

GCE

Moles, Formulae and Equations

Edexcel Advanced GCE in Chemistry (9080)

Edexcel Advanced GCE in Chemistry (Nuffield) (9086)

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Workbook for GCE students

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Introduction

This workbook has been developed from an earlier version offering support to students in transition from GCSE Science (Double Award) and the GCE Advanced Subsidiary.

The aim of the booklet is to help students to practise their skills in the areas of formulae, equations and simple mole equations. The booklet gives examples for students to work through to help build their confidence. There are some sections involving multi-step calculations.

Edexcel acknowledges the help and support received from teachers in updating this latest edition. It replaces previous versions issued in January 1998 and August 2000.



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Section 1

Atoms

All matter is made of particles. At one time, it was thought that the tiniest particle was the *atom*; the word comes from the Greek word meaning 'indivisible'.

We now know that atoms can be split and that there are smaller particles than atoms, the so-called sub-atomic particles, electrons, protons and neutrons. You will need to know something about these particles which make up the different kinds of atoms.

However, you must understand that chemistry is all about rearrangements of atoms *that do not themselves* change.

Atoms are *very* small. The hydrogen atom, the smallest and lightest of all atoms, has a diameter of about 10^{-8} mm. 1 g of hydrogen atoms contains about 6×10^{23} atoms. It is very difficult to 'see' an individual atom and to find its mass.

An *atom* is the smallest, electrically neutral, particle of an element that can take part in a chemical change.

A *molecule* is the smallest, electrically neutral, particle of an element or compound that can exist on its own.

An *ion* is an atom, or group of atoms, which carries an electric charge.

You need to know these definitions by heart, but you also need to be able to recognise the formulae of atoms and molecules when you see them. Li, O, Cl, C are all formulae which represent atoms. Some of these can exist on their own, but not all of them. Oxygen, for example, always exists as oxygen *molecules*, O_2 , which contain two atoms, unless it is combined with something else. Water contains only one atom of oxygen but here it is combined with two hydrogen atoms.

Make sure that you really understand these ideas:

- a single oxygen atom, O, cannot exist on its own
- a single oxygen atom can exist when it is combined with something else, but then it is part of a molecule
- an oxygen molecule has two oxygen atoms, O_2
- a few elements exist as single atoms: for these elements, an atom is the same as a molecule.





Structure of the atom

The atom is composed of electrons, neutrons and protons. You have to remember the relative mass of, and the electric charge on, each.

Particle	Relative mass (Carbon -12 scale)	Relative charge (on scale electron charge = -1 unit)
Proton	1	+1
Electron	1/1840	-1
Neutron	1	0

The atom is mostly empty space. It has a solid core or *nucleus*, the centre that contains the protons and neutrons. The electrons circulate round the nucleus in specific *orbits* or *shells*.

We can picture the hydrogen atom - the simplest of all atoms with one electron, and one proton in the nucleus - by considering a pea placed in the centre of a football pitch, to represent the nucleus with its proton. On this scale the electron will revolve in a circular orbit round the goal posts. Between the electron and the nucleus is empty space.

Atoms are the particles whose symbols are found in the periodic table given in all your examination papers and also on page 113 of this book. You can see there are only about 100 of them. The middle part of the atom, the nucleus, contains one or more protons. It is the number of protons that make the atom what it is. An atom with one proton is always a hydrogen atom; one with two protons is a helium atom and so on.

There are more substances in the world than the 100 or so different kinds of atom. The other substances are made by combining atoms in various ways to make molecules.

When a chemical reaction takes place the atoms are rearranged to make different molecules but no atoms can be made or destroyed. To show this you have to be able to find a method of counting the atoms that take part in a reaction and its products.

The mass of an individual atom is very small and it is much more convenient to measure atomic masses as *relative masses*.

The definition of *relative atomic mass* A_r is:

The mass of a single atom on a scale on which the mass of an atom of carbon — 12 has a mass of 12 atomic mass units. The *relative atomic mass* does not have units.

The definition of *Relative Molecular Mass* M_r (also referred to as *Molar Mass*) is

The mass of a single molecule on a scale on which the mass of an atom of carbon — 12 has a mass of 12 atomic mass units.

Relative Molecular Mass of a molecule is calculated by adding together the relative atomic masses of the atoms in the chemical formulae.

Relative formula mass: in many ways this is more accurate than Relative Molecular Mass. Many salts, even in the solid state, exist as ions rather than molecules. Although the formula of sodium chloride is normally given as NaCl, it is not a simple molecule but a giant lattice and it is more accurately written as $(\text{Na}^+\text{Cl}^-)_n$. Since this compound does not have molecules, it cannot have relative 'molecular' mass. However, the principle is the same: add the relative atomic masses of sodium (23) and chlorine (35.5) to give 58.5, the relative formula mass of NaCl.

Remember: relative atomic mass, molecular mass and formula mass have no units.



Examples: Calculation of Molar Mass from Relative Atomic Mass data

Before you start any of these questions make sure you read the *Section 4* of this booklet (The mole on page 27).

When you carry out experiments you will weigh chemicals in grams. Molar mass has the same numerical value as the *Relative Molecular Mass*; it is calculated by adding together the relative atomic masses of the elements in the molecule. The total is expressed in units of grams per mol or g mol^{-1} .

Example 1

Calculate the Molar Mass of sulphuric acid H_2SO_4

This molecule contains

2 atoms of hydrogen each of mass 1	= 2 x 1	= 2 g mol^{-1}
1 atom of sulphur of mass 32	= 1 x 32	= 32 g mol^{-1}
4 atoms of oxygen of mass 16	= 4 x 16	= 64 g mol^{-1}
Total mass		= 98 g mol^{-1}

Example 2

Calculate the Molar Mass of lead nitrate $\text{Pb}(\text{NO}_3)_2$

Care! This molecule contains **TWO** nitrate groups

1 atom of lead of mass 207	= 1 x 207	= 207 g mol^{-1}
2 atoms of nitrogen of mass 14	= 2 x 14	= 28 g mol^{-1}
6 atoms of oxygen of mass 16	= 6 x 16	= 96 g mol^{-1}
Total mass		= 331 g mol^{-1}

Example 3

Calculate the Molar Mass of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

Care! This molecule has 5 molecules of water attached to each molecule of copper sulphate. Many students make the mistake of thinking that there are 10 hydrogens and only 1 oxygen.

In CuSO_4	1 atom of copper of mass 63.5	= 1 x 63.5	= 63.5 g mol^{-1}
	1 atom of sulphur of mass 32	= 1 x 32	= 32 g mol^{-1}
	4 atoms of oxygen of mass 16	= 4 x 16	= 64 g mol^{-1}
In $5\text{H}_2\text{O}$	5 x 2 atoms of hydrogen of mass 1	= 10 x 1	= 10 g mol^{-1}
	5 x 1 atoms of oxygen of mass 16	= 5 x 16	= 80 g mol^{-1}
	Total mass		= 249.5 g mol^{-1}

Calculations of this type are generally written as follows

$$\text{CuSO}_4 \cdot 5\text{H}_2\text{O} = [63.5 + 32 + (4 \times 16) + 5\{(2 \times 1) + 16\}] = 249.5 \text{ g mol}^{-1}$$



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Exercise 1

Calculation of the Molar Mass of compounds

Calculate the Molar Mass of the following. You will find data concerning Relative Atomic Masses on the periodic table (on page 113). When you have finished this set of calculations keep the answers for reference. You will find them useful in some of the other questions in this workbook.

1	H ₂ O
2	CO ₂
3	NH ₃
4	C ₂ H ₅ OH
5	C ₂ H ₄
6	SO ₂
7	SO ₃
8	HBr
9	H ₂ SO ₄
10	HNO ₃
11	NaCl
12	NaNO ₃
13	Na ₂ CO ₃
14	NaOH
15	Na ₂ SO ₄
16	KMnO ₄
17	K ₂ CrO ₄
18	KHCO ₃
19	KI
20	CsNO ₃
21	CaCl ₂

22 $\text{Ca}(\text{NO}_3)_2$

23 $\text{Ca}(\text{OH})_2$

24 CaSO_4

25 BaCl_2

26 AlCl_3

27 $\text{Al}(\text{NO}_3)_3$

28 $\text{Al}_2(\text{SO}_4)_3$

29 FeSO_4

30 FeCl_2

31 FeCl_3

32 $\text{Fe}_2(\text{SO}_4)_3$

33 PbO

34 PbO_2

35 Pb_3O_4

36 $\text{Pb}(\text{NO}_3)_2$

37 PbCl_2

38 PbSO_4

39 CuCl

40 CuCl_2

41 CuSO_4

42 ZnCl_2

43 AgNO_3

44 NH_4Cl

45 $(\text{NH}_4)_2\text{SO}_4$

46 NH_4VO_3

47 KClO_3

48 KIO_3

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49 NaClO

50 NaNO₂

51 CuSO₄·5H₂O

52 FeSO₄·7H₂O

53 (NH₄)₂SO₄·Fe₂(SO₄)₃·24H₂O

54 Na₂S₂O₃·5H₂O

55 (COOH)₂·2H₂O

56 MgSO₄·7H₂O

57 Cu(NH₃)₄SO₄·2H₂O

58 CH₃CO₂H

59 CH₃COCH₃

60 C₆H₅CO₂H



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Section 2

Chemical formulae

A chemical formula is a useful shorthand method for describing the atoms in a chemical: sometimes you will see the formula used instead of the name, but you should **not** do this if you are asked for a name.

The chemical formula of an element or compound tells you:

- Which elements it contains: eg FeSO_4 contains iron, sulphur and oxygen
- How many atoms of each kind are in each molecule: eg H_2SO_4 contains two atoms of hydrogen, one atom of sulphur and four atoms of oxygen in each molecule
- How the atoms are arranged: eg $\text{C}_2\text{H}_5\text{OH}$ contains a group of atoms known as the ethyl group, $-\text{C}_2\text{H}_5$, and a hydroxyl group, $-\text{OH}$
- The masses of the various elements in a compound: eg 18 g of water, H_2O , contains 2 g of hydrogen atoms and 16 g of oxygen since the relative atomic mass of hydrogen is 1 (x 2 because there two hydrogen atoms) and that of oxygen is 16.

