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Simple Solutions

Counting Moles ...

Nigel P. Freestone

Bentham e Books

COUNTING MOLES

SIMPLE SOLUTIONS - CHEMISTRY COUNTING MOLES

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FOREWORD

When people I meet ask me what I do and I mention that I trained as a chemist, they usually wince. This is often followed by a comment such as, “there’s too much maths in chemistry”; “learning chemistry is like learning a foreign language”; “there’s a lot to learn”. People very rarely say, “I wish I’d learned more chemistry at school” or “chemistry is so interesting”. This is hugely frustrating for me as chemistry is everywhere – everything we see, touch, taste, smell and even hear involves chemicals and chemistry. I think it is a brilliant, fascinating, complex and simple subject that shapes and changes our world for the better.

However, having taught various aspects of (mainly environmental) chemistry from primary school to PhD levels for more than a quarter of a century, I’m very aware that chemistry can be viewed as difficult, theoretical and even irrelevant to people’s everyday lives. This is a terrible shame and may have unforeseen adverse consequences. The world is facing a large number of global challenges that include anthropogenic climate change, sustainable development, the need for adequate supplies of clean water and other key resources such as metals, growing energy demands, concerns about food supply and quality, and the impacts on nutrition and health. Chemistry can make a significant contribution to helping us to mitigate or solve these challenges – but for that to happen, we need more chemists.

I know from talking to colleagues in industry, business and academia that there is a potential shortage of human resources in key scientific professions. There have been numerous pleas for the rejuvenation of science teaching and initiatives to increase the interest and attainment of students in science and mathematics. Basic scientific skills such as communication, classification, measurement, observation, inference and prediction contribute to a larger purpose - problem solving. These skills allow scientists to conduct systematic, objective investigations and to reach conclusions based on the results. Such skills are vital in chemistry. We need young people who are mathematically and scientifically literate and who engage with scientific problems in an orderly, “can-do” fashion. We cannot allow a shortfall in these skills to emerge simply through neglect or fatalism or worse – because subjects such as chemistry are “hard”, “dangerous”, “geeky”, “too mathematical” or require “too much work”.

Therefore I’m delighted to welcome Nigel Freestone’s refreshing book “Counting Moles” to the stable of chemistry textbooks. Nigel has a wealth of experience in delivering chemical concepts to young people in an engaging and inspiring fashion. This book is designed to provide multiple opportunities for students to develop a firm understanding of chemical concepts that underpin the whole subject. Students will be able to use the multiple worked examples and calculating frame to hone their skills and become more confident in fundamental problem-solving activities, culminating in the award of a Counting Moles

Driving Licence that confirms their achievement.

I'm sure Nigel's book will make a significant contribution to addressing the chemistry skills shortage and to ensuring that this wonderful subject continues to make its distinct contribution to solving the world's challenges.

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PREFACE

Students studying chemistry often struggle with the **MOLE.** This user-friendly self-teach package provides an effective aid to learning by giving clear and confident presentation of the essentials of the mole needed by those starting chemistry courses. This self teach package contains over 200 questions, with detailed solutions, so if you get stuck, you can see where you went wrong.

After successfully completing Counting Moles, you will be able to:

- understand what is meant by the terms relative atomic mass and relative formula mass
- calculate average atomic mass from isotope abundance;
- determine relative formula mass;
- define of the chemical term, the 'mole';
- calculate the molar mass of a substance from its chemical formula;
- calculate percentage composition;
- interconvert between mass, number of molecules, and moles;
- calculate the amount (mass or moles) of product expected to be formed in a chemical reaction, given the amounts of reactants used;
- calculate the amount (mass or moles) of reactants which need to be used in a chemical reaction in order to produce a specified amount of product;
- define the term concentration, and calculate the molarity of solutions from volumetric analysis and other data;
- determine the mass and/or volume of gases consumed or formed in a chemical reaction.

Counting Moles is split into SIX (6) Chapter. Each Chapter should be thought of as a separate lesson that should take between 15 and 45 minutes to complete. It is **ESSENTIAL** that you answer each question. The Counting Moles approach to learning is based on understanding developed through repetition and introduces the '**mole calculating frame'** to help you solve problems. Answers to ALL questions, with full explanations are provided, should you get into difficulty.

The Chapters prepare you for the **Counting Moles Driving Test**. If you PASS, you can throw you M-plate away and claim your **Counting Moles Driving Licence**. You will then be able to answer moles-based coursework and examination questions correctly and with confidence.

CONFLICT OF INTEREST

The author confirms that author has no conflict of interest to declare for this publication.

ACKNOWLEDGEMENTS

Declared none.

CHAPTER 1

Relative Atomic & Formula Mass, Percentage Composition, Empirical & Molecular Formulae

Keywords: Atoms, Average relative atomic mass, Calculations, Empirical and molecular formula, Isotopes, Percentage composition, Relative atomic mass (A_r), Relative formula mass (M_r).

1.1. RELATIVE ATOMIC MASS (A_R)

Chemical elements are the building blocks from which everything is constructed, from specks of dust to mobile phones and from flora and fauna to the clothes we wear. There are over a 100 known elements. An **element** is a pure substance that cannot be chemically broken down. The smallest unit of an element is the **atom**. Different atoms have different masses. The mass of an atom is so small that it is more convenient to compare atom masses, rather than refer to their actual mass. The standard for this relative scale is an atom of carbon-12, which has a relative atomic mass (A_r) of 12.

The Table below lists the relative atomic mass (A_r) values of some common elements.

Selected Relative Atomic Mass Values

Element	Approximate Relative Atomic Mass (A_r)
Hydrogen	1
Carbon	12
Nitrogen	14
Oxygen	16
Sodium	23
Magnesium	24
Silicon	28
Calcium	40

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Table contd.....

Element	Approximate Relative Atomic Mass (A_r)
Bromine	80

These relative atomic mass values tell us for example that sodium atoms (A_r = 23) are 23 times heavier than hydrogen atoms (A_r = 1), two atoms of neon (A_r = 20) have the same mass as one atom of calcium (A_r = 40) and that three oxygen atoms (A_r = 16) weigh the same as two magnesium atoms (A_r = 24).

Relative atomic masses are listed in the periodic table.

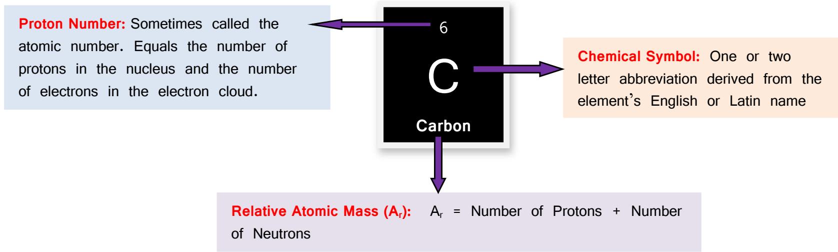
Isotopes

Elements are defined by their proton (atomic) number. An atom with 7 protons is always nitrogen (N), an atom with 20 protons is always calcium and an atom with 70 protons must therefore always be gold (Au). Isotopes are atoms that have the same number of protons, but have different numbers of neutrons. For example, carbon has three naturally occurring isotopes, often referred to as simply carbon-12, carbon-13 and carbon-14, with relative atomic masses of 12, 13 and 14, respectively. Since carbon has a proton number of 6, the isotopes contain 6, 7 and 8 neutrons, respectively.

