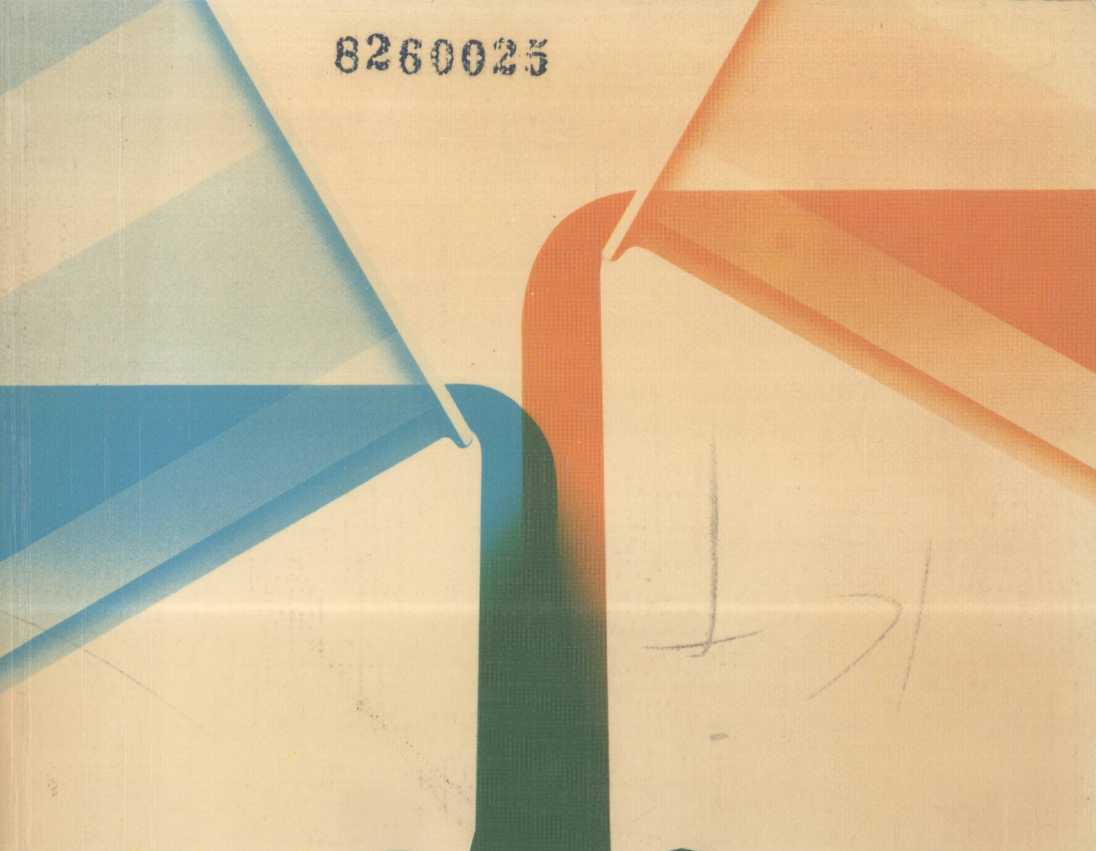

Calculations for O-level Chemistry

E. N. Ramsden

CALCULATIONS FOR O-LEVEL CHEMISTRY

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Preface

This book has been written to provide practice in calculations for O-Level Chemistry students. All the types of numerical problems met in GCE O-level syllabuses are covered. A brief treatment of the theoretical background to each type of numerical problem is given, and is followed by a series of worked examples. The problems are divided into three sections of increasing difficulty, the third section being a collection of questions from past O-level papers. The arithmetic in Sections 1 and 2 has been kept simple by basing the problems on a selection of compounds with relative formula masses which are round numbers, such as NH_4NO_3 = 80, $MgSO_4$ = 120, $CaBr_2$ = 200. If a pupil has difficulty with a problem, he or she can return for help to the theoretical section and to the worked examples. Thus, the book can be used for private study, as well as for class work.

The concept of the mole is the thread which knits together the calculations on reacting masses of solids, reacting volumes of gases, empirical formulae, volumetric analysis, electrolysis and heats of reaction. The pupil learns to look at the equation and ask himself or herself how many moles of reactant are involved.