You should not learn large numbers of chemical formulae by heart. However, it is useful to know a few of them and when you do you should be able to work out the rest. The table on page 10 shows the names, formulae and valency of the more common elements and some groups of atoms, called radicals, that you will study and you should refer to it when necessary.

Although it's best to learn formulae by using the valency of the common parts, it is sometimes useful to be able to work out the formula of a compound. This set of rules helps you to do this using information in the table.

You can think of valency as the combining power and use it to show the simplest ratio in which the atoms of the elements and radicals combine together in the formula. The following rules can now be applied:

- Write down the symbols of the elements and radicals given in the chemical name of the compound
- Now write down the valency of each element or radical under the corresponding symbols for the element or radical
- Now cross them over as shown in the example on page 10
- The valency shows the *simplest* combining ratio and may be cancelled down but only the valency can be simplified in this way
- If an element has more than one valency, the name of the compound will indicate which valency is to be used.

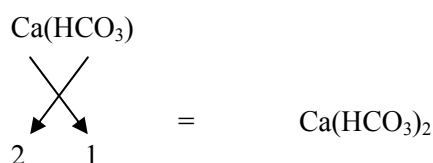


Here are a few examples:

- Sodium Sulphate



- Calcium hydrogen carbonate



Note: A bracket *must* be placed around the radical if it is multiplied by 2 or more *and* composed of more than one element.

Eg MgBr_2 no bracket required
 $\text{Ca}(\text{OH})_2$ bracket *essential* as CaOH_2 is incorrect.

- Often you can cancel the numbers on the two formulae:
 $\text{Ca}_2(\text{CO}_3)_2 = \text{CaCO}_3$
 However, you should **not** do this for organic compounds: C_2H_4 has two atoms of carbon and four of hydrogen so it cannot be cancelled down to CH_2 .
- Copper(I) oxide means use copper valency 1, ie Cu_2O : lead(II) nitrate means use lead valency 2, ie $\text{Pb}(\text{NO}_3)_2$

The periodic table can help you to find the valency of an element and hence the formula of its compounds.

Although you can use the table above to work out the formulae of many compounds it is important to realise that all formulae were originally found by experiment.

On page 11 you will find a table of the more common elements and groups that you may have met at GCSE. Also included are a few that you will meet in the first few weeks of your Advanced course or are mentioned in some of the calculations in this booklet. These are in italics.



Symbols and Valences of Common Elements and Radicals

ELEMENTS			RADICALS		
	Symbol	Valency		Symbol	Valency
Aluminium	Al	3	Ammonium	NH ₄	1
Barium	Ba	2	Carbonate	CO ₃	2
Bromine	Br	1	Chloride	Cl	1
Calcium	Ca	2	Hydrogen-carbonate	HCO ₃	1
Carbon	C	4	Hydrogen-sulphate	HSO ₃	1
Chlorine	Cl	1	Hydroxide	OH	1
Cobalt	Co	2	Nitrate	NO ₃	1
Copper	Cu	1&2	Nitrite	NO ₂	1
Hydrogen	H	1	Sulphate	SO ₄	2
Iodine	I	1	Sulphite	SO ₃	2
Iron	Fe	2&3			
Lead	Pb	2&4	<i>Chlorate(I)</i>	<i>ClO</i>	<i>1</i>
Magnesium	Mg	2	<i>Chlorate(V)</i>	<i>ClO₃</i>	<i>1</i>
Manganese	Mn	2&4	<i>Vanadate(V)</i>	<i>VO₃</i>	<i>1</i>
Mercury	Hg	1&2	<i>Manganate(VII)</i>	<i>MnO₄</i>	<i>1</i>
Nitrogen	N	3&5	<i>Chromate(VI)</i>	<i>CrO₄</i>	<i>2</i>
Oxygen	O	2	<i>Dichromate(VI)</i>	<i>Cr₂O₇</i>	<i>2</i>
Phosphorus	P	3&5			
Potassium	K	1			
Silicon	Si	4			
Silver	Ag	1			
Sodium	Na	1			
Sulphur	S	2,4,6			



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Exercise 2

Writing formulae from names

Use the data in the table on page 11 to write the formulae of the following. Before you start this exercise, make sure you have read *Section 3 (Naming of compounds)* on page 19) of this booklet.

1 Sodium Chloride

2 Sodium Hydroxide

3 Sodium Carbonate

4 Sodium Sulphate

5 Sodium Phosphate

6 Potassium Chloride

7 Potassium Bromide

8 Potassium Iodide

9 Potassium Hydrogen Carbonate

10 Potassium Nitrite

11 Magnesium Chloride

12 Magnesium Nitrate

13 Magnesium Hydroxide

14 Magnesium Oxide

15 Magnesium Carbonate

16 Calcium Oxide

17 Calcium Chloride

18 Calcium Sulphate

19 Calcium Carbonate

20 Barium Chloride

21 Barium Sulphate

22 Aluminium Chloride

23 Aluminium Oxide

24 Aluminium Hydroxide

25 Aluminium Sulphate

26 Copper(II) Sulphate

27 Copper(II) Oxide

28 Copper(II) Chloride

29 Copper(II) Nitrate

30 Copper(I) Oxide

31 Copper(I) Chloride

32 Zinc Nitrate

33 Zinc Carbonate



34 Zinc Oxide

35 Silver Chloride

36 Silver Bromide

37 Silver Iodide

38 Silver Nitrate

39 Silver Oxide

40 Lead(II) Nitrate

41 Lead(II) Carbonate

42 Lead(II) Oxide

43 Lead(IV) Oxide

44 Lead(II) Chloride

45 Lead(IV) Chloride

46 Lead(II) Sulphide

47 Tin(II) Chloride

48 Tin(IV) Chloride

49 Iron(II) Sulphate

50 Iron(II) Chloride

51 Iron(III) Sulphate

52 Iron(III) Chloride

53 Iron(III) Hydroxide

54 Iron(II) Hydroxide

55 Ammonium Chloride

56 Ammonium Carbonate

57 Ammonium Hydroxide

58 Ammonium Nitrate

59 Ammonium Sulphate

60 Ammonium Phosphate

61 Phosphorus Trichloride

62 Phosphorus Pentachloride

63 Phosphorus Trioxide

64 Phosphorus Pentoxide

65 Hydrogen Phosphate (Phosphoric Acid)

66 Hydrogen Sulphate (Sulphuric Acid)

67 Hydrogen Nitrate (Nitric Acid)

68 Hydrogen Chloride (Hydrochloric Acid)

69 Carbon Tetrachloride

70 Silicon Tetrachloride



71 Silicon Dioxide

72 Sulphur Dioxide

73 Sulphur Trioxide

74 Hydrogen Sulphide

75 Chlorine(I) Oxide

76 Nitrogen Dioxide

77 Nitrogen Monoxide

78 Carbon Dioxide

79 Carbon Monoxide

80 Hydrogen Hydroxide



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Section 3

Naming of compounds

At Advanced Level you will meet many compounds that are new to you; a lot of these will be organic compounds. In this section, we will be looking at the naming of compounds that you may already have met at GCSE level. Many of these compounds are named using simple rules. However, there are some that have 'trivial' names not fixed by the rules. It is important that you learn the names and formulae of these compounds. Later in the course, you will learn the rules for naming most of the organic compounds you will meet.

Naming inorganic compounds

The name must show which elements are present and, where confusion is possible, the valency of the elements concerned.

- 1 You need to remember that if there are only two elements present then the name will end in **-ide**

Thus, **oxides** contain an element and oxygen

eg Na_2O is **Sodium Oxide**
 CaO is **Calcium Oxide**

Chlorides contain an element and chlorine

eg MgCl_2 is **Magnesium Chloride**
 AlCl_3 is **Aluminium Chloride**

Bromides and **Iodides** have an element and either bromine or iodine

eg KBr is **Potassium Bromide**
 ZnI is **Zinc Iodide**

Hydrides contain an element and hydrogen and **Nitrides** an element and nitrogen.

eg LiH is **Lithium Hydride**
 Mg_3N_2 is **Magnesium Nitride**

Other elements also form these types of compounds and the name always ends in **-ide**. The exceptions to this are **hydroxides** that have the **-OH** group and **cyanides**, which have the **-CN** group.

eg NaOH is **Sodium Hydroxide**
 Ca(OH)_2 is **Calcium Hydroxide**
 KCN is **Potassium Cyanide**



- 5 Because most non-metals can have more than one valency they can also produce more than one acid upon which these groups are based. Thus sulphur can form **sulphates** and **sulphites**. The ending **-ite** is used when an element forms more than one such compound. In all cases the **-ite** is used for the compound with the lower number of oxygens atoms. **Sulphate** can also be referred to as **sulphate(VI)** and **sulphite** can also be referred to as **sulphate(IV)**. In the case of nitrogen with oxygen the compounds would be **nitrate** and **nitrite** or **nitrate(V)** and **nitrate(III)**.

In summary:

<i>Common name</i>	<i>Systematic name</i>	<i>Formulae</i>
Sulphate	Sulphate(VI)	-SO₄
Sulphite	Sulphate(IV)	-SO₃
Nitrate	Nitrate(V)	-NO₃
Nitrite	Nitrate(III)	-NO₂
Chlorate	Chlorate(V)	-ClO₃
Hypochlorite	Chlorate(I)	-ClO

Great care needs to be taken when using these systematic names, as they are called, because the properties of the two groups of compounds will be very different. In some cases the use of the wrong compound in a reaction could cause considerable danger. For this reason you should always read the label on a bottle or jar and make sure it corresponds exactly to what you should be using.

