases (1)
  6.3.1
    ● The increase in [X3] is opposed. The change is opposed.
   ● The reverse reaction is favoured. X_3 is used [X_3] decreases

(2)

12.4 Higher than (1)

12.5 Exothermic
  ● [X3] decrease and [X_2] increases.
  ● K_c decreases if temperature increases.
  ● The decrease in temperature favoured the forward reaction. (4)

12.6

Kinetic energy / Kinetiese energie
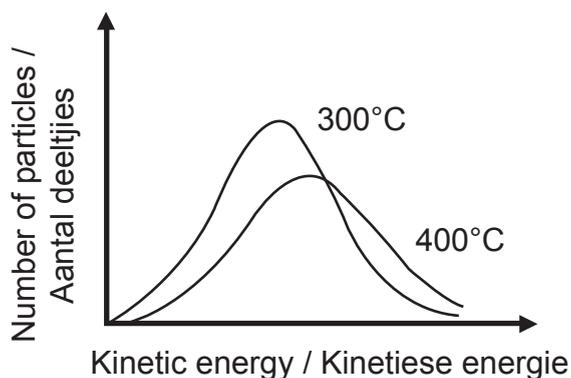
Marking criteria	
The peak of the curve at 400°C is lower than at 300°C and it shifted to the right.	
The curve at 400°C has a larger area at the higher E_k.	

(2)

Exercise 13

Initially, *2,2 g* of pure $CO_{2(g)}$ is sealed in an empty 5 dm³ 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:

$$CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$$

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_c) for this reaction is 2,6 × 10⁻⁶ at 727°C.
Is the reaction EXOTHERMIC or ENDOTHERMIC? Explain in full how you arrived at the answer. (4)

[19]

Solutions

6.1 $M(CO_2) = 12 + 2(16) = 44 \text{ g·mol}^{-1}$

$$c = m/MV$$
$$= 2,2/(44)(5)$$
$$= 0,01 \text{ mol·dm}^{-3} \qquad (4)$$

6.2 If only CO_2 is present, the reverse reaction cannot take place. (1)

6.3 CO_2 is a gas that will escape if the container is not <u>sealed.</u> (1)

6.3 $K_c = [CO_2]$
    $= 0,0108$
    $[CO_2] = 0,0108 \text{ mol·dm}^{-3}$

		$CaCO_3$	CaO	CO_2
Ratio		1	1	1
Initial quantity (mol)		0,004	0	0,05
Change (mol)	Used	0,004		
	Form			0,004
Quantity at equilibrium (mol)				0,054
Equilibrium concentration (mol·dm³)				0,0108

$$m(CaCO_3) = nM$$
$$= (0,004)(100)$$
$$= 0,4 \text{ g}$$

$\qquad (7)$

6.5
   8.5.1 Remains the same. (1)
   8.5.2 Decreases. (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. (4)

[19]

Exercise 14

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

$$CH_3OH_{(aq)} + HC\ell_{(aq)} \rightarrow CH_3C\ell_{(aq)} + H_2O(\ell)$$
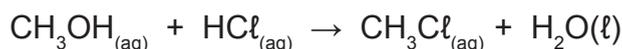
14.1 State TWO factors that can INCREASE the rate of this reaction. (2)

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 $HC\ell_{(aq)}$ was measured at different time intervals. The following results were obtained:

TIME (MINUTES)	HCℓ CONCENTRATION (mol· dm⁻³)
0	1,90
15	1,45
55	1,10
100	0,85
215	0,60

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 $HC\ell_{(aq)}$ 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 $CH_3 C\ell_{(aq)}$ in the flask at the 215th minute. The volume of the reagents remains 60 cm³ during the reaction. (5)

Exercise 15

15.1 Increase temperature.
Increase the concentration of acid.
Add a catalyst. (2)

15.2 Change in concentration of products / reactants per unit time.
OR
Rate of change in concentration. (2)

15.2
6.1.3 Average rate$=-\Delta c/\Delta t$
$$=\frac{-(1,45 - 1,9)}{(15 - 0)}$$

$=0,03 \ (mol \cdot dm^{-3}) \cdot min^{-1}$ (3)

Graph of concentration versus time
Grafiek van konsentrasie teenoor tyd
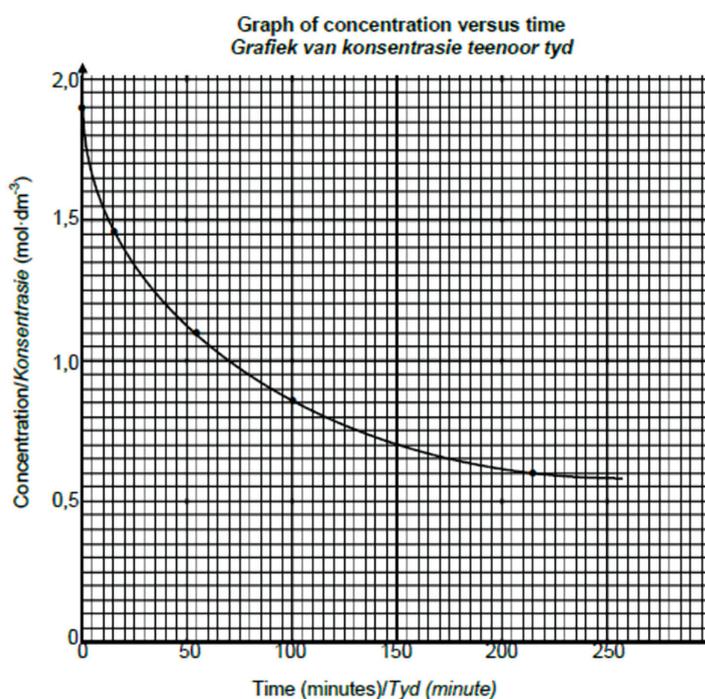
Marking criteria	
Four points correctly plotted.	
Curve drawn as shown.	

15.2.3 1,2 mol$\cdot$ dm^{-3} (1)

15.2.4
● The concentration of reactants decreases.
● Less particles per unit volume.
● Less effective collisions per unit time. (3)

15.2.5 $\Delta c(HCl) = 0,6-1,9$

$$= -1,3$$
$$= 1,3 \text{ mol} \cdot \text{dm}^{-3}$$

$$c(HCl) = n/V$$

$$1,3 = \frac{n}{60 \times 10^{-3}}$$
$$n(HCl) = 0,078 \text{ mol}$$

$$n(CH_3l\ Cl) = n(HCl) = 0.078 \text{ mol}$$

$$n(CH_3\ Cl) = \frac{m}{M}$$

$$0.078 = \frac{m}{(50,5)}$$

$m(CH_3\ Cl) = 3,94 \text{ g}$

(5)

PHYSICAL SCIENCE_PAPER 2

STOICHIOMETRY

GRADE 12