The Association for Science Education publication, *Chemical Nomenclature, Symbols and Terminology in School Science* (2nd edn, 1979), which embodies the latest recommendations of the International Union of Pure and Applied Chemistry, is followed in matters of terminology. Concentrations are expressed as the number of moles of solute per cubic decimetre of solution (mol dm^{-3}). The older term, 'molarity', is also explained as it continues to be used in schools and in some examination papers. In addition to the units recommended by IUPAC and ASE, millimetres of mercury and degrees Celsius are used for pressure and temperature. Reference is made to the obsolescent unit of charge, the Faraday.

Some Examination Boards use terms which differ from those recommended by IUPAC, and students should familiarise themselves with the terminology used by their own Board. Reference to the questions from past papers will assist them in this.

Acknowledgements

I thank the following Examination Boards for permission to print questions from recent examination papers. In a few cases, the wording of the questions has been changed slightly, for example, to ask readers to plot results on a piece of graph paper, instead of on the question paper. The Boards are not responsible for the accuracy of the numerical answers.

The Associated Examining Board (at whose request the terminology of some questions has been updated)
University of Cambridge Schools Local Examinations Syndicate
Joint Matriculation Board
Oxford and Cambridge Schools Examination Board
Oxford Delegacy of Local Examinations
Southern Universities' Joint Board
University of London School Examinations Council
Welsh Joint Education Committee

My thanks are also offered to those who have helped me during the preparation of this book, including the pupils who have read the text and given me the benefit of their comments. I am grateful to Chris Baker for helpful discussions and to Stephanie Cox for checking the numerical answers. I thank Stanley Thornes (Publishers) for the care they have taken over the preparation of the text. Finally, I thank my family for their support and encouragement.

E N Ramsden
Hull, 1981

Contents

1. Formulae and Equations

Calculations are based on formulae and on equations. In order to tackle the calculations in this book, you will have to be quite sure you can work out the formulae of compounds correctly, and that you can balance equations. This section is a revision of work on formulae and equations.

Formulae

Electrovalent compounds consist of oppositely charged ions. The compound formed is neutral because the charge on the positive ion (or ions) is equal to the charge on the negative ion (or ions). In sodium chloride, NaCl, one sodium ion, Na^+, is balanced in charge by one chloride ion, Cl^-.

This is how the formulae of electrovalent compounds can be worked out

Compound	*Zinc chloride*
Ions present are	Zn^{2+} and Cl^-
Now balance the charges	One Zn^{2+} ion needs two Cl^- ions
Ions needed are	Zn^{2+} and $2Cl^-$
The formula is	$ZnCl_2$
Compound	*Sodium sulphate*
Ions present are	Na^+ and SO_4^{2-}
Now balance the charges	Two Na^+ balance one SO_4^{2-}
Ions needed are	$2Na^+$ and SO_4^{2-}
The formula is	Na_2SO_4
Compound	*Aluminium sulphate*
Ions present are	Al^{3+} and SO_4^{2-}
Now balance the charges	Two Al^{3+} balance three SO_4^{2-}
Ions needed are	$2Al^{3+}$ and $3SO_4^{2-}$
The formula is	$Al_2(SO_4)_3$
Compound	*Iron(II) sulphate*
Ions present are	Fe^{2+} and SO_4^{2-}
Now balance the charges	One Fe^{2+} balances one SO_4^{2-}
Ions needed are	Fe^{2+} and SO_4^{2-}
The formula is	$FeSO_4$

Compound	*Iron(III) sulphate*
Ions present are	Fe^{3+} and SO_4^{2-}
Now balance the charges	Two Fe^{3+} balance three SO_4^{2-}
Ions needed are	$2Fe^{3+}$ and $3SO_4^{2-}$
The formula is	$Fe_2(SO_4)_3$

You need to know the charges of the ions in Table 1.1. Then you can work out the formula of any electrovalent compound.

You will notice that the compounds of iron are named iron(II) sulphate and iron(III) sulphate to show which of its valencies iron is using in the compound. This is always done with the compounds of elements of variable valency.