Other elements can form compounds involving oxygen in this way. These include **Chlorate(V)**, **Chromate(VI)**, **Manganate(VII)** and **Phosphate(V)**.

eg KNO₂ is **Potassium Nitrite** or **Potassium Nitrate(III)**
 Na₂SO₃ is **Sodium Sulphite** or **Sodium Sulphate(IV)**
 K₂CrO₄ is **Potassium Chromate(VI)**
 KMnO₄ is **Potassium Manganate(VII)**
 KClO₃ is **Potassium Chlorate(V)**

- 6 When a compound is considered it is usual to put the metal down first both in the name and the formula. The exceptions to this rule are in organic compounds where the name has the metal first but the formula has the metal at the end.

eg CH₃COONa is **Sodium Ethanoate**



- 7 The elements nitrogen and **hydrogen** can join together to form a group called the **ammonium** group. This must not be confused with the compound **ammonia**, but more of that later. This **ammonium** group has the formula NH_4^+ and sits in the place generally taken by a metal in the formula.

eg NH_4Cl is **Ammonium Chloride**
 $(\text{NH}_4)_2\text{SO}_4$ is **Ammonium Sulphate**
 NH_4ClO_3 is **Ammonium Chlorate(V)**

- 8 There are a small number of simple molecules that do not follow the above rules. You will need to learn their names and formulae.

They include:

Water which is H_2O
Sulphuric Acid which is H_2SO_4
Nitric Acid which is HNO_3
Hydrochloric Acid which is HCl
Ammonia which is NH_3
Methane which is CH_4

- 8 Organic compounds have their own set of rules for naming but you will need to learn some of the basic rules. The names are generally based on the names of the simple hydrocarbons. These follow a simple pattern after the first four:

CH_4 is **Methane**
 C_2H_6 is **Ethane**
 C_3H_8 is **Propane**
 C_4H_{10} is **Butane**

After butane the names are based on the prefix for the number of carbons C_5 -**pent**, C_6 - **hex** and so on.

Thus organic compounds with 2 carbons will either start with **Eth-** or have **-eth-** in their name.

eg C_2H_4 is **Ethene**
 $\text{C}_2\text{H}_5\text{OH}$ is **Ethanol**
 CH_3COOH is **Ethanoic Acid**
 $\text{C}_2\text{H}_5\text{Cl}$ is **Chloroethane**

Exercise 3

Names from formulae

Use the notes in this section, the data in the table on page 11 and the copy of the periodic table on page 113 to write the names of the following. Before you start this exercise make sure you have read *Section 2* of this booklet (*Chemical formulae* on page 9).

1 H_2O

2 CO_2

3 NH_3

4 O_2

5 H_2

6 SO_2

7 SO_3

8 HCl

9 HI

10 HF

11 CH_4

12 H_2S

13 HBr

14 H_2SO_4

15 HNO_3

16 NaCl

17 NaNO_3

18 Na_2CO_3

19 NaOH

20 Na_2SO_4

21 CaCl_2

22 $\text{Ca}(\text{NO}_3)_2$

23 $\text{Ca}(\text{OH})_2$

24 CaSO_4

25 BaCl_2

26 AlCl_3

27 $\text{Al}(\text{NO}_3)_3$

28 $\text{Al}_2(\text{SO}_4)_3$

29 FeSO_4

30 FeCl_2

31 FeCl_3

32 $\text{Fe}_2(\text{SO}_4)_3$

33 PbO

34 PbO_2

35 $\text{Pb}(\text{NO}_3)_2$

36 PbCl_2

37 PbSO_4

38 $\text{Cu}(\text{NO}_3)_2$

39 CuCl

40 CuCl_2

41 CuSO_4

42 ZnCl_2

43 AgNO_3

44 NH_4Cl

45 $(\text{NH}_4)_2\text{SO}_4$

46 NH_4VO_3 (V is Vanadium)

47 KClO_3

48 KIO_3

49 NaClO

50 NaNO_2



51 C_2H_6

52 C_4H_{10}

53 C_8H_{18}

54 $(NH_4)_2CO_3$

55 $KMnO_4$

56 K_2CrO_4

57 $KHCO_3$

58 KI

59 $Co(NO_3)_2$

60 KAt



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Section 4

The mole

When chemists measure how much of a particular chemical reacts they measure the amount in grams; or they measure the volume of a gas. However, chemists find it convenient to use a unit called a **mole**. You need to know several definitions of a mole and be able to use them.

- The **mole** is the amount of substance, which contains the same number of particles (atoms, ions, molecules, formulae or electrons) as there are carbon atoms in 12 g of carbon -12
- This **number** is known as the *Avogadro constant*, L , and is equal to $6.02 \times 10^{23} \text{ mol}^{-1}$
- The **molar mass** of a substance is the mass, in grams, of one mole
- The **molar volume** of a gas is the volume occupied by one mole at room temperature and atmospheric pressure (r.t.p). It is equal to 24 dm^3 at r.t.p
- *Avogadro's Law* states that equal volumes of all gases, under the same conditions of temperature and pressure contain the same number of moles or molecules. If the volume is 24 dm^3 , at room temperature and pressure, this number, once again, is the Avogadro constant.

When you talk about moles, you must always state whether you are dealing with atoms, molecules, ions, formulae etc. To avoid any ambiguity it is best to show this as a formula.

Example calculations involving the use of moles

These calculations form the basis of many of the calculations you will meet at A level.

Example 1**Calculation of the number of moles of material in a given mass of that material**

- a Calculate the number of moles of oxygen atoms in 64 g of oxygen atoms. *You need the mass of one mole of oxygen atoms. This is the Relative Atomic Mass in grams; in this case it is 16 g mol^{-1} .*

$$\text{number of moles of atoms} = \frac{\text{mass in grams}}{\text{molar mass of atoms}}$$

$$\begin{aligned} \therefore \text{number of moles of oxygen} &= \frac{64 \text{ g of oxygen atoms}}{\text{molar mass of oxygen of } 16 \text{ g mol}^{-1}} \\ &= \mathbf{4 \text{ moles of oxygen atoms}} \end{aligned}$$

- b Calculate the number of moles of chlorine molecules in 142 g of chlorine gas.

$$\text{number of moles of atoms} = \frac{\text{mass in grams}}{\text{molar mass of atoms}}$$

The first stage of this calculation is to calculate the molar mass of Chlorine molecules.
Molar mass of $\text{Cl}_2 = 2 \times 35.5 = 71 \text{ g mol}^{-1}$

$$\begin{aligned} \therefore \text{number of moles of chlorine} &= \frac{142 \text{ g of chlorine gas}}{\text{molar mass of chlorine of } 71 \text{ g mol}^{-1}} \\ &= \mathbf{2 \text{ moles of chlorine molecules}} \end{aligned}$$



- c Calculate the number of moles of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in 100 g of the solid.

The Relative Molecular Mass of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ =

$$[63.5 + 32 + (4 \times 16) + 5\{(2 \times 1) + 16\}] = 249.5 \text{ g mol}^{-1}$$

$$\therefore \text{ number of moles of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O} = \frac{100 \text{ g of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O}}{\text{molecular mass of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O of } 249.5 \text{ g mol}^{-1}}$$

$$= \mathbf{0.4008 \text{ moles of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O} \text{ molecules}}$$

Example 2

Calculation of the mass of material in a given number of moles of that material

The mass of a given number of moles	=	the mass of 1 mole	x	the number of moles of material concerned
--	---	-------------------------------	---	--

- a Calculate the mass of 3 moles of sulphur dioxide SO_2

$$1 \text{ mole of sulphur dioxide has a mass} = 32 + (2 \times 16) = 64 \text{ g mol}^{-1}$$

$$\therefore 3 \text{ moles of } \text{SO}_2 = 3 \times 64 = \mathbf{192 \text{ g}}$$

- b What is the mass of 0.05 moles of $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$?

$$1 \text{ mole of } \text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O} = [(23 \times 2) + (32 \times 2) + (16 \times 3)] + 5[(2 \times 1) + 16]$$

$$= 248 \text{ g mol}^{-1}$$

$$\therefore 0.05 \text{ moles of } \text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O} = 0.05 \times 248 = \mathbf{12.4 \text{ g}}$$

Example 3**Calculation of the volume of a given number of moles of a gas**

All you need to remember is that **1 mole of any gas has a volume of 24 dm³ (24000 cm³) at room temperature and pressure.**

$\therefore \frac{\text{The volume of a given number of moles of gas}}{\text{of moles of gas}} = \text{number of moles} \times 24000 \text{ cm}^3$
--

- a What is the volume of 2 mol of carbon dioxide?

Remember you do not need to work out the molar mass to do this calculation as it does not matter what gas it is.

$$\therefore 2 \text{ moles of carbon dioxide} = 2 \times 24000 \text{ cm}^3 = 48000 \text{ cm}^3 = \mathbf{48 \text{ dm}^3}$$

- b What is the volume of 0.0056 moles of chlorine molecules?

$$\text{Volume of 0.0056 moles of chlorine} = 0.0056 \times 24000 \text{ cm}^3 = \mathbf{134.4 \text{ cm}^3}$$

Example 4**Calculation of the number of moles of gas in a given volume of that gas**

$\text{number of moles of gas} = \frac{\text{volume of gas in cm}^3}{24000 \text{ cm}^3}$

- a Calculate the number of moles of hydrogen molecules in 240 cm³ of the gas.

$$\text{number of moles} = \frac{240 \text{ cm}^3}{24000 \text{ cm}^3} = 0.010 \text{ moles}$$

- b How many moles of a gas are there in 1000 cm³ of the gas?

$$\text{number of moles of gas} = \frac{1000 \text{ cm}^3}{24000 \text{ cm}^3} = 0.0147 \text{ moles}$$

**Example 5****Calculation of the volume of a given mass of gas**

This calculation require you to apply the skills covered in the previous examples

Calculate the volume of 10 g of hydrogen gas.