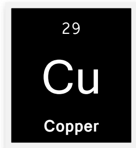
Three Isotopes of Carbon

Isotope	No. of protons	No of neutrons	No of electrons	Relative Atomic Mass (A_r)
$^{12}C_6$	6	6	6	12
$^{13}C_6$	6	7	6	13
$^{14}C_6$	6	8	6	14

A typical periodic table information box for the element carbon is given below:



The relative atomic masses listed in the periodic table are an average of the masses of the different naturally occurring isotopes. The rounded value gives you the mass of the most abundant isotope. For example, copper occurs naturally as Cu-63 and Cu-65. Given that the average relative atomic mass of copper is 63.546, we can conclude that Cu-63 is the more abundant isotope since the average value is closer to 63 than to 65.



Calculating Average Relative Atomic Mass

Chlorine has a relative atomic mass of 35.5, which is the average of the masses of its two naturally occurring isotopes. This is calculated by working out the relative abundance of each isotope. For example, in any sample of chlorine, 25% will be Cl-37 and 75% Cl-35. The relative atomic mass is therefore calculated using the equation:

(% of isotope 1/100 × mass of isotope 1) + (% of isotope 2/100 × mass of isotope 2)

So in the case of chlorine:

Average relative atomic mass = (75/100 x 35) + (25/100 × 37) = 26.25 + 9.25 = **35.5**

Example 1.1: *Boron has three naturally occurring isotopes: 9% boron-10 and 80.1% boron-11. Calculate the relative atomic mass of boron.*

Answer:

Average relative atomic mass = (19.9/100 x 10) + (80.1/100 x 11) = 1.99 + 8.811 = **10.801**

Example 1.2: *Magnesium has three isotopes, Given that the natural abundances are Mg-24 (78.70%), Mg-25 (10.13%), and Mg-26 (11.17%) calculate the relative*

CHAPTER 2

The Mole

Keywords: Avogadro's number, Calculations, Formula units, Molar mass, Relative formula mass, The mole.

The Mole is simply a number. Just as the term dozen refers to the number (12) twelve and a score to the number (20) twenty, the mole refers to the number 6.023 x 10^{23}. Thus 12 eggs is a dozen of eggs, 20 eggs is a score of eggs and 6.023 x 10^{23} eggs is a mole of eggs. Commonly referred to as Avogadro's constant, 6.023 x 10^{23} is the number of atoms found in exactly 12 grams of carbon-12. Carbon-12 is used as the standard from which atomic masses are measured: its mass number is 12 by definition. Since 12 g of carbon contains one mole of carbon atoms, the mass of one mole of any element is equal to its relative atomic mass in grams. Magnesium has relative atomic mass of 24. Therefore one mole of magnesium has a mass of 24 g. Thus 24 g of magnesium contains 6.02 x 10^{23} magnesium atoms and the mass of one atom of magnesium = 24/(6.02 x 10^{23}) = 3.987 x 10^{-23} g. Similarly, the mass of one mole of lithium (A_r = 7) is 7g, 27 g of aluminium (A_r = 27) contains one mole of atoms and the mass of one mole of calcium (A_r = 40) is 40g *etc.* You can also work with fractions (or multiples) of moles:

Mole/Mass Relationship Examples Using Magnesium (A_r =24)		
Moles Magnesium	**Number of Magnesium Atoms**	**Mass of Magnesium**
0.25	1.505 x 10^{23}	6 g
0.5	3.01 x 10^{23}	12 g
1	6.02 x 10^{23}	24 g
2	1.204 x 10^{24}	48 g
10	6.02 x 10^{24}	240 g
50	3.01 x 10^{25}	1200 g

Some elements exist as molecules rather than atoms. The following elements all exist as diatomic molecules: hydrogen (H_2), nitrogen (N_2), oxygen (O_2) and the

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halogens (F_2, Cl_2, Br_2, I_2). Hydrogen has a relative atomic mass of 1. Therefore, the relative formula mass of (M_r) of H_2 = (2 x 1) = 2. Therefore, one mole of hydrogen molecules will have a mass of 2g and will cont 6.02 x 10^{23} molecules of hydrogen. Oxygen has a relative atomic mass of 16. Thus one mole of oxygen gas (O_2) has a mass of 32 g and 6.02 x 10^{23} molecules of nitrogen gas (N_2) have a mass of 28g.

The concept of a mole is equally applicable to compounds as well as elements. The mass of one mole of a compound is its relative formula mass (M_r) in grams. To avoid any ambiguity it is convenient to use the term **formula unit**. Formula unit refers to the smallest repeating unit of a substance and is the chemical formula normally used for the substance. For instance, the formula unit of graphite is an atom of carbon (C). Similarly, the formula unit of oxygen gas is an oxygen molecule (O_2); NaCl is the formula unit for the ionic compound sodium chloride and the formula unit for silicon dioxide is SiO_2.

Equimolar amounts of substances contain the same number of formula units. Thus 0.5 moles any substance will contain the same number of formula units (particles), *i.e.* 0.5 x 6.02 x 10^{23} = 3.01 x 10^{23}.

The idea of the **mole** links the **mass of a substance** to the **number of formula units** (particles) it contains. The mass of one mole of an element or compound is referred to as its **molar mass**, which is its **relative atomic mass (A_r)** or **relative formula mass(M_r)** in grams.

Molar Mass (M_r) = Relative Formula Mass in grams (g mol^{-1})

If you have *m* grams of a substance which has a molar mass of M_r g mol^{-1}, then the amount of a substance in moles, *n*, is given by:-

Number of moles = Mass (g) / Molar Mass (g mol^{-1})
Number of formula units (particles) = Number of moles x 6.23 x 10^{23}
Number of moles = mass (*g*) / M_r (g mol^{-1})

Thus if you know the values of any two of n, m or M_r you can calculate the third using the equations above.

Water has a relative formula mass of 18. Thus:

- one mole of water has a mass of 18 g;
- 18 g of water contains 6.02×10^{23} formula units (molecules) of water;
- 0.5 moles of water has a mass of 9 g and contains 3.01×10^{23} molecules of water;
- one molecule of water has a mass of $18/(6.02 \times 10^{23}) = 2.99 \times 10^{-23}$ g.

Example 2.1: *Determine the mass of one mole of O_2?*

Answer:

Relative formula mass of O_2 = (2 x 16) = 32
Mass of one mole, *i.e.* molar mass = M_r in g
Molar Mass, $M_r O_2$ = **32 g mol^{-1}**

Example 2.2: *What is the mass of 0.05 moles of ammonium sulfate?*

Answer:

Relative formula mass of $(NH_4)_2SO_4$ = (2 x 14) + (8 x 1) + 32 + (4 x 16) = 132
Mass of one mole, *i.e.* molar mass = M_r in g
Molar Mass, M_r $(NH_4)_2SO_4$ = 132 g mol^{-1}
Therefore, 0.05 moles of ammonium sulfate has a mass of 132 x 0.05 = **6.6 g**

Example 2.3: *How many moles of substance are present in 0.250 g of calcium carbonate?*

Answer:

Relative formula mass of $CaCO_3$ = 40 + 12 + (3 x 16) =100
Mass of one mole, *i.e.* molar mass = M_r in g
Molar Mass, M_r $CaCO_3$ = 100g mol^{-1}
So the number of moles in 0.250 g of $CaCO_3$ = mass/M_r = 0.250/100 = **2.5×10^{-3} moles**

Example 2.4: *How many formula units are present in 9 g of KNO_3?*

CHAPTER 3

Moles & Balanced Chemical Equations

Keywords: Balancing chemical equations moles, Calculations, Calculating frame, Chemical equations, Products, Reactants, Reaction coefficients, Stoichiometry.