Table 1.1 *Symbols and valencies of common ions*

Name	Symbol	Valency	Name	Formula	Valency
Hydrogen	H^+	1	Hydroxide	OH^-	1
Ammonium	NH_4^+	1	Nitrate	NO_3^-	1
Potassium	K^+	1	Chloride	Cl^-	1
Sodium	Na^+	1	Bromide	Br^-	1
Silver	Ag^+	1	Iodide	I^-	1
Copper(I)	Cu^+	1	Hydrogen-carbonate	HCO_3^-	1
Barium	Ba^{2+}	2	Oxide	O^{2-}	2
Calcium	Ca^{2+}	2	Sulphide	S^{2-}	2
Copper(II)	Cu^{2+}	2	Sulphite	SO_3^{2-}	2
Iron(II)	Fe^{2+}	2	Sulphate	SO_4^{2-}	2
Lead	Pb^{2+}	2	Carbonate	CO_3^{2-}	2
Magnesium	Mg^{2+}	2			
Zinc	Zn^{2+}	2			
Aluminium	Al^{3+}	3	Phosphate	PO_4^{3-}	3
Iron(III)	Fe^{3+}	3			

The formulae of covalent compounds

To work out the formulae of covalent compounds, you need to know the symbols and the valencies of the elements present. These are listed in Table 1.2. The valency is the number of electrons which an element shares in forming a compound. The method of working out the formulae is the same as for electrovalent compounds, although here electrons are shared, not given and accepted.

Table 1.2 *Symbols and valencies of some common elements*

Element	Symbol	Valency	Element	Symbol	Valency
Bromine	Br	1	Iodine	I	1
Carbon	C	4	Nitrogen	N	3 and 5
Chlorine	Cl	1	Oxygen	O	2
Fluorine	F	1	Phosphorus	P	3 and 5
Hydrogen	H	1	Sulphur	S	2, 4 and 6

Method of working out the formulae of covalent compounds

Compound	*Compound of carbon and hydrogen*
Symbols of elements	C H
Valencies (no. of shared electrons)	4 1
Balance the electrons	One C with four e^- needs four H with one e^-
Atoms needed	C and 4 H
The formula is	CH_4
	You will recognise the formula of methane.

Compound	*Compound of sulphur and hydrogen*
Symbols of elements	S H
Valencies (no. of shared electrons)	2 1
Balance the electrons	One S with two e^- needs two H with one e^-
Atoms needed	S and 2 H
The formula is	H_2S
	This is the formula of hydrogen sulphide.

Equations

Having symbols for elements and formulae for compounds gives us a way of representing chemical reactions.

Example 1 Instead of writing, 'Copper carbonate dissociates into copper oxide and carbon dioxide', we can write

$$CuCO_3 \rightarrow CuO + CO_2$$

The atoms we finish with are the same in number and kind as the atoms we start with. We start with one atom of copper, one atom of carbon and three atoms of oxygen, and we finish with the same. This makes the two sides of the expression equal, and we call it an *equation*. A simple way of conveying a lot more information is to include *state symbols* in the equation. These are (s) = solid, (l) = liquid, (g) = gas, (aq) = in solution in water. The equation

$$CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$$

tells you that solid copper carbonate dissociates to form solid copper oxide and the gas carbon dioxide.

Example 2 The equation

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

tells you that solid zinc reacts with a solution of sulphuric acid to give a solution of zinc sulphate and hydrogen gas. Hydrogen is written as H_2, since each molecule of hydrogen gas contains two atoms.

Example 3 Sodium carbonate reacts with dilute hydrochloric acid to give carbon dioxide and a solution of sodium chloride. The equation could be

$$Na_2CO_3(s) + HCl(aq) \rightarrow CO_2(g) + NaCl(aq) + H_2O(l)$$

but, when you add up the atoms on the right, you find that they are not equal to the atoms on the left. The equation is not 'balanced', so the next step is to balance it. Multiplying NaCl by two gives

$$Na_2CO_3(s) + HCl(aq) \rightarrow CO_2(g) + 2NaCl(aq) + H_2O(l)$$

This makes the number of sodium atoms on the right-hand side equal to the number on the left-hand side. But there are two chlorine atoms on the right-hand side, therefore the HCl must be multiplied by two:

$$Na_2CO_3(s) + 2HCl(aq) \rightarrow CO_2(g) + 2NaCl(aq) + H_2O(l)$$

The equation is now balanced.

When you are balancing a chemical equation, the only way you do it is to multiply the number of atoms or molecules. You never try to alter a formula. In the above example, you got two chlorine atoms by multiplying HCl by 2, not by altering the formula to HCl_2, which does not exist.

The steps in writing an equation are

1. write a word equation
2. put in the symbols and formulae
3. balance the equation

Example 4 The reaction between sodium and water to form hydrogen and sodium hydroxide solution. Work through the three steps:

1. Sodium + water → Hydrogen + Sodium hydroxide solution
2. $Na(s) + H_2O(l) \rightarrow H_2(g) + NaOH(aq)$
3. $2Na(s) + 2H_2O(l) \rightarrow H_2(g) + 2NaOH(aq)$

Example 5 Methane burns to form carbon dioxide and water:

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

There is one carbon atom on the left-hand side and one carbon atom on the right-hand side. There are four hydrogen atoms on the left-hand side, and therefore we need to put four hydrogen atoms on the right-hand side. Putting $2H_2O$ on the right-hand side will accomplish this:

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

There is one molecule of O_2 on the left-hand side and four O atoms on the right-hand side. We can make the two sides equal by putting $2O_2$ on the left-hand side:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

This is a balanced equation. The numbers of atoms of carbon, hydrogen and oxygen on the left-hand side are equal to the numbers of atoms of carbon, hydrogen and oxygen on the right-hand side.

Example 6 Propane also burns to form carbon dioxide and water:

$$C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

As there are three C atoms on the left-hand side, there must be $3CO_2$ molecules on the right-hand side:

$$C_3H_8(g) + O_2(g) \rightarrow 3CO_2(g) + H_2O(g)$$

As there are eight H atoms on the left-hand side, there must be $4H_2O$ on the right-hand side:

$$C_3H_8(g) + O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

Counting the oxygen atoms, there are two on the left-hand side and ten on the right-hand side. Putting $5O_2$ on the left-hand side will make the two sides equal:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

This is a balanced equation.

Practice with equations

1. For practice, try writing the equations for the reactions:
   (a) Hydrogen + Copper oxide → Copper + Water
   (b) Carbon + Carbon dioxide → Carbon monoxide
   (c) Carbon + Oxygen → Carbon dioxide
   (d) Magnesium + Sulphuric acid → Hydrogen + Magnesium sulphate
   (e) Copper + Chlorine → Copper(II) chloride

2. Now try writing balanced equations for the reactions:
   (a) Calcium + Water → Hydrogen + Calcium hydroxide solution
   (b) Copper + Oxygen → Copper(II) oxide
   (c) Sodium + Oxygen → Sodium oxide
   (d) Iron + Hydrochloric acid → Iron(II) chloride solution
   (e) Iron + Chlorine → Iron(III) chloride

3. Balance these equations:
   (a) $Na_2O(s) + H_2O(l) \rightarrow NaOH(aq)$
   (b) $KClO_3(s) \rightarrow KCl(s) + O_2(g)$
   (c) $H_2O_2(aq) \rightarrow H_2O(l) + O_2(g)$
   (d) $Fe(s) + O_2(g) \rightarrow Fe_3O_4(s)$
   (e) $Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$
   (f) $NH_3(g) + O_2(g) \rightarrow N_2(g) + H_2O(g)$
   (g) $Fe(s) + H_2O(g) \rightarrow Fe_3O_4(s) + H_2(g)$
   (h) $H_2S(g) + O_2(g) \rightarrow H_2O(g) + SO_2(g)$
   (i) $H_2S(g) + SO_2(g) \rightarrow H_2O(l) + S(s)$

2. Relative Atomic Mass; Relative Molecular Mass and Relative Formula Mass; Percentage Composition

Relative Atomic Mass

Atoms are tiny: one atom of hydrogen has a mass of 1.66×10^{-24}g; one atom of carbon has a mass of 1.99×10^{-23}g. Numbers as small as this are awkward to handle, and, instead of the actual masses, we use relative atomic masses. Since hydrogen atoms are the smallest of all atoms, one atom of hydrogen was taken as the mass with which all other atoms would be compared. Then,

$$\text{Relative atomic mass} = \frac{\text{Mass of one atom of the element}}{\text{Mass of one atom of hydrogen}}$$

Thus, on this scale, the relative atomic mass of hydrogen is 1, and, since one atom of carbon is 12 times as heavy as one atom of hydrogen, the relative atomic mass of carbon is 12.

The modern method of finding relative atomic masses is to use an instrument called a mass spectrometer. The most accurate measurements are made with volatile compounds of carbon, and it was therefore convenient to change the standard of reference to carbon. There are three isotopes of carbon. Isotopes are forms of an element which have the same number of protons and electrons but have different numbers of neutrons, and therefore different masses. It was decided to use the most plentiful carbon isotope, carbon-12. Thus,

$$\text{Relative atomic mass} = \frac{\text{Mass of one atom of an element}}{(1/12)\ \text{Mass of one atom of carbon-12}}$$

On this scale, carbon-12 has a relative atomic mass of 12, and hydrogen has a relative atomic mass 1.00797. Since relative atomic masses are ratios of two masses, they have no units. As this value for hydrogen is very close to one, the value of H = 1 is used in most calculations. A table of approximate relative atomic masses is given on page 145.

Relative Molecular Mass

A molecule consists of a combination of atoms. You can find the mass of a molecule by adding up the masses of all the atoms in it. You can find the relative molecular mass of a compound by adding the relative atomic masses of all the atoms in a molecule of the compound. For example, you can work out the relative molecular mass of carbon dioxide as follows:

The formula is CO_2

1 atom of C, relative atomic mass	12	= 12
2 atoms of O, relative atomic mass	16	= 32
Total		= 44
Relative molecular mass of CO_2		= 44

Relative Formula Mass

There are, however, a vast number of compounds which consist of ions, not molecules. The compound sodium chloride, for example, consists of sodium ions and chloride ions. One cannot correctly refer to a 'molecule of sodium chloride' or a 'molecule of copper sulphate'. For ionic compounds, the term *formula unit* is used to describe the ions which make up the compound. A formula unit of sodium chloride is NaCl. A formula unit of copper sulphate is $CuSO_4$. A formula unit of copper sulphate-5-water is $CuSO_4 \cdot 5H_2O$. The term relative formula mass is used for ionic compounds:

$$\text{Relative formula mass} = \frac{\text{Mass of one formula unit}}{(1/12)\ \text{Mass of one atom of carbon-12}}$$

Thus, for sodium chloride:

The formula is NaCl

1 atom of Na, relative atomic mass	23	= 23
1 atom of Cl, relative atomic mass	35.5	= 35.5
Total		= 58.5
Relative formula mass of NaCl		= 58.5

We work out the relative formula mass of calcium chloride as follows:

The formula is $CaCl_2$

1 atom of Ca, relative atomic mass	40	= 40
2 atoms of Cl, relative atomic mass	35.5	= 71

Total = 111

Relative formula mass of $CaCl_2$ = 111

We work out the relative formula mass of aluminium sulphate as follows:

The formula is $Al_2(SO_4)_3$

2 atoms of Al, relative atomic mass 27 = 54

3 atoms of S, relative atomic mass 32 = 96

12 atoms of O, relative atomic mass 16 = 192

Total = 342

Relative formula mass of $Al_2(SO_4)_3$ = 342

The term relative formula mass is convenient as it can be used *for both* covalent compounds and ionic compounds. The term relative molecular mass will be used *only for* covalent compounds which consist of molecules (e.g., methane, CH_4; ethanol, C_2H_5OH). You will find both terms used in the section of examination questions.

Problems on Relative Formula Mass

Work out the relative formula masses of these compounds:

SO_2	$NaOH$	KNO_3
$MgCO_3$	$PbCl_2$	$MgCl_2$
$Mg(NO_3)_2$	$Zn(OH)_2$	$ZnSO_4$
H_2SO_4	HNO_3	$MgSO_4 \cdot 7H_2O$
$CaSO_4$	Pb_3O_4	P_2O_5
Na_2CO_3	$Ca(OH)_2$	$CuCO_3$
$CuSO_4$	$Ca(HCO_3)_2$	$CuSO_4 \cdot 5H_2O$
$Al_2(SO_4)_3$	$Na_2CO_3 \cdot 10H_2O$	$FeSO_4 \cdot 7H_2O$

Percentage Composition

From the formula of a compound, one can work out the percentage by mass of each element present in the compound.

Example 1 Calculate the percentage of silicon and oxygen in silicon(IV) oxide (silica).