This is a two-stage calculation a) you need to calculate how many moles of hydrogen gas are present, and b) you need to convert this to a volume.

$$\begin{aligned} \therefore \text{number of moles of hydrogen (H}_2\text{)} &= \frac{10 \text{ g of hydrogen (H}_2\text{)}}{\text{molecular mass of hydrogen (H}_2\text{) of } 2 \text{ g mol}^{-1}} \\ &= 5 \text{ moles} \end{aligned}$$

$$\therefore 5 \text{ moles of hydrogen} = 5 \times 24000 \text{ cm}^3 = 120000 \text{ cm}^3 = \mathbf{120 \text{ dm}^3}$$

Example 6**Calculation of the mass of a given volume of gas**

This calculation require you to apply the skills covered in the previous examples

Calculate the mass of 1000 cm³ of carbon dioxide

Again this is a two-stage calculation a) you need to calculate the number of moles of carbon dioxide and then b) convert this to a mass.

$$\begin{aligned} \therefore \text{number of moles of CO}_2 &= \frac{1000 \text{ cm}^3 \text{ of CO}_2}{\text{volume of 1 mole of CO}_2 \text{ of } 24000 \text{ cm}^3} \\ &= 0.0147 \text{ moles} \end{aligned}$$

$$\therefore 0.0147 \text{ moles of carbon dioxide} = 0.0147 \times 44 \text{ g} = \mathbf{1.833 \text{ g}}$$



Example 7

Calculation of the molar mass of a gas from mass and volume data for the gas

Calculations of this type require you to find the mass of 1 mole of the gas, ie 24 000 cm³. This is the molar mass of the gas.

eg Calculate the Relative Molecular Mass of a gas for which 100 cm³ of the gas at room temperature and pressure, have a mass of 0.0667 g

100 cm³ of the gas has a mass of 0.0667 g

$$\begin{aligned}\therefore 24000 \text{ cm}^3 \text{ of the gas must have a mass of} &= \frac{0.0667 \text{ g} \times 24000 \text{ cm}^3}{100 \text{ cm}^3} \\ &= 16 \text{ g}\end{aligned}$$

\therefore The molar mass of the gas is 16 g mol⁻¹

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Exercise 4a

Calculation of the number of moles of material in a given mass of that material

In this set of calculations all the examples chosen are from the list of compounds whose molar mass you calculated in exercise 1.

In each case calculate the number of moles of the material in the mass stated.

1 9.00 g of H₂O

2 88.0 g of CO₂

3 1.70 g of NH₃

4 230 g of C₂H₅OH

5 560 g of C₂H₄

6 0.640 g of SO₂

7 80.0 g of SO₃

8 18.0 g of HBr

9 0.0960 g of H₂SO₄

10 3.15 g of HNO₃

11 19.3 g of NaCl

12 21.25 g of NaNO₃

13 2.25 g of Na₂CO₃

14 0.800 g of NaOH

15 17.75 g of Na₂SO₄

16 3.16 g of KMnO_4

17 32.33 g of K_2CrO_4

18 100 g of KHCO_3

19 7.63 g of KI

20 3.90 g of CsNO_3

21 0.111 g of CaCl_2

22 41.0 g of $\text{Ca}(\text{NO}_3)_2$

23 1.48 g of $\text{Ca}(\text{OH})_2$

24 3.40 g of CaSO_4

25 41.6 g of BaCl_2

26 14.95 g of CuSO_4

27 13.64 g of ZnCl_2

28 1.435 g of AgNO_3

29 13.76 g of NH_4Cl

30 13.76 g of $(\text{NH}_4)_2\text{SO}_4$

31 23.4 g of NH_4VO_3

32 10.0 g of KClO_3

33 10.7 g of KIO_3

34 100 g of NaClO

35 1.70 g of NaNO_2

36 50.9 g of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

37 19.6 g of $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$

38 9.64 g of $(\text{NH}_4)_2\text{SO}_4 \cdot \text{Fe}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$

39 12.4 g of $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$

40 32.0 g of $(\text{COOH})_2 \cdot 2\text{H}_2\text{O}$

41 3.075 g of $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$

42 40.0 g of $\text{Cu}(\text{NH}_3)_4\text{SO}_4 \cdot 2\text{H}_2\text{O}$

43 6.00 g of $\text{CH}_3\text{CO}_2\text{H}$

44 3.10 g of CH_3COCH_3

45 0.530 g of $\text{C}_6\text{H}_5\text{CO}_2\text{H}$

46 4.79 g of AlCl_3

47 56.75 g of $\text{Al}(\text{NO}_3)_3$

48 8.35 g of $\text{Al}_2(\text{SO}_4)_3$

49 3.8 g of FeSO_4

50 200 g of FeCl_2



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Exercise 4b

Calculation of the mass of material in a given number of moles of the material

In each case calculate the mass in grams of the material in the number of moles stated.

1 2 moles of H₂O

2 3 moles of CO₂

3 2.8 moles of NH₃

4 0.50 moles of C₂H₅OH

5 1.2 moles of C₂H₄

6 0.64 moles of SO₂

7 3 moles of SO₃

8 1 mole of HBr

9 0.012 moles of H₂SO₄

10 0.15 moles of HNO₃

11 0.45 moles of NaCl

12 0.70 moles of NaNO₃

13 0.11 moles of Na₂CO₃

14 2.0 moles of NaOH

15 0.90 moles of Na₂SO₄

16 0.050 moles of KMnO₄

17 0.18 moles of K_2CrO_4

18 0.90 moles of $KHCO_3$

19 1.5 moles moles of KI

20 0.12 moles of $CsNO_3$

21 0.11 moles of $CaCl_2$

22 4.1 moles of $Ca(NO_3)_2$

23 0.0040 moles of $Ca(OH)_2$

24 0.10 moles of $CaSO_4$

25 0.21 moles of $BaCl_2$

26 0.10 moles of $CuSO_4$

27 0.56 moles of $ZnCl_2$

28 0.059 moles of $AgNO_3$

29 0.333 moles of NH_4Cl

30 1.1 moles of $(NH_4)_2SO_4$

31 0.025 moles of NH_4VO_3

32 0.10 moles of $KClO_3$

33 0.10 moles of KIO_3

34 10 moles of $NaClO$

35 0.0010 moles of $NaNO_2$



36 0.20 moles of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

37 0.10 moles of $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$

38 0.0050 moles of $(\text{NH}_4)_2\text{SO}_4 \cdot \text{Fe}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$

39 0.040 moles of $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$

40 2.4 moles of $(\text{COOH})_2 \cdot 2\text{H}_2\text{O}$

41 3.075 moles of $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$

42 0.15 moles of $\text{Cu}(\text{NH}_3)_4\text{SO}_4 \cdot 2\text{H}_2\text{O}$

43 0.17 moles of $\text{CH}_3\text{CO}_2\text{H}$

44 0.20 moles of CH_3COCH_3

45 0.080 moles of $\text{C}_6\text{H}_5\text{CO}_2\text{H}$

46 0.0333 moles of AlCl_3

47 0.045 moles of $\text{Al}(\text{NO}_3)_3$

48 0.12 moles of $\text{Al}_2(\text{SO}_4)_3$

49 2.0 moles of FeSO_4

50 11 moles of FeCl_2



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Exercise 4c

Calculation of the volume of a given number of moles of a gas

In each case calculate the volume of the number of moles of gas stated.

(Assume that all volumes are measured at room temperature and pressure and that 1 mole of gas has a volume of $24\,000\text{ cm}^3$ under these conditions).

1 1 mole of CO_2

2 0.1 moles of NH_3

3 0.5 moles of C_2H_4

4 2 moles of SO_2

5 0.12 moles of SO_3

6 3.4 moles of HBr

7 0.11 moles of Cl_2

8 0.0040 moles of CH_4

9 10 moles of H_2

10 0.45 moles of O_2

11 0.0056 moles of C_2H_6

12 0.0090 moles of C_3H_8

13 0.040 moles of C_2H_2

14 0.123 moles of NO

15 0.0023 moles of HCl

16 8.0 moles of HBr

17 0.000010 moles of HI

18 6.0 moles of NO₂

19 0.0076 moles of F₂

20 3.0 moles of N₂

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Exercise 4d

Calculation of the number of moles of gas in a given volume of that gas

In each case calculate the volume of the number of moles of gas stated.

(Assume that all volumes are measured at room temperature and pressure and that 1 mol of gas has a volume of $24\,000\text{ cm}^3$ under these conditions).

1 200 cm^3 of CO_2

2 500 cm^3 of NH_3

3 1000 cm^3 of C_2H_4

4 2000 cm^3 of SO_2

5 234 cm^3 of SO_3

6 226 cm^3 of HBr

7 256 cm^3 of Cl_2

8 200 cm^3 of CH_4

9 2000 cm^3 of H_2

10 2400 cm^3 of O_2

11 700 cm^3 of C_2H_6

12 5600 cm^3 of C_3H_8

13 2200 cm^3 of C_2H_2

14 210 cm^3 of NO

15 800 cm^3 of HCl



16 80 cm³ of HBr

17 2 cm³ of HI

18 20 000 cm³ of NO₂

19 420 cm³ of F₂

20 900 cm³ of N₂

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Exercise 4e

Calculation of the mass of a given volume of gas

Calculate the mass of the volume of gases stated below.

(Assume that all volumes are measured at room temperature and pressure and that

1 mol of gas has a volume of 24 000 cm³ under these conditions).

1 200 cm³ of CO₂

2 500 cm³ of NH₃

3 1000 cm³ of C₂H₄

4 2000 cm³ of SO₂

5 234 cm³ of SO₃

6 226 cm³ of HBr

7 256 cm³ of Cl₂

8 200 cm³ of CH₄

9 2000 cm³ of H₂

10 2400 cm³ of O₂

11 700 cm³ of C₂H₆

12 5600 cm³ of C₃H₈

13 2200 cm³ of C₂H₂

14 210 cm³ of NO

15 800 cm³ of HCl



16 80 cm³ of HBr

17 2 cm³ of HI

18 20 000 cm³ of NO₂

19 420 cm³ of F₂

20 900 cm³ of N₂

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Exercise 4f

Calculation of the volume of a given mass of gas

In each case calculate the volume in cm^3 of the mass of gas given.