3.1. CHEMICAL EQUATIONS

Chemical equations are a form of shorthand that scientists use to describe chemical reactions, where the reactants are given on the left-hand side and the product on the right hand side.

$$\text{Reactants} \rightarrow \text{Products}$$

Chemical reactions simply involve the rearrangement of the same atoms through the breaking of chemical bonds and the formation of new bonds. Equations are balanced when the number and type of reactant and product atoms are identical. To balance equations, whole numbers, known as **reaction coefficients** are placed to the left hand side of chemical formula. **Remember** you cannot change the chemical formula of a substance to balance a chemical equation. Mass is always conserved. Atoms of gold cannot be produced from atoms of lead.

Sodium carbonate breaks down when heated to produce sodium carbonate, carbon dioxide and water. This thermal degradation can be represented by the following word equation:

$$\text{sodium hydrogen carbonate} \rightarrow \text{sodium carbonate} + \text{carbon dioxide} + \text{water}$$

Although this word equation is a convenient way of expressing a chemical reaction, it really does not provide us with any more information than the opening sentence of the paragraph above. But this all changes when we replace the names of the chemicals with chemical formula.

$$NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$$

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Atom Count (unbalanced):

	Reactants	Products
Na atoms	1	2
H atoms	1	2
C atoms	1	2
O atoms	3	6

The above reaction is unbalanced since the number of atoms of each element on the left hand side of the equation is not the same as the number of atoms of each element on the right hand side. To balance the equation the relative number of molecules or formula units that participate in the reaction are changed. Reaction coefficients are placed to the left of the chemical formula, *e.g.*, $2NaHCO_3$, until the number of atoms of each element on either side of the equation balance. For most chemical equations it is often best to start by balancing the element with the least atoms present.

Once one element is balanced, proceed to balance another, and another, until all elements are balanced. Thus for the reaction above, we can see that H appears in one reactant ($NaHCO_3$) and one product (H_2O). H's can be balanced by using '2' as the coefficient for $NaHCO_3$.

$$2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$$

Atom Count (balanced):

	Reactants	Products
Na atoms	2	2
H atoms	2	2
C atoms	2	2
O atoms	6	6

As you can see from the atom count, the equation is now balanced. Equations are completed by indicating the states of matter of each chemical species present, *i.e.* - (g) for gaseous substances, (s) for solids, (l) for liquids, and (aq) for species in solution in water.

$$2NaHCO_{3\,(s)} \rightarrow Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{\,(l)}$$

Example 3.1: *Balance the following equations:*

a) $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$

Start by balancing the element with the least atoms present, *i.e.* C, then the H atoms and lastly the O atoms.

Step 1: Balance C's

$$C_2H_6 + O_2 \rightarrow 2CO_2 + H_2O$$

Step 2: Balance H's

$$C_2H_6 + O_2 \rightarrow 2CO_2 + 3H_2O$$

Step 3: Balance O's

$$C_2H_6 + 7/2\ O_2 \rightarrow 2CO_2 + 3H_2O$$

Although the equation is balanced in terms of the atom count, the reaction coefficients need to be the lowest possible whole numbers. To get whole numbers we multiply all reaction coefficients by the lowest common denominator of any fractions. In this example we multiply by 2 to remove the 7/2 fraction.

Balanced equation: $\mathbf{2C_2H_6 + 7O_2 \rightarrow 2CO_2 + 3H_2O}$

b) $N_2 + H_2 \rightarrow NH_3$

Start by balancing the element with the least atoms present, *i.e.* N.

Step 1: Balance the N's. Since there are 2 N's on the left hand side of the equation and one on the right had side, the N's can be balanced by using the coefficient 2 in front of the NH_3.

$$N_2 + H_2 \rightarrow 2NH_3$$

Step 2: Balance the H's. Since there are 3 H's on the left hand side and 6 on the right hand side, the H's can be balanced by placing the coefficient 3 in front of H_2.

CHAPTER 4

Molarity & Concentration

Keywords: Calculations, Concentration, Molarity, Mole fraction, Parts per million, Percent by mass, Percent by volume.

4.1. CONCENTRATION

Most chemical processes occur in solution. "Life" has been described as the sum of a series of complex processes occurring in solution. Air, seawater, tea, beer and toilet bleach are all solutions. A solution is simply a homogenous mixture of substances of variable composition. The most abundant substance is called the **solvent**, whereas the substance present in lesser amounts is called the **solute**. If a small quantity of ethanol is added to water, the ethanol is the solute and the water is the solvent. But if we add a smaller amount of water to a larger amount of ethanol, then the water is the solute and the ethanol is the solvent. Although we will predominantly be concerned with solutions produced by solids dissolved in liquids, there are as many types of solutions as there are different combinations of solids, liquids, and gases. Brass is solid solution of copper and zinc, whilst the atmosphere is a solution in which a gaseous solvent (nitrogen) dissolves other gases (such as oxygen, carbon dioxide, water vapour, and neon).

Scientists often refer to concentrated solutions, dilute solutions, or very dilute solutions. Dilute solutions contain a relatively small amount of the solute in a given volume of solvent. Tap water is an example of a dilute solution since it contains small quantities of dissolved minerals. A concentrated solution on the other hand has a large amount of solute in the solvent.

Concentration can be expressed in several ways:

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Solution Type (solute - solvent)	Concentration Units	Concentration Equation
Solid-Liquid	Molarity (M or mol/L or mol dm^{-3})	Number of moles of solute per litre of solvent *
Solid-Liquid	Mass per volume (g/L or g dm^{-3})	Number of g per litre of solvent
Solid-Solid	Percent by Mass (m/m %) *	Mass of solute/ Mass of solution x 100
Liquid-Liquid	Percent by volume (v/v %)	Volume of solute / Volume of Solution x 100
Solid-Liquid	Percent by mass/volume (m/v %) **	Mass of Solute/ Volume of Solution x 100
All Solutions	Parts per million (ppm or mg /L or mg/kg)	Mass of solute/ Mass of Solution x 10^6

IMPORTANT NOTE:

* m/m % is often improperly referred to as weight percent (wt %) or weight-weight percent (w/w %)
** m/v % is often improperly referred to as % weight-volume percent (w/v %)
Since mass and weight are different quantities, both w/w % and w/v % are both incorrect.

4.2. MOLARITY

The concentration of a solution is expressed in terms of the amount of solute present in a standard volume of solvent. The standard volume is 1 litre, which is the same as is the same as 1 cubic decimetre (1 dm^3), which is the same as 1,000 cm^3. The most commonly used unit of concentration is molarity, often abbreviated as M. Molarity is simply the concentration of a solution expressed as the number of moles of solute per litre of solution.

A 1 M solution contains a molar mass (M_r) of solute in 1 litre of the solvent.

The following units of concentration are all the same:

$$M = mol/L = mol\ dm^{-3}$$

Concentration (c) = Amount of Solute (n) / Volume of solution (v)

$$c = n/v$$

Amount of Solute (n) = Volume of Solution (v) x Concentration (c)

$$n = v \times c$$

Volume of Solution (v) = Concentration (c) / Amount of Solute (n)

$$v = c/n$$

Concentration can also be expressed as stated above in terms of mass (g) of solute per unit of volume (litre), *i.e.* g/L.

Units:

V = Volume in litres (L)

Note: to convert cm^3 to L divide by 1000, *e.g* 10 cm^3 = 10/1000 = 0.01 L

If the amount of solute is measured in moles, then concentration unit is M
If amount of solute is measured in g then concentration unit is g/L or g dm^{-3}

For example, glucose, $C_6H_{12}O_6$ has a molar mass (M_r) of 180 g mol^{-1}. Therefore, a solution of 180 g of glucose dissolved in a total volume of 1 litre (L) has a concentration of 1.0 M or 180 g/L. A 0.1 M solution of glucose therefore must contain 18.0 g (0.1 x M_r) of dissolved glucose in 1 L of solution. Similarly 1.8 g of glucose in 100 cm^3 is equivalent to 18 g in 1000 cm^3 (1 L), giving a concentration of 0.1 M or 18 g/L.

To convert from M to g/L simply multiply by M_r
To convert g/L to M simply divide by M_r

Example 4.1: *What is the molarity of an aqueous solution of sodium chloride (NaCl) with a concentration of 5.85 g/L?*

Answer:

Molar Mass (M_r) of NaCl = 58.5
To convert g/L to M divide by M_r
Therefore, Molarity = 5.85/50.5 = **0.1 M**

Example 4.2: *A solution of $KMnO_4$ has a concentration of 0.2M. Express this concentration in terms of g/L.*

Answer:

Molar Mass (M_r) of $KMnO_4$ = 158 g mol^{-1}
To convert M to g/L multiply by M_r
Thus concentration = 0.2 x 158 = **31.6 g/L**

CHAPTER 5

Volumetric Analysis

Keywords: Acid-Base titrations, Back titrations, Calculations, Concentration, End point, Redox titration, Standard solutions, Volumetric analysis.

5.1 INTRODUCTION

A solution of accurately known concentration is called a **standard solution**. Standard solutions are used in volumetric (titrimetric) analysis, to determine the concentration an unknown solution. The volume of one solution that will react with a known volume of a standard solution (titrant) is determined. The point at which the exact amount of titrant added to just react with all of the other reagent present is called the **end point** or **equivalence point**. Indicators, normally added to the solution of known volume, which change colour, are often used to determine the end point.

Volumetric analysis is widely used to determine the concentration of a broad range variety of parameters including alkalinity, acidity, total hardness and chloride levels.

Titrations can be categorised based on chemical reactions:

> **Acid–Base** titrations involve the exact neutralisation of an acid or base with an acid or base of known concentration, thus allowing the concentration of the unknown acid or base solution to be determined.
> **Redox** titrations can be used to determine oxidizing or reducing agents in a solution. The reducing or oxidizing agent is used as the titrant against the other agent.
> **Back Titrations** are used where analytes are either partially soluble or too slow to give a reaction. A known amount of excess reagent is used. The remaining excess reagent is then titrated with another second reagent to determine how much of the excess reagent was used in the first

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titration, allowing the original analyte's concentration to be determined.

Example 5.1: *Chloride concentrations in water can be determined by titration against standardised silver nitrate solutions. A 50 cm^3 water sample was titrated against 0.05 M silver nitrate solution, using potassium chromate as the indicator. This indicator changes colour when all the chloride has been precipitated out of solution as silver chloride. This colour change occurred after the addition of 10.9 cm^3 of 0.05 M silver nitrate. What is the concentration of chloride ions in the water sample?*

Step 1: Write the balanced chemical equation, insert the information given in the question and identify what you are trying to calculate.

	$Cl^-_{(aq)}$	+	$AgNO_{3(aq)}$	→	$AgCl_{(s)} + NO_3^-{}_{(aq)}$
Reaction coefficients	1		1		
Volume (cm^3)	50 (0.05 L)		10.9		
No. of Moles					
Concentration (M)	?		0.05		

Please note that volumetric analysis (titrations) are predominantly concerned with the reaction coefficient (stoichiometric) relationship between the reactants. The products are included in the above table for the sake of completeness.

Step 2: If two pieces of information from volume, number of moles, and concentration are known for a given species, the third can be calculated using:

Concentration = Number of Moles / Volume of solution

c = n / v

OR

Number of Moles = Volume of Solution x Concentration

n = c x v

OR

Volume of Solution = Number of Moles / Concentration

v = n / c

Units: Concentration: M,
Volume: Litres (L) or {volume (cm^3)/1000}

Number of moles in 10.9 cm^3 0.05 M $AgNO_3$ = concentration x volume
= 0.05 x 10.9/1000 = **5.45 x 10^{-4}**

	$Cl^-_{(aq)}$	+	$AgNO_{3(aq)}$	→	$AgCl_{(s)} + NO^-_{3\,(aq)}$
Reaction Coefficients	1		1		
Volume (cm^3)	50 (0.05 L)		10.9		
No. of Moles			5.45 x 10^{-4}		
Concentration (M)	?		0.05		

Step 3: At the end point the amount of each species present is in accordance with the reaction coefficient (stoichiometric) relationship of the balanced chemical equation for the reaction. In this example, the number of moles of Cl^- in 50 cm^3 of tap water = the number of moles Ag^+ in 10.89 cm^3 of the standard solution = 5.45 x 10^{-4}

	$Cl^-_{(aq)}$	+	$AgNO_{3(aq)}$	→	$AgCl_{(s)} + NO^-_{3\,(aq)}$
Reaction Coefficients	1		1		
Volume (cm^3)	50 (0.05 L)		10.9		
No. of Moles	5.45 x 10^{-4}		5.45 x 10^{-4}		
Concentration (M)	?		0.05		

Step 4: Since we know both the number of moles of Cl^- and its volume, we can calculate its concentration.

Concentration of Cl^- = number of moles / volume (L)
= 5.45 x 10^{-4} /0.05
= 0.0109 M

	$Cl^-_{(aq)}$	+	$AgNO_{3(aq)}$	→	$AgCl_{(s)} + NO^-_{3\,(aq)}$
Reaction coefficients	1		1		
Volume (cm^3)	50 (0.05 L)		10.9		
No. of Moles	5.45 x 10^{-4}		5.45 x 10^{-4}		
Concentration (M)	0.0109		0.05		

CHAPTER 6

Molar Volumes & Reacting Volumes

Keywords: Calculations, Gases, Molar volume, Moles to volume, Reacting volumes.

6.1. MOLES & GASES

At Standard Temperature and Pressure (STP, 0ºC, 1 atmosphere pressure), one mole of gas occupies a volume of 22.4 litres (22.4 L; 22,400 cm^3). Thus at STP, 28g of nitrogen gas (N_2) occupies a volume of 22.4 litres, 32 g of oxygen (O_2) gas occupies a volume of 22.4 litres and 71 g of chlorine gas (Cl_2) occupies a volume of 22.4 litres. Similarly, 22.4 litres of hydrogen gas has a mass of 2 g and 22.4 litres of helium gas has a mass of 4 g at STP.

In other words, a given volume of gas always contains the same number of formula units (particles) regardless of the substance at a given temperature and pressure. But, the mass of one mole of gas (M_r in grams) differs from substance to substance.

ONE MOLE = M_r in grams = 6.02 x 10^{23} particles = 22.4 litres of a gas at STP

Given that the relative formula mass of CO_2 = 44 (12 +16 +16):

44g of CO_2 contains 6 x 10^{23} molecules of CO_2 which equals 1 mole and under STP occupies a volume of 22,400 cm^3 (22.4 L)

Thus,

0.5 moles of CO_2 has a mass of 22g, contains 3 x10^{23} molecules of CO_2 and occupies a volume of 11,200 cm^3 (11.2 litres) at STP

Equal volumes of gases at the same temperature and pressure contain the same number of moles.

Nigel P. Freestone

The molar volume (V_m) is simply defined as the **volume occupied by one mole of a gas**. The units of molar volume are L/mol (dm^3/mol or cm^3/mole).

At STP one mole of a gas occupies a volume of 22.4 L (22,400 cm^3). Molar volume at Room Temperature and Pressure (RTP, 25°C, 1 atmosphere) is 24.0 L (24,000 cm^3).

Molar Volume varies with both temperature and pressure.

Volume of gas = Number of moles x V_m
Number of moles = Volume of gas / V_m
V_m = Volume of Gas/Number of moles

Ensure that the volume units for V and V_m are the same, *i.e.* either L (dm^3) or cm^3.

6.2. MOLES TO VOLUME

Volume of gas = Number of moles x V_m

Example 6.1: *If the molar volume is 24 L/mol, what is the volume is occupied by 0.025 mole of H_2?*

Answer:

1 mole occupies a volume of 24 L
Volume of gas = Number of moles x V_m
Thus 0.25 moles will occupy 0.025 x 24 = 0.6 L = **600 cm^3**

Example 6.2: *Determine the volume occupied by 0.5 moles of chlorine gas given that the molar volume is 23,000 cm^3/mol.*

Answer:

V_m = 23,000 cm^3
Number of moles = 0.5
Volume of gas = Number of moles x V_m
Thus 0.5 moles will occupy 0.5 x 23000 = 11,500 cm^3 = **11.5 L**

Exercise 6.1

Calculate the volume occupied by the following gases at STP (V_m = 22.4 litres)

a. *2 mole of CH_4*
b. *0.3 moles of NH_3*
c. *1.6 moles of C_2H_4*
d. *3 moles of SO_2*
e. *0.26 moles of NO*
f. *5.7 moles of HBr*
g. *0.22 moles of Cl_2*
h. *0.020 moles of CO_2*
i. *15 moles of O_2*
j. *3.5×10^{-3} moles of H_2*

6.3. VOLUME TO MOLES (AND MASS)

$$\text{Number of moles} = \text{Volume} / V_m$$
$$\text{Mass} = \text{Number of moles} \times M_r$$

Example 6.3: *Calculate the number of moles of carbon dioxide in 250 cm^3 of the gas. Assume that one mole of a gas occupies a volume of 23,000 cm^3*

Answer:

V_m = 23 L = 23,000 cm^3
Number of moles = Volume / V_m
Thus number of moles present in 250 cm^3 = 250/23000 = **0.011 moles**

Example 6.4: *What is the mass 5 litres of oxygen gas at STP?*

Answer:

V_m (at STP) = 22.4 L = 22,400 cm^3
$M_r[O_2]$ = 32
Number of moles = Volume of gas / V_m
Number of moles present in 5 L = 5/22.4 = 0.223 moles
Mass of O_2 = number of moles x M_r

Mole Driving Test No. I

M Counting Moles .... Driving Test M

To Pass the Mole Driving Test and throw away your **M**-plates you will need to achieve a score of greater than 70%. Each question is equally weighted (5%).

Relative Atomic Masses: Hydrogen (H) = 1, Carbon (C) = 12, Nitrogen (N) = 14, Oxygen (O) = 16, Sodium (Na) = 23, Sulfur (S) = 32, Chlorine (Cl) = 35.5, Potassium (39), Calcium (Ca) = 40, Copper (Cu) = 63.5, Barium (Ba) = 137, Mercury(Hg) 200.5, Lead (Pb) = 207.
Molar Volume: 22.4 L or 22,4000 cm^3 at STP (273.15 K and 1 atm)
Avogadro's Number: 6.02 x 10^{23}

1. Rubidium has two naturally occurring two isotopes. 72.2% of all rubidium atoms have a relative atomic mass of 85 and the rest have a relative atomic mass of 87. What is the average relative atomic mass of rubidium?

2. Compound A was found to have the following composition: 47.0 % potassium, 14.5 % carbon, and 38.5 % oxygen and a relative molecular molar mass of 166.22. Determine the empirical and the molecular formula of compound A.

3. Lead (IV) oxide reacts with concentrated hydrochloric acid as follows:

$$PbO_{2(s)} + 4HCl_{(aq)} \rightarrow PbCl_{2(s)} + Cl_{2(g)} + 2H_2O_{(l)}$$

Calculate the mass of both lead chloride and chlorine gas that could be produced from 37.2g of PbO_2.

4. In 1774 Joseph Priestly produced oxygen by heating a sample of mercury II oxide with a large lens, according to the following equation:

$$2HgO(s) \rightarrow 2Hg(l) + O_{2(g)}$$

Calculate the volume of O_2 that could be generated from 1.08g of mercury (II) oxide. Assume that 1 mole of a gas occupies 24 dm^3 under the experimental conditions.

5. Copper (II) nitrate thermally decomposes:

$$2Cu(NO_3)_{2(s)} \rightarrow 2CuO_{(s)} + 4NO_{2(g)} + O_{2(g)}$$

   a. What mass of copper (II) oxide could be produced from the thermal decomposition of 20.0 g of copper (II) nitrate?

   b. Calculate the mass of NO_2 produced?

6. Calculate the concentration of a solution containing 0.732 moles of ammonia in 250 cm^3 of solution.

7. Sodium hydride reacts violently with water forming flammable hydrogen gas:

$$NaH_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2\,(g)}$$

A 1.00 g sample of sodium hydride was added to water and the resulting solution was diluted to a volume of exactly 250 cm^3

   a. What is the concentration sodium hydroxide solution formed?

   b. What volume of hydrogen gas, measured at STP, was generated?

   c. What volume of 0.112 M hydrochloric acid would be required to react exactly with a 25.0 cm^3 sample of the sodium hydroxide solution?

8. The hydrates of sodium carbonate can be represented by the general formula $Na_2CO_3.xH_2O$. A student dissolved a 3.01 g sample of one of these hydrates in water and made the solution up to 250 cm^3. In a titration, a 25.0 cm^3 aliquot of this solution required 24.3 cm^3 of 0.200 M hydrochloric acid for complete reaction.

The equation for this reaction is shown below.

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$$

   a. Calculate the number of moles of HCl in 24.3 cm^3 of 0.200 M hydrochloric acid.

   b. Determine the number of moles of Na_2CO_3 present in 25.0 cm^3 of the Na_2CO_3 solution.

c. Deduce the number of moles of Na_2CO_3 in the original 250 cm^3 of solution.

d. Calculate the relative formula mass (M_r) of the hydrated sodium carbonate and hence the value of x.

9. A student prepared a calcium hydroxide solution by adding 0.00131 mole of calcium to a beaker containing about 100 cm^3, according to the reaction:

$$Ca_{(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$$

a. Calculate the mass of calcium that the student added to the beaker of water.

b. What volume of hydrogen gas, at STP, would be generated?

c. The contents of the beaker were transferred to a 250 cm^3 volumetric flask and water was added to make the solution up to 250 cm^3. What is the concentration of hydroxide ions in the volumetric flask?

10. What is the volume of oxygen gas, at STP, required for the complete combustion of 1.0 kg of methane.

11. A 5.175 g sample of lead when heated in air at 300°C produced 5.708 g of an oxide of lead as the only product. Write a balanced chemical equation for this reaction.

12. 4.90 g of pure sulfuric acid was dissolved in water and the resulting solution was made up to 200 cm^3. On titration, a 10.0 cm^3 aliquot of a sodium hydroxide solution was completely neutralized by 20.7 cm^3 of this solution.

What is the concentration of the sodium hydroxide solution?

Mole Driving Test No. II

To Pass the Mole Driving Test and throw away your **M**-plates you will need to achieve a score of greater than 70%. Each question is equally weighted (5%).

Relative Atomic Masses: Hydrogen (H) = 1, Carbon (C) = 12, Nitrogen (N) = 14, Oxygen (O) = 16, Sodium (Na) = 23, Magnesium (Mg = 24, Silicon (Si) = 28, Sulfur (S) = 32, Chlorine (Cl) = 35.5, Lead (Pb) = 207
Molar Volume: 22.4 L (22,400 cm^3) at STP (273.15 K and 1 atm)
Avogadro's Number: 6.02 x 10^{23}

1. Calculate the average relative atomic mass of titanium given that it has five common isotopes, ^{46}Ti (8.0%), ^{47}Ti (7.8%), ^{48}Ti (73.4%), ^{49}Ti (5.5%), ^{50}Ti (5.3%)

2. Calculate the empirical and molecular formula of a compound of containing 17.04% Na, 47.41% S, with a relative formula mass 270.

3. In the sixteenth century, a large deposit of graphite was discovered in the Lake District. People at the time thought that the graphite was a form of lead. Graphite used in pencils is still referred to as 'pencil lead'. A student found that the mass of the 'pencil lead' in a school pencil was 0.321 g.

   a. How many moles of carbon atoms are present in 0.321g of student's pencil lead? Assume that the 'pencil lead' is pure graphite.

   b. Determine the number carbon atoms are present in the student's 'pencil lead'.

4. 19.6 g of hydrogen chloride, HCl, was dissolved in water and the volume made up to 250 cm^3. What is the concentration of the resulting solution?

5. Determine the mass of H_2O required to completely react with 5.0 g of $SiCl_4$:

$$SiCl_4 + 2H_2O \rightarrow SiO_2 + 4HCl$$

6. Magnesium carbonate and hydrochloric acid react according to the following equation:

$$MgCO_3 + 2HCl \rightarrow MgCl_2 + H_2O + CO_2$$

When a student added 75.0 cm^3 of 0.500 M hydrochloric acid to 1.25 g of impure $MgCO_3$ some acid was left unreacted. 21.6 cm^3 of a 0.500 M solution of sodium hydroxide was required to neutralize the unreacted acid.

a. How many moles of HCl are present in 75.0 cm^3 of 0.500 M hydrochloric acid?

b. How many moles of NaOH were used to neutralize the unreacted HCl.

c. Determine the number of moles of HCl that reacted with the $MgCO_3$ in the sample.

d. Calculate both the number of moles and the mass of $MgCO_3$ in the sample, and hence its purity.

7. The explosive nitroglycerine, $C_3H_5N_3O_9$, decomposes rapidly on detonation to form a large volume of gas, according to the following equation:

$$4C_3H_5N_3O_{9(l)} \rightarrow 12CO_{2(g)} + 10H_2O_{(g)} + 6N_{2(g)} + O_{2(g)}$$

0.350 g of oxygen gas was produced from the detonation of a sample of nitroglycerine.

a. Determine the number of moles of oxygen gas produced from the detonation of nitroglycerine, and the total number of moles of gas generated.

b. Determine the number of moles, and the mass, of nitroglycerine detonated.

8. Lead(II) sulfate can be produced from the reaction between lead nitrate and dilute sulfuric acid. What is the maximum amount of lead sulfate that could be obtained from 10 g of lead nitrate dissolved in water?

$$Pb(NO_3)_{2(aq)} + H_2SO_{4(aq)} \rightarrow PbSO_{4(s)} + 2HNO_{3(aq)}$$

9. The reaction between ammonium sulfate and aqueous sodium hydroxide is shown by the equation below.

$$(NH_4)_2SO_4 + 2NaOH \rightarrow 2NH_3 + Na_2SO_4 + 2H_2O$$

A sample of ammonium sulfate was heated with 100 cm^3 of 0.500 M aqueous sodium hydroxide. An excess of sodium hydroxide was used to ensure that all of the ammonium sulfate reacted. The unreacted sodium hydroxide required 27.3 cm^3 of 0.600 M hydrochloric acid for neutralisation.

a. Calculate the original number of moles of NaOH in 100 cm^3 of 0.500 M aqueous sodium hydroxide.

b. How many moles of HCl are present in 27.3 cm^3 of 0.600 M hydrochloric acid?

c. Deduce the number of moles of the unreacted NaOH neutralized by the hydrochloric acid.

d. How many moles of NaOH reacted with the ammonium sulfate?

e. Determine the number of moles and the mass of ammonium sulfate in the sample.

10. The purity of commercially available sodium hydrogencarbonate was tested as follows. A 0.400g sample was dissolved in 100.0 cm^3 of water and titrated against 0.200 M hydrochloric acid using methyl orange indicator.

$$NaHCO_3 + HCl \rightarrow NaCl + CO_2 + H_2$$

23.75 cm^3 of acid was required for complete neutralisation.

a. How many moles of acid were used in the titration?

b. Calculate the mass of sodium hydrogen carbonate titrated and hence the purity of the sample.

Mole Driving Test No. III

Counting Moles ....

Driving Test

To Pass the Mole Driving Test and Throw Away Your **M**-plates You Will Need to Achieve a Score of Greater than 70%. Each Question is Equally Weighted (5%).

Relative Atomic Masses: Hydrogen (H) = 1, Carbon (C) = 12, Nitrogen (N) = 14, Oxygen (O) = 16, Sodium (Na) = 23, Magnesium (Mg) = 24, Silicon (Si) = 28, Sulfur (S) = 32, Chlorine (Cl) = 35.5, Potassium (K) = 39, Titanium (Ti) = 48, Barium (Ba) = 137
Molar Volume: 22.4 L (22,4000 cm^3) at STP (273.15 K and 1 atm)
Avogadro's Number: 6.02 x 1023 mol^{-1}

1. Calculate the mass, in grams, of a single atom of sodium-23.

2. An organic compound was found to contain 12.8% carbon, 2.13% hydrogen and 85.07% bromine. If the compound has a relative molecular mass of 188, determine its empirical formula and the molecular formula.

3. What the mass of phosphorus is required to produce 200 g of phosphine (PH_3)?

$$P_{4(s)} + 3NaOH_{(aq)} + 3H_2O_{(l)} \rightarrow 3NaH_2PO_{4(aq)} + PH_{3(g)}$$

4. A student heats 5.29g of $Sr(NO_3)_2$ and collects the gas.

$$2Sr(NO_3)_{2(s)} \rightarrow 2SrO_{(s)} + 4NO_{2(g)} + O_{2(g)}$$

Determine the volume of gas, at STP, obtained by the student. Molar mass of $Sr(NO_3)_2$ = 211.6 g mol^{-1}.

5. A 12.41 g sample of hydrated sodium thiosulfate, $Na_2S_2O_3$•$5H_2O$, was heated to remove the water of crystallization.

    a. What is the relative formula mass of $Na_2S_2O_3$•$5H_2O$?

b. Calculate the expected mass of anhydrous sodium thiosulfate that forms.

6. The Kroll process is used to convert ore into titanium metal. Titanium chloride produced from the chlorination of ore, is reduced to titanium metal using magnesium under an inert atmosphere.

$$2Mg + TiCl_4 \rightarrow 2MgCl_2 + Ti$$

Calculate the maximum mass of titanium that could be produced from the addition of 3800 kg of titanium chloride to 1500 kg of magnesium.

7. A 1 mg sample of octane, C_8H_{18} was totally combusted in air.

   a. How many moles are present in 1 mg of octane ?

   b. Determine the number of moles and the volume of carbon dioxide generated.

8. Metal **M** forms a carbonate (M_2CO_3), which reacts with hydrochloric acid according to the following equation:

$$M_2CO_3 + 2HCl \rightarrow 2MCl + CO_2 + H_2O$$

A 0.394 g sample of M_2CO_3, was found to require the addition of 21.7 cm^3 of a 0.263 M solution of hydrochloric acid (HCl) for complete reaction.

   a. Determine the number of moles of hydrochloric acid added.

   b. Determine the relative molecular mass of M_2CO_3 and hence the identity of M

9. An ammonia solution was reacted with sulfuric acid:

$$2NH_{3(aq)} + H_2SO_{4(aq)} \rightarrow (NH_4)_2SO_{4(aq)}$$

A 25.0 cm^3 aliquot of 1.24 M sulfuric acid required 30.8 cm^3 of this ammonia solution for complete reaction.

Calculate the concentration of the ammonia and the mass of ammonium sulfate present in the solution at the end of this titration.

10. Hydrogen is produced by the addition of hydrochloric acid to magnesium metal:

$$2HCl + Mg \rightarrow MgCl_2 + H_2$$

What mass of hydrogen is produced from the addition of 100 cm^3 of 5M hydrochloric acid of to an excess of magnesium ?

11. Ammonium nitrate is produced industrially by the reaction between ammonia and nitric acid:

$$NH_3 + HNO_3 \rightarrow NH_4NO_3$$

Calculate the volume of 2M nitric acid required to react with exactly 20.0 g of ammonia.

12. Potassium chlorate, $KClO_3$, thermally decomposes according to the following equation:

$$2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$$

   a. What mass of oxygen could be produced from the complete decomposition of 1.47 g of $KClO_3$?

   b. What mass of $KClO_3$ is required to generate 1.00 dm^3 of oxygen at STP ?

13. 25.0 cm^3 of 0.25 M sodium hydroxide required 22.5 cm^3 of a hydrochloric acid solution for complete neutralisation. Calculate the concentration of the HCl solution.

14. Barium nitrate thermally decomposes as follows:

$$Ba(NO_3)_{2(s)} \rightarrow BaO_{(s)} + 2NO_{2(g)} + \tfrac{1}{2}\, O_{2(g)}$$

   a. What is the total volume of gas, at STP, generated by the thermal decomposition of 5.00 g of barium nitrate ?

   b. What volume of 1.20 M hydrochloric acid is required to neutralize the barium oxide produced by the thermal decomposition of 5.00 g of barium nitrate. Barium oxide reacts with hydrochloric acid as follows:

$$BaO_{(s)} + 2HCl_{(aq)} \rightarrow BaCl_{2(aq)} + H_2O_{(l)}$$

15. Silver nitrate thermally decomposes. 0.720 g of silver and 0.307 g of nitrogen dioxide was produced from heating a 1.133 g sample of silver nitrate in an open tube. The rest of the mass loss was due to oxygen. Using this data write the balanced chemical equation for the reaction.

Mole Driving Test Answers

M Counting Moles ....
Driving Test M

To Pass the Mole Driving Test and Throw Away Your **M**-plates You Will Need To Achieve a Score of Greater than 70%. Each Question is Equally Weighted (5%).

Relative Atomic Masses:
Hydrogen (H) = 1, Carbon (C) = 12, Nitrogen (N) = 14, Oxygen (O) = 16, Sodium (Na) = 23, Sulfur (S) = 32, Chlorine (Cl) = 35.5, Potassium (39), Calcium (Ca) = 40, Copper (Cu) = 63.5, Barium (Ba) = 137, Mercury(Hg) 200.5, Lead (Pb) = 207.
Molar Volume: 22.4 L or 22,4000 cm^3 at STP (273.15 K and 1 atm)
Avogadro's Number: 6.02 x 10^{23}

1. Rubidium has two naturally occurring two isotopes. 72.2% of all rubidium atoms have a relative atomic mass of 85 and the rest have a relative atomic mass of 87. What is the average relative atomic mass of rubidium?

Answer:

Average Atomic Mass=(72.2/100 x 85) + (27.8/100 x 87)
=61.37 + 24.186
=**85.56**

2. Compound A was found to have the following composition: 47.0 % potassium, 14.5 % carbon, and 38.5 % oxygen and a relative molecular molar mass of 166.22. Determine the empirical and the molecular formula of compound A.

Answer:

	Potassium	Carbon	Oxygen
% composition	47.0	14.5	38.5
A_r	39	12	16
% composition/A_r	47/39 = 1.2	14.5/12 = 1.2	38.5/16 = 2.4

Table contd.....

	Potassium	**Carbon**	**Oxygen**
Ratio	1	1	2

Empirical formula: KCO_2
Relative formula mass KCO_2 = (1 x 39) + 12 + (2 x 16) = 83
Molecular formula = $(KCO_2)_n$
n = 166/83 = 2
Molecular formula: $(KCO_2)_2$ = $\mathbf{K_2C_2O_4}$

3. Lead (IV) oxide reacts with concentrated hydrochloric acid as follows:

$$PbO_{2(s)} + 4HCl_{(aq)} \rightarrow PbCl_{2(s)} + Cl_{2(g)} + 2H_2O_{(l)}$$

Calculate the mass of both lead chloride and chlorine gas that could be produced from 37.2 g of PbO_2.

Answer:

	$PbO_{2(s)}$	+	**$4HCl_{(aq)}$**	→	**$PbCl_{2(s)}$**	+	**$Cl_{2(g)}$**	+	**$2H_2O_{(l)}$**
M_r	239		36.5		278		71		18
Mass Balance	239		146		278		71		36
		385				385			
Reaction Coefficients	1		4		1		1		1
Mass (g)	37.2				0.1556 x 278 = **43.27**		0.1556 x 71 = **11.05**		
No. of Moles	0.1556				0.1556		0.1556		

Number of moles in 37.2 g of PbO_2 = mass/M_r = 347.2/239 = 0.1556
According to the balanced chemical equation, 1 mole of PbO_2 produces 1 mole $PbCl_2$
Therefore, 0.1556 moles of PbO_2 can produce 0.1556 moles of $PbCl_2$
Mass of 0.1556 moles of $PbCl_2$ = number of moles x M_r = 0.1556 x 278 = 43.27 g
Mass of $PbCl_2$ = **43.27 g**

According to the balanced chemical equation, 1 mole of PbO_2 produces 1 mole of Cl_2
Therefore, 0.1556 moles of PbO_2 can produce 0.1556 moles of Cl_2
Mass of 0.01556 moles of Cl_2 = number of moles x M_r = 0.1556 x 71 = 11.05 g
Mass of Cl_2 = **11.05 g**

4. In 1774 Joseph Priestly produced oxygen by heating a sample of mercury II oxide with a large lens, according to the following equation:

$$2HgO(s) \rightarrow 2Hg(l) + O_{2(g)}$$

Calculate the volume of O_2 that could be generated from 1.08g of mercury (II) oxide. Assume that 1 mole of a gas occupies 24 dm^3 under the experimental conditions.

Answer:

	$2HgO_{(s)}$	→	**$2Hg_{(l)}$**	+	**$O_{2(g)}$**
A_r/M_r	216.5		200.5		32
Mass Balance	433		401		32
				433	
Reaction Coefficients	2		2		1
Mass	1.08				
No. of Moles	1.08/216.5 = 0.005		0.005		0.0025
Volume (L)					0.025 x 22.4 = **0.06**

Number of moles in 1.08 g of HgO = 1.08/216.5 = 0.005
According to the balanced chemical equation 2 moles of HgO generates 1 mole of O_2
Thus number of moles of O_2 produced from 0.005 moles of HgO = 0.005/2 = 0.0025
Volume of 0.0025 moles of O_2 = 0.0025 x 24 = **0.06 L** = **60 cm^3**

5. Copper (II) nitrate thermally decomposes:

$$2Cu(NO_3)_{2(s)} \rightarrow 2CuO_{(s)} + 4NO_{2(g)} + O_{2(g)}$$

a) What mass of copper (II) oxide could be produced from the thermal decomposition of 20.0 g of copper (II) nitrate?

Answer:

	$2Cu(NO_3)_{2(s)}$	→	**2CuO(s)**	+	**$4NO_{(g)}$**	+	**$O_{2(g)}$**
M_r	187.5		79.5		46		32
Mass Balance			159		184		32
	375			375			

PERIODIC TABLE

The Periodic Table of Elements

Key

relative atomic mass	7
	Li lithium
atomic number	3

1	2											3	4	5	6	7	0
1 **H** hydrogen 1																	4 **He** helium 2
7 **Li** lithium 3	9 **Be** beryllium 4											11 **B** boron 5	12 **C** carbon 6	14 **N** nitrogen 7	16 **O** oxygen 8	19 **F** fluorine 9	20 **Ne** neon 10
23 **Na** sodium 11	24 **Mg** magnesium 12											27 **Al** aluminium 13	28 **Si** silicon 14	31 **P** phosphorus 15	32 **S** sulphur 16	35.5 **Cl** chlorine 17	40 **Ar** argon 18
39 **K** potassium 19	40 **Ca** calcium 20	45 **Sc** scandium 21	48 **Ti** titanium 22	51 **V** vanadium 23	52 **Cr** chromium 24	55 **Mn** manganese 25	56 **Fe** iron 26	59 **Co** cobalt 27	59 **Ni** nickel 28	63.5 **Cu** copper 29	65 **Zn** zinc 30	70 **Ga** gallium 31	73 **Ge** germanium 32	75 **As** arsenic 33	79 **Se** selenium 34	80 **Br** bromine 35	84 **Kr** krypton 36
85 **Rb** rubidium 37	88 **Sr** strontium 38	89 **Y** yttrium 39	91 **Zr** zirconium 40	93 **Nb** niobium 41	96 **Mo** molybdenum 42	**Tc** technetium 43	101 **Ru** ruthenium 44	103 **Rh** rhodium 45	106 **Pd** palladium 46	108 **Ag** silver 47	112 **Cd** cadmium 48	115 **In** indium 49	119 **Sn** tin 50	122 **Sb** antimony 51	128 **Te** tellurium 52	127 **I** iodine 53	131 **Xe** xenon 54
133 **Cs** caesium 55	137 **Ba** barium 56	139 **La** lanthanum 57 ❶	178 **Hf** hafnium 72	181 **Ta** tantalum 73	184 **W** tungsten 74	186 **Re** rhenium 75	190 **Os** osmium 76	192 **Ir** iridium 77	195 **Pt** platinum 78	197 **Au** gold 79	201 **Hg** mercury 80	204 **Tl** thallium 81	207 **Pb** lead 82	209 **Bi** bismuth 83	**Po** polonium 84	**At** astatine 85	**Rn** radon 86
Fr francium 87	226 **Ra** radium 88	227 **Ac** actinium 89 ❷															

❶ 58 – 71 Lanthanum series	140 **Ce** cerium 58	141 **Pr** praseodymium 59	144 **Nd** neodymium 60	**Pm** promethium 61	150 **Sm** samarium 62	152 **Eu** europium 63	157 **Gd** gadolinium 64	159 **Tb** terbium 65	162 **Dy** dysprosium 66	165 **Ho** holmium 67	167 **Er** erbium 68	169 **Tm** thulium 69	173 **Yb** ytterbium 70	175 **Lu** lutetium 71
❷ 90 – 103 Actinium series	232 **Th** thorium 90	**Pa** protoactinium 91	238 **U** uranium 92	**Np** neptunium 93	**Pu** plutonium 94	**Am** americium 95	**Cm** curium 96	**Bk** berkelium 97	**Cf** californium 98	**Es** einsteinium 99	**Fm** fermium 100	**Md** mendelevium 101	**No** nobelium 102	**Lr** lawrencium 103

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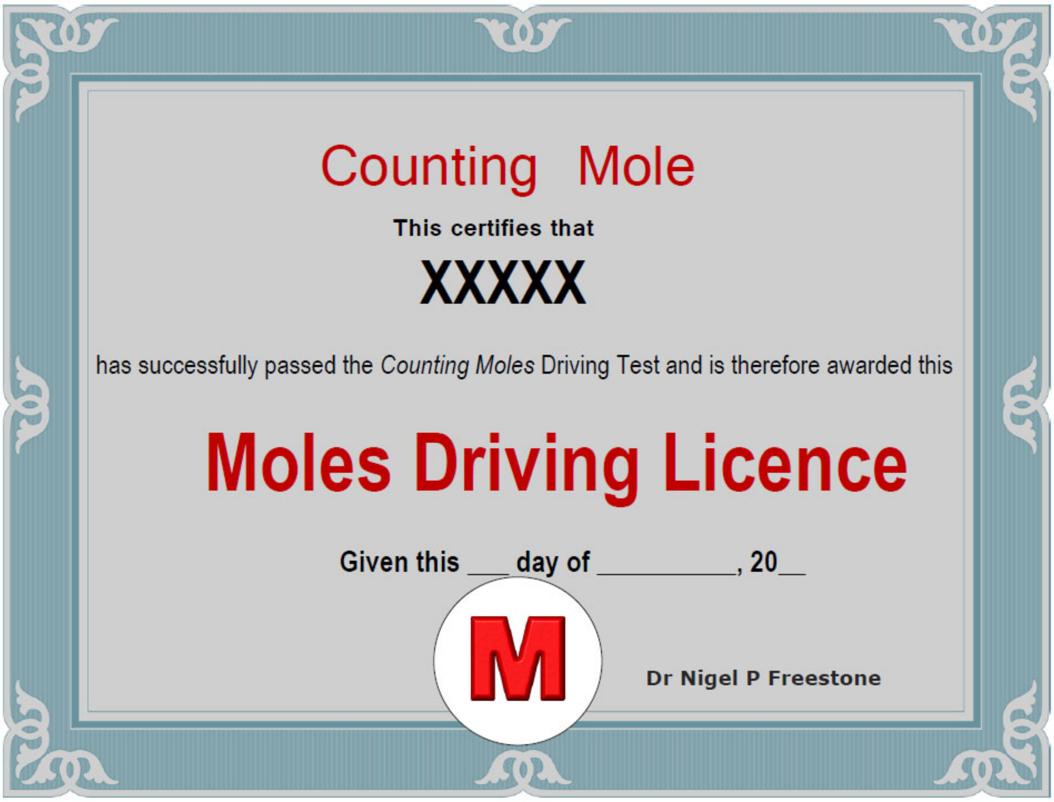
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Dr Nigel P Freestone

SUBJECT INDEX

Nigel P. Freestone