(Assume that all volumes are measured at room temperature and pressure and that

1 mol of gas has a volume of $24\,000\text{ cm}^3$ under these conditions).

1 2 g of CO_2

2 5 g of NH_3

3 10 g of C_2H_4

4 20 g of SO_2

5 2.34 g of SO_3

6 2.26 g of HBr

7 10 g of Cl_2

8 20 g of CH_4

9 200 g of H_2

10 240 g of O_2

11 70 g of C_2H_6

12 56 g of C_3H_8

13 22 g of C_2H_2

14 20 g of NO

15 8 g of HCl



16 8 g of HBr

17 2 g of HI

18 23 g of NO₂

19 42 g of F₂

20 90 g of N₂

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Exercise 4g

Calculation of the Relative Molecular Mass of a gas from mass and volume data for the gas

In each case you are given the mass of a certain volume of an unknown gas. From each set of data calculate the Relative Molecular Mass of the gas.

(Assume that all volumes are measured at room temperature and pressure and that

1 mol of gas has a volume of 24 000 cm³ under these conditions).

1 0.373 g of gas occupy 56 cm³

2 0.747 g of gas occupy 280 cm³

3 0.467 g of gas occupy 140 cm³

4 0.296 g of gas occupy 100 cm³

5 0.0833 g of gas occupy 1000 cm³

6 0.175 g of gas occupy 150 cm³

7 0.375 g of gas occupy 300 cm³

8 0.218 g of gas occupy 90 cm³

9 0.267 g of gas occupy 200 cm³

10 1.63 g of gas occupy 1400 cm³

11 0.397 g of gas occupy 280 cm³

12 0.198 g of gas occupy 280 cm³

13 0.0602 g of gas occupy 38 cm³

14 0.0513 g of gas occupy 44 cm³



15 0.0513 g of gas occupy 28 cm³

16 1.33 g of gas occupy 1000 cm³

17 8.79 g of gas occupy 1000 cm³

18 0.0760 g of gas occupy 50 cm³

19 0.338 g of gas occupy 100 cm³

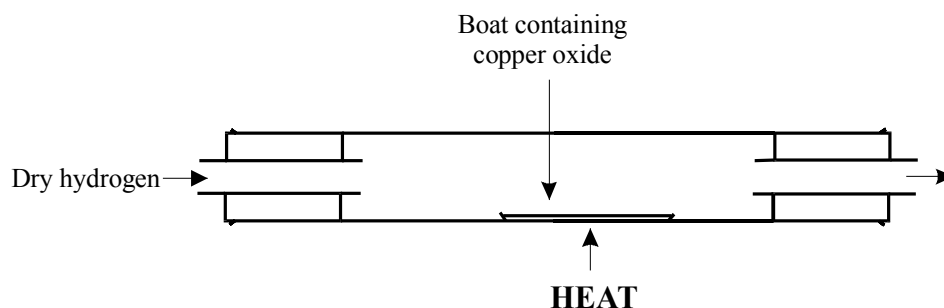
20 0.667 g of gas occupy 125 cm³

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Section 5

Using the idea of moles to find formulae

You can find the formula of copper(II) oxide by passing a stream of hydrogen over a known mass of copper oxide and weighing the copper formed.



- A known mass of copper(II) oxide is used.
- A stream of hydrogen from a cylinder is passed over the copper until all the air has been swept out of the apparatus.
- It is heated to constant mass (until two consecutive mass determinations at the end of the experiment are same) in a stream of *dry* hydrogen.
- The mass of the copper is finally determined.

Note:

- Excess hydrogen must **not** be ignited until it has been tested (by collection in a test tube) to make sure that all the air has been expelled from the apparatus. If the hydrogen in the test tube burns quietly, without a squeaky pop, then it is safe to ignite it at the end of the tube.
- The combustion tube is tilted to prevent the condensed steam from running back on to the hot part of the tube.
- When the reduction process is complete, ie after heating to constant mass, the tube is allowed to cool with hydrogen still being passed over the remaining copper. This is to prevent the copper from being oxidized to copper(II) oxide.

The working on the next page shows you how to calculate the results:

Typical results

Mass of copper (II) oxide	= 5 g
Mass of copper	= 4 g
Mass of oxygen	= 1 g



	÷ by relative atomic mass (r.a.m)	÷ by smallest	Ratio of atoms
Moles of copper atoms	$\frac{4}{64} = 0.0625$	$\frac{0.0625}{0.0625} = 1$	1
Moles of oxygen atoms	$\frac{1}{16} = 0.0625$	$\frac{0.0625}{0.0625} = 1$	1

Therefore, the simplest (or empirical) formula is CuO.

The apparatus may be modified to determine the formula of water. Anhydrous calcium chloride tubes are connected to the end of the combustion tube and the excess hydrogen then ignited at the end of these tubes. Anhydrous calcium chloride absorbs water; the mass of the tubes is determined at the beginning and end of the experiment. The increase in mass of the calcium chloride tubes is equal to the mass of water produced.

Typical results

Mass of water = 1.125 g
 Mass of oxygen (from previous results) = 1.000 g
 Mass of hydrogen = 0.125 g

	÷ by r.a.m	÷ by smallest	Ratio of atoms
Moles of hydrogen atoms	$\frac{0.125}{1} = 0.125$	$\frac{0.125}{0.0625} = 2$	2
Moles of oxygen atoms	$\frac{1}{16} = 0.0625$	$\frac{0.0625}{0.0625} = 1$	1

Since the ratio of hydrogen to oxygen is 2:1 the simplest (or empirical) formula is H₂O.

In examination calculations of this type the data is often presented not as mass, but as percentage composition of the elements concerned. In these cases the calculation is carried out in an identical fashion as percentage composition is really the mass of the element in 100 g of the compound.

**Example 1**

Sodium burns in excess oxygen to give a yellow solid oxide that contains 58.97% of sodium. What is the empirical formula of the oxide?

N.B. This is an oxide of sodium. It must contain only Na and O. Since the percentage of Na is 58.97 that of O must be $100 - 58.97 = 41.03\%$

	÷ by r.a.m	÷ by smallest	Ratio of atoms
Moles of sodium atoms	$\frac{58.97}{23} = 2.564$	$\frac{2.564}{2.564} = 1$	1
Moles of oxygen atoms	$\frac{41.03}{16} = 2.564$	$\frac{2.564}{2.564} = 1$	1

Therefore the empirical formula is **NaO**.

The result of the above calculation does not seem to lead to a recognisable compound of sodium. This is because the method used only gives the **simplest** ratio of the elements - but see below.

Consider the following series of organic compounds:

C₂H₄ ethene, C₃H₆ propene, C₄H₈ butene, C₅H₁₀ pentene. These all have the same empirical formula CH₂.

To find the Molecular Formula for a compound it is necessary to know the Relative Molecular Mass (M_r).

Molecular Formula Mass = Empirical Formula Mass × a whole number (n)

In the example above the oxide has an M_r = 78 g mol⁻¹.

Thus

$$\begin{aligned} \text{Molecular Formula Mass} &= 78 \\ \text{Empirical Formula Mass} &= (\text{Na} + \text{O}) = 23 + 16 = 39 \\ \therefore 78 &= 39 \times n \\ \therefore n &= 2 \end{aligned}$$

The Molecular Formula becomes **(NaO)₂ or Na₂O₂**

**Example 2**

A compound **P** contains 73.47% carbon and 10.20% hydrogen by mass, the remainder being oxygen. It is found from other sources that **P** has a Relative Molecular Mass of 98 g mol^{-1} . Calculate the molecular formula of **P**.

NB It is not necessary to put in all the details when you carry out a calculation of this type. The following is adequate.

	C	H	O
	73.47	10.20	$(100 - 73.47 - 10.20)$ = 16.33
By r.a.m	$\frac{73.47}{12}$	$\frac{10.20}{1}$	$\frac{16.33}{16}$
	= 6.1225	= 10.20	= 1.020
By smallest	$\frac{6.1255}{1.020}$	$\frac{10.20}{1.020}$	$\frac{1.020}{1.020}$
Ratio of atoms	6	10	1

Therefore the empirical formula is **C₆H₁₀O**.

To find molecular formula:

$$\text{Molecular Formula Mass} = \text{Empirical Formula Mass} \times \text{whole number (n)}$$

$$98 = [(6 \times 12) + (10 \times 1) + 16] \times n = 98 \times n$$

$$\therefore n = 1$$

The molecular formula is the same as the empirical formula **C₆H₁₀O**.

**A warning**

In calculations of this type at GCE Advanced level you may meet compounds that are different yet have very similar percentage composition of their elements. When you carry out a calculation of this type you should never round up the figures until you get right to the end. For example NH_4OH and NH_2OH have very similar composition and if you round up the data from one you may well get the other. If you are told the Relative Molecular Mass and your Empirical Formula Mass is not a simple multiple of this you are advised to check your calculation.

Example 3

Calculate the empirical formula of a compound with the following percentage composition:

C 39.13%; O 52.17%; H 8.700%

	C	O	H
By r.a.m	$\frac{39.13}{12}$	$\frac{52.17}{16}$	$\frac{8.700}{1}$
	= 3.26	= 3.25	= 8.70
Divide by smallest	1	1	2.66

It is clear at this stage that dividing by the smallest has not resulted in a simple ratio. **You must not round up or down at this stage.** You must look at the numbers and see if there is some factor that you could multiply each number by to get each one to a whole number. In this case if you multiply each by 3 you will get:

C	O	H
3	3	8

Thus $\text{C}_3\text{H}_8\text{O}_3$ is the empirical formulae not $\text{C}_1\text{H}_{2.66}\text{O}_1$

You need to watch carefully for this, the factors will generally be clear and will be 2 or 3. What you must not do is round 1.33 to 1 or 1.5 to 2.



Calculations involving the moles of water of crystallization

In calculations of this type you need to treat the water as a *molecule* and divide by the *Relative Molecular Mass*.

Example 4

24.6 grams of a hydrated salt of $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$, gives 12.0 g of anhydrous MgSO_4 on heating. What is the value of x ?

Your first job is to find the mass of water driven off.

$$\text{Mass of water evolved} = 24.6 - 12.0 = 12.6 \text{ g}$$

	MgSO_4	H_2O
	12.0	12.6
Divide by M_r	$\frac{12.0}{120}$	$\frac{12.6}{18}$
	= 0.100	= 0.700
Ratio of Atoms	1	7

Giving a formula of $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$

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Exercise 5

Calculation of a formula from experimental data

In *Section a*. calculate the empirical formula of the compound from the data given. This may be as percentage composition or as the masses of materials found in an experiment. For *Section b*. you are given the data for analysis plus the Relative Molecular Mass of the compound, in these cases you are to find the empirical formula and thence the molecular formula. *Section c*. is more difficult, the data is presented in a different fashion but the calculation of the empirical formula/molecular formula is essentially the same.

Section a

1 Ca 40%; C 12%; O 48%

2 Na 32.4%; S 22.5%; O 45.1%

3 Na 29.1%; S 40.5%; O 30.4%

4 Pb 92.8%; O 7.20%

5 Pb 90.66%; O 9.34%

6 H 3.66%; P 37.8%; O 58.5%

7 H 2.44%; S 39.0%; O 58.5%

8 C 75%; H 25%



9 C 81.81%; H 18.18%

10 H 5.88% ; O 94.12%

11 H 5%; N 35%; O 60%

12 Fe 20.14%; S 11.51%; O 63.31%; H 5.04%

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Section b

-
- 13 A hydrocarbon with a Relative Molecular Mass (M_r) of 28 g mol^{-1} has the following composition: Carbon 85.7%; Hydrogen 14.3%. Calculate its molecular formula.
-
- 14 A hydrocarbon with a Relative Molecular Mass (M_r) of 42 g mol^{-1} has the following composition: Carbon 85.7%; Hydrogen 14.3%. Calculate its molecular formula.
-
- 15 P 10.88%; I 89.12%. $M_r = 570 \text{ g mol}^{-1}$
-
- 16 N 12.28%; H 3.51%; S 28.07%; O 56.14%. $M_r = 228 \text{ g mol}^{-1}$
-
- 17 P 43.66%; O 56.34%. $M_r = 284 \text{ g mol}^{-1}$
-
- 18 C 40%; H 6.67%; O 53.3%. $M_r = 60 \text{ g mol}^{-1}$
-
- 19 Analysis of a compound with a $M_r = 58 \text{ g mol}^{-1}$ shows that 4.8 g of carbon are joined with 1.0 g of hydrogen. What is the molecular formula of the compound?
-
- 20 3.36 g of iron join with 1.44 g of oxygen in an oxide of iron that has a $M_r = 160 \text{ g mol}^{-1}$. What is the molecular formula of the oxide?
-
- 21 A sample of an acid with a $M_r = 194 \text{ g mol}^{-1}$ has 0.5 g of hydrogen joined to 16 g of sulphur and 32 g of oxygen. What is the molecular formula of the acid?
-
- 22 Analysis of a hydrocarbon showed that 7.8 g of the hydrocarbon contained 0.6 g of hydrogen and that the $M_r = 78 \text{ g mol}^{-1}$. What is the formula of the hydrocarbon?
-



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Section c

23 22.3 g of an oxide of lead produced 20.7 g of metallic lead on reduction with hydrogen. Calculate the empirical formula of the oxide concerned.

24 When 1.17 g of potassium is heated in oxygen 2.13 g of an oxide is produced. In the case of this oxide the empirical and molecular formulae are the same. What is the molecular formula of the oxide produced?

25 A hydrocarbon containing 92.3% of carbon has a Relative Molecular Mass of 26 g mol^{-1} . What is the molecular formula of the hydrocarbon?

26 When 1.335 g of a chloride of aluminium is added to excess silver nitrate solution 4.305 g of silver chloride is produced. Calculate the empirical formula of the chloride of aluminium.

Hint; you will need to work out how much chlorine there is in 4.305 g of AgCl. This will be the amount of chlorine in the initial 1.335 g of the aluminium chloride.

27 16 g of a hydrocarbon burn in excess oxygen to produce 44 g of carbon dioxide. What is the empirical formula of the hydrocarbon.

Hint; you will need to work out what mass of carbon is contained in 44 g of CO_2 . This is the mass of C in 16 g of the hydrocarbon.

28 When an oxide of carbon containing 57.1% oxygen is burnt in air the percentage of oxygen joined to the carbon increases to 72.7%. Show that this data is consistent with the combustion of carbon monoxide to carbon dioxide.

29 When 14.97 g of hydrated copper(II) sulphate is heated it produces 9.60 g of anhydrous copper(II) sulphate. What is the formula of the hydrated salt?

30 When the chloride of phosphorus containing 85.1% chlorine is heated a second chloride containing 77.5% chlorine is produced. Find the formulae of the chlorides and suggest what the other product of the heating might be.



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Section 6

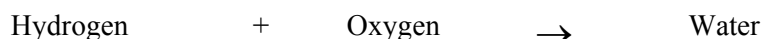
Chemical equations

Chemical equations do much more than tell you what reacts with what in a chemical reaction. They tell you how many of each type of molecule is needed and are produced and so they also tell you what masses of the reactants are needed to produce a given mass of products.

Often you will learn equations that have been given to you. However, if you are to interpret equations correctly you must learn to write them for yourself.

Equations in words

Before you can begin to write an equation, you must know what the reacting chemicals are and what is produced in the reaction. You can then write them down as a *word equation*. For instance, hydrogen reacts with oxygen to give water, or, as a word equation:

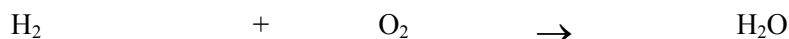


Writing formulae

When you have written the equation in words you can write the formula for each of the substances involved; you may know them or you may have to look them up. In our example:

Hydrogen is represented as H_2 ;
Oxygen is represented as O_2 ; and
Water is H_2O .

So we get:

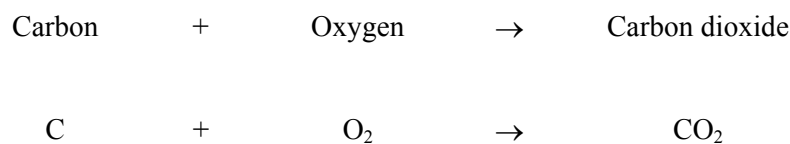


However this will not do as a full equation as you will discover if you read on!

Balancing the equation

One of the most important things you must understand in chemistry is that atoms are **rearranged** in chemical reactions: they are never produced from 'nowhere' and they do not simply 'disappear'. This means that in chemical equation you must have the same number of atoms of each kind on the left-hand side of the equation as on the right. Sometimes you need to start with two or more molecules of one of the reactants and you may end up with more than one molecule of one of the products.

Let us look at two very simple examples:

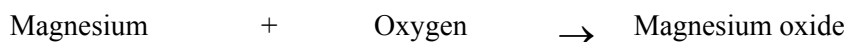




It so happens that carbon dioxide has one atom of carbon and two atoms of oxygen in one molecule: carbon is written as C (one atom) and oxygen molecules have two atoms each: written as O₂.

This equation does not need balancing because the number of atoms of carbon is the same on the left as on the right (1) and the number of atoms oxygen is also the same (2) ie it is already balanced.

Now let us try one that does not work out:



Magnesium is written as Mg (one atom just like carbon) and oxygen is, of course, O₂, but magnesium oxide has just one atom of oxygen per molecule and is therefore written as MgO.

So we might write:



The magnesium balances, one atom on the left and one on the right, but the oxygen does not as there are two atoms on the left-hand side of the equation and only one on the right hand side.

You cannot change the formulae of the reactants or products.

Each 'formula' of magnesium oxide has only one atom of oxygen: each molecule of oxygen has two atoms of oxygen, so you can make *two* formulae of magnesium oxide for each molecule of oxygen. So we get:



Even now the equation does not balance, because we need two atoms of magnesium to make two formulae of MgO, and the final equation is:



Sometimes, you will need to show in the equation whether the chemicals are solid, liquid or gas. You do this by putting in *state symbols*: (aq) for aqueous solution, (g) for gas, (l) for liquid and (s) for solid or precipitate:

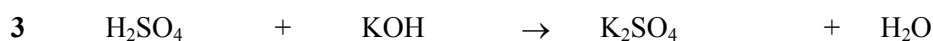
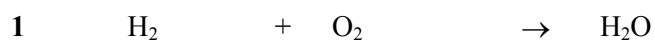


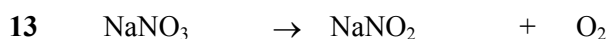
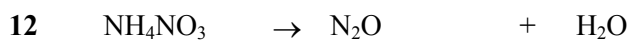
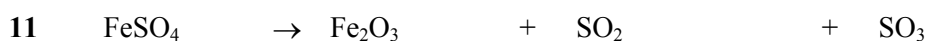


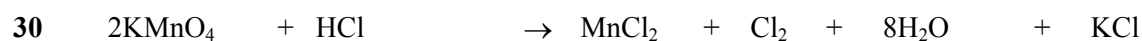
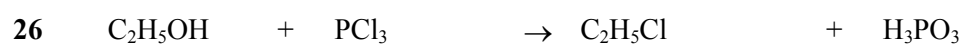
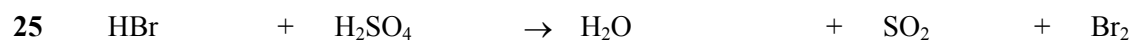
Exercise 6a

Balancing equations

Balance the following equations. To get you started $_$ indicates the first six questions where numbers need to be inserted to achieve the balance. In one or two difficult examples some of the numbers have been added. You will not need to change these. Also remember all the formulae are correct!









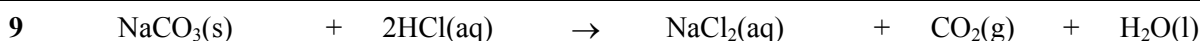
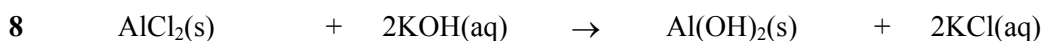
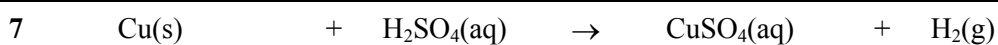
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Exercise 6b

What's wrong here?

The following equations have one or more mistakes. These may be in a formula, in the balancing, in the state symbols or even in the chemistry. Your job is to identify the error and then write a correct equation.





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Exercise 6c

Writing equations in symbols from equations in words

In the following examples you will need to convert the names of the materials into formulae and then balance the resulting equation. In some cases more than one experiment is described. In these cases you will need to write more than one equation.

1 Zinc metal reacts with copper sulphate solution to produce solid copper metal and zinc sulphate solution.

2 Solid calcium hydroxide reacts with solid ammonium chloride on heating to produce solid calcium chloride, steam and ammonia gas.

3 When lead(II) nitrate is heated in a dry tube lead(II) oxide, nitrogen dioxide gas and oxygen are produced.

4 Silicon tetrachloride reacts with water to produce solid silicon dioxide and hydrogen chloride gas.

5 When a solution of calcium hydrogen carbonate is heated a precipitate of calcium carbonate is produced together with carbon dioxide gas and more water.

6 When octane (C_8H_{18}) vapour is burnt with excess air in a car engine carbon dioxide and water vapour are produced.

7 All the halogens, apart from fluorine, react with concentrated sodium hydroxide solution to produce a solution of the sodium halide (NaX) the sodium halate ($NaXO_3$) and water.

8 The elements of Group 1 of the periodic table all react with water to produce a solution of the hydroxide of the metal and hydrogen gas.

The last two examples in this section will need a lot of thought as they involve changes in the valency of the elements concerned. Before you start to balance the equations check with your teacher that you have the formulae correct.

8 Tin(II) chloride solution reacts with mercury(II) chloride solution to produce a precipitate of mercury(I) chloride and a solution of tin(IV) chloride. This precipitate of mercury(I) chloride then reacts with further tin(II) chloride solution to produce liquid mercury and more tin(IV) chloride solution.

9 Concentrated sulphuric acid reacts with solid potassium iodide to produce solid potassium hydrogen sulphate, iodine vapour, water and hydrogen sulphide gas.

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Section 7

How equations are found by experiment

Although equations are often printed in books for you to learn, you must remember that they have all been found originally by someone doing experiments to measure how much of each chemical reacted and how much of each product was formed.

Below are set out some of the methods you could use:

- *Direct mass determinations*, eg the reaction of magnesium with oxygen

A known mass of magnesium is heated in a crucible to constant mass and hence the mass of magnesium oxide is found. Supposing 0.12 g of magnesium produce 0.20 g of magnesium oxide. By subtraction, the mass of oxygen combined with the magnesium is 0.080 g.

Each of these masses is then converted to moles and it is found that every 2 moles of magnesium react with one mole oxygen molecules and produce two moles of magnesium oxide: hence



- *Reacting volumes in solution*: usually you have to calculate concentrations of acids or alkalis by reaction with the appropriate standard solution and use the chemical equation for the reaction

However, you can calculate the ratio of reacting moles from experimental data, in order to construct the equation. To do this you use solutions, both of whose concentrations you know. You then do a titration in the usual way and use the volumes used in the titration to find the number of moles of each reagent which react.

These are then used in the equation straight away, just as in the magnesium oxide example above.

- *Measurement of gas volumes*: the molar volume of a gas is taken as 24 dm³ at room temperature and atmospheric pressure (r.t.p.)

Examples

- 1 In an experiment a solution containing 3.31 g of lead(II) nitrate reacts with a solution containing 1.17 g of sodium chloride to produce 2.78 g of lead(II) chloride solid and leave a solution that contains 1.70 g of sodium nitrate. What is the equation for the reaction?

In this type of question you are required to calculate the ratio of the reacting moles and then use these to write the equation.

$\text{Pb}(\text{NO}_3)_2$	NaCl	PbCl_2	NaNO_3
$M_r = 331$	$M_r = 58.5$	$M_r = 278$	$M_r = 85$

$$\therefore 3.31 \text{ g of Pb}(\text{NO}_3)_2 = (3.31/331) \text{ mol} = 0.010 \text{ mol}$$

$$1.17 \text{ g of NaCl} = (1.17/58.5) \text{ mol} = 0.020 \text{ mol}$$

$$2.78 \text{ g of PbCl}_2 = (2.78/278) \text{ mol} = 0.010 \text{ mol}$$

$$1.70 \text{ g of NaNO}_3 = (1.70/85) \text{ mol} = 0.020 \text{ mol}$$

$$\therefore 0.01 \text{ mol of Pb}(\text{NO}_3)_2 \text{ reacts with } 0.02 \text{ mol of NaCl} \text{ to give } 0.01 \text{ mol of PbCl}_2 \text{ and } 0.02 \text{ mol of NaNO}_3$$

$$\text{ie } 1 \text{ mol of Pb}(\text{NO}_3)_2 \text{ reacts with } 2 \text{ mol of NaCl} \text{ to give } 1 \text{ mol of PbCl}_2 \text{ and } 2 \text{ mol of NaNO}_3$$



It is not necessary to write all of this out each time.

- 2 When 5.175 g of lead are heated at 300°C the lead reacts with the oxygen in the air to produce 5.708 g of an oxide of lead. This is the only product. What is the equation for this reaction? *In a question of this type you seem to be short of information but in fact you know the mass of oxygen reacting. Remember this is oxygen molecules that are reacting not oxygen atoms.*

$$\text{Mass of oxygen used is } 5.708 - 5.175 \text{ g} = 0.533 \text{ g}$$

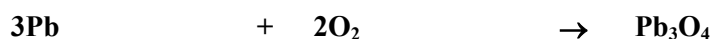
$$\text{Moles of lead reacting} = (5.175/207) \text{ mol} = 0.025 \text{ mol}$$

$$\text{Moles of oxygen reacting} = (0.533/32) \text{ mol} = 0.0167 \text{ mol}$$

$$\therefore 0.025 \text{ mol of Pb react with } 0.0167 \text{ mol of O}_2 \text{ to give product}$$

$$\therefore 1.5 \text{ mol of Pb react with } 1 \text{ mol of O}_2 \text{ to give product}$$

$$\therefore 3 \text{ mol of Pb react with } 2 \text{ mol of O}_2 \text{ to give product}$$



You do not have the information to write the full equation but as you know there is only one product and this has 3 lead atoms and 4 oxygen you can suggest a formula.



- 3 25 cm³ of 2M sulphuric acid solution react with 50 cm³ of 2M sodium hydroxide solution to produce sodium sulphate and water. Construct the equation for this reaction.

You will need to look at the start of chapter 11 before you can follow this question.

25 cm³ of 2M H₂SO₄ contain (25 x 2/1000) mol of H₂SO₄ = 0.050 mol

50 cm³ of 2M NaOH contain (50 x 2/1000) mol of NaOH = 0.10 mol

∴ 0.05 mol of H₂SO₄ react with 0.10 mol of NaOH to give Na₂SO₄ plus H₂O

∴ 1 mol of H₂SO₄ react with 2 mol of NaOH to give Na₂SO₄ plus H₂O



- 4 2 cm³ of nitrogen gas react completely with 6 cm³ of hydrogen gas to produce 4cm³ of ammonia gas. Use the data to write the equation for this reaction.

2 cm³ of nitrogen = (2/24000) mol = 8.33 x 10⁻⁵ mol

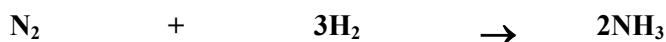
6 cm³ of hydrogen = (6/24000) mol = 2.50 x 10⁻⁴ mol

4 cm³ of ammonia = (4/24000) mol = 1.67 x 10⁻⁴ mol

∴ ratios are (8.33 x 10⁻⁵/8.33 x 10⁻⁵) of nitrogen = 1

(2.50 x 10⁻⁴/8.33 x 10⁻⁵) of hydrogen = 3

(1.67 x 10⁻⁴/8.33 x 10⁻⁵) of ammonia = 2



- 5 1 g of CaCO₃ reacts with 10 cm³ of 2M HNO₃ to produce 1.64 g of Ca(NO₃)₂, 240 cm³ of CO₂ and water.

In practice the acid will be in water and it is almost impossible to measure the amount of water produced by the reaction.

1/ 100 mol of CaCO₃ + (10x2)/1000 mol of HNO₃ → 1.64/164 mol of Ca(NO₃)₂ + 240/24000 mol of CO₂ + H₂O

∴ 0.01 mol of CaCO₃ + 0.02 mol of HNO₃ → 0.01 mol of Ca(NO₃)₂ + 0.01 mol of CO₂ + H₂O





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Exercise 7

Writing chemical equations from experimental data

Use the data below to write the equations for the reactions listed. In some cases you may not be able to calculate the moles of all the materials involved. In these cases you should indicate that you have 'balanced' this part yourself.

In examples involving gases you should assume 1 Mole of gas occupies 24 000 cm³ at room temperature and pressure.

-
- 1 In an experiment a solution containing 6.675 g of aluminium chloride reacted with a solution containing 25.50 g of silver nitrate. 21.52 g of silver chloride was produced together with a solution of 10.65 g of aluminium nitrate, Al(NO₃)₃. What is the equation for the reaction taking place?
-
- 2 100 cm³ of a solution of potassium chromate(VI), containing 97.05 g dm⁻³, reacts with 50 cm³ of a solution, containing 331 g dm⁻³ of lead nitrate, to produce 16.15 g of a precipitate of lead(II) chromate and 150 cm³ of a solution of potassium nitrate, which gives 10.1 g of solid when the water is evaporated off from the solution. Write the equation for the reaction.
-
- 3 1.133 g of silver nitrate was heated in an open tube. The silver residue weighed 0.720 g. During the reaction 0.307 g of nitrogen dioxide was also produced. The rest of the mass loss was due to oxygen. Use the data to write the equation for the reaction.
-
- 4 In a titration using methyl orange as an indicator 25.0 cm³ of a solution of 0.1 M sodium hydroxide reacted with 25.0 cm³ of 0.1 M phosphoric acid, H₃PO₄, solution. If the experiment is repeated using phenolphthalein in place of the methyl orange as the indicator the volume of the sodium hydroxide used to cause the indicator to change colour is 50.0 cm³.
-
- i Use the data to calculate the number of moles of sodium hydroxide that reacts with one mole of phosphoric acid in each case
-
- ii Suggest the formula of the salt produced in each case
-
- iii Write the equations
-
- iv What volume of the alkali would be needed to produce the salt Na₃PO₄?
-

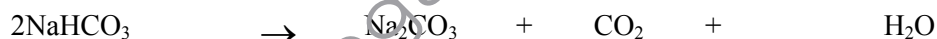
5 50 cm³ of a solution of citric acid, $M_r = 192$, containing 19.2 g dm⁻³ reacted with 50 cm³ of a solution of sodium hydroxide containing 12 g dm⁻³. Citric acid can be represented by the formula H_xA, where x represents the number of hydrogen atoms in the molecule. Use the data above to calculate the number of moles of sodium hydroxide that react with one mole of citric acid and hence find the value of x.

6 When 12.475 g of hydrated copper(II) sulphate, CuSO₄.xH₂O, was heated 7.980 g of anhydrous salt were produced. Use the data to find the value of x and hence write the equation for the reaction.

7 When 20 cm³ ammonia gas is passed over a catalyst with excess oxygen 20 cm³ of nitrogen monoxide (NO) and 30 cm³ of water vapour are produced. Use this data to write the equation for the reaction.

8 10 cm³ of a hydrocarbon C_aH_b reacted with 50 cm³ of oxygen gas to produce 30 cm³ of carbon dioxide and 40 cm³ of water vapour. Use the data to calculate the reacting moles in the equation and suggest value for a and b.

9 When 8.4 g of sodium hydrogen carbonate are heated 5.30 g of solid residue are produced 1200 cm³ of carbon dioxide are produced and 0.900 g of water are evolved. Show that this data is consistent with the following equation.



10 When 13.9 g of FeSO₄.xH₂O is heated 4 g of solid iron (III) oxide is produced together with the loss of 1.6 g of sulphur dioxide and 2.0 g of sulphur trioxide. The rest of the mass loss being due to the water of crystallization being lost. Use the data to write the full equation for the action of heat.



Section 8

Amounts of substances

Equations can also tell you how much of a chemical is reacting or is produced. The equation in *Section 7* tells us that 2 moles of (solid) magnesium atoms react with 1 mole of (gaseous) oxygen molecules to produce 2 moles of (solid) magnesium oxide molecules.

We know that the relative atomic mass of magnesium is 24, and that of oxygen is 16, (see periodic table on page 113). And from the equation we balanced in *Section 6* we can suggest that 48 g of magnesium react with 32 g of oxygen (because an oxygen molecule contains two atoms) to give 80 g of magnesium oxide.

Since we know the ratio of reacting masses (or volumes in the case of gases) we can calculate any reacting quantities based on the equation.

Example 1

a What mass of magnesium oxide would be produced from 16 g of magnesium in the reaction between magnesium and oxygen?

i Write the full balanced equation



ii Read the equation in terms of moles

2 moles of magnesium reacts to give 2 moles of magnesium oxide

iii Convert the moles to masses using the M_r values

$$\therefore (2 \times 24\text{g}) \text{ of magnesium gives } 2 \times (24+16) = 80 \text{ g of Magnesium oxide}$$

$$\therefore 16 \text{ g of magnesium gives } \frac{80 \times 16}{2 \times 24} = \mathbf{26.7 \text{ g of Magnesium oxide}}$$

- b What volume of oxygen would react with 16 g of magnesium in the above reaction?

In this case the oxygen is a gas so the volume of each mole is 24000 cm³ at room temperature and pressure and you do not have to worry about the molecular mass of the gas.

From the equation:

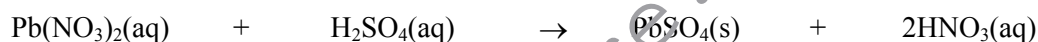
2 moles of Mg reacts with 1 mole of O₂

∴ 2 x 24 g of Mg reacts with 1 x 24000 cm³ of O₂(g)

$$\therefore 16 \text{ g of Mg reacts with } \frac{1 \times 24000 \text{ cm}^3 \times 16 \text{ g}}{2 \times 24 \text{ g}} = \mathbf{8000 \text{ cm}^3 \text{ of oxygen}}$$

Example 2

What mass of lead(II) sulphate would be produced by the action of excess dilute sulphuric acid on 10 g of lead nitrate dissolved in water ?



∴ 1 mole of lead nitrate gives 1 mole of lead sulphate

∴ 331 g of lead nitrate gives 303 g of lead sulphate

$$\therefore 10 \text{ g of lead nitrate gives } \frac{303 \text{ g} \times 10 \text{ g of lead sulphate}}{331 \text{ g}} = \mathbf{9.15 \text{ g of lead sulphate}}$$

Example 3

What is the total volume of gas produced by the action of heat on 1 g of silver nitrate?



2 moles of silver nitrate give 2 moles of nitrogen dioxide gas plus 1 mole of oxygen gas = 3 moles of gas

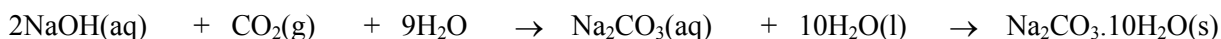
∴ 2 x 170 g of silver nitrate give 3 x 24000 cm³ of gas

$$\therefore 1 \text{ g of silver nitrate gives } \frac{3 \times 24000 \text{ cm}^3 \times 1 \text{ g of gas}}{2 \times 170 \text{ g}} = \mathbf{211.8 \text{ cm}^3 \text{ of gas}}$$

**Example 4**

When excess carbon dioxide is passed into sodium hydroxide solution, sodium carbonate solution is formed. This can be crystallised out as $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$. What mass of crystals would be produced from 5 g of sodium hydroxide in excess water.

Care. You need the water as moles in the equation.



\therefore 2 moles of sodium hydroxide give 1 mole of the crystals of sodium carbonate

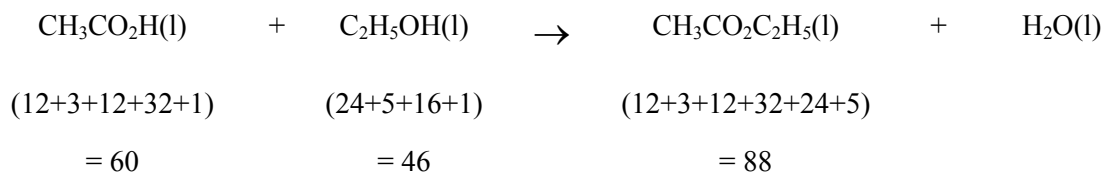
\therefore 2 x 40 g of sodium hydroxide give 286 g of crystals

$$\therefore 5 \text{ g of sodium hydroxide give } \frac{286 \times 5}{2 \times 40} = \mathbf{17.88 \text{ g of crystals}}$$

Example 5

What mass of ethanoic acid and what mass of ethanol would be needed to produce 100 g of ethyl ethanoate assuming the reaction went to completion?

Care! In this question you know how much you want to get and are asked how much you will need to start with. In these cases you must read the equation from the other end ie 1 mole of the ethyl ethanoate is produced from 1 mole of acid and 1 mole of alcohol.



\therefore 88 g of ethyl ethanoate are produced from 60 g of ethanoic acid and 46 g of ethanol

$$\therefore 100 \text{ g of ethyl ethanoate are produced from } \frac{60 \text{ g} \times 100 \text{ g}}{88 \text{ g}} = \mathbf{68.2 \text{ g of ethanoic acid}}$$

$$\text{and } \frac{46 \text{ g} \times 100 \text{ g}}{88 \text{ g}} = \mathbf{52.3 \text{ g of ethanol}}$$



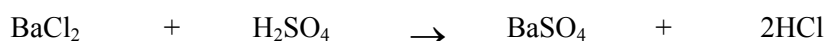
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Exercise 8

Calculations of products/reactants based on equations

In this exercise the equations you need are given in the question, unless they were included in Exercise 6a.

-
- 1 What mass of barium sulphate would be produced from 10 g of barium chloride in the following reaction?



-
- 2 What mass of potassium chloride would be produced from 20 g of potassium carbonate?

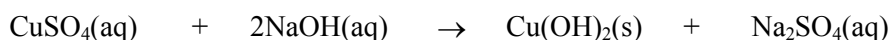
-
- 3 What masses of ethanol and ethanoic acid would need to be reacted together to give 1 g of ethyl ethanoate?

-
- 4 What mass of iron(III) oxide would need to be reduced to produce 100 tonnes of iron in a blast furnace?

-
- 5 What mass of silver nitrate as a solution in water would need to be added to 5 g of sodium chloride to ensure complete precipitation of the chloride?

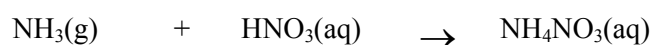


-
- 6 A solution of copper sulphate reacts with sodium hydroxide solution to produce a precipitate of copper hydroxide according to the following equation:



What mass of sodium hydroxide would be needed to convert 15.95 g of copper sulphate to copper hydroxide and what mass of copper hydroxide would be produced?

-
- 7 What volume of ammonia gas would be needed to produce 40 g of ammonium nitrate in the following reaction?



-
- 8 In the reaction between calcium carbonate and nitric acid what mass of calcium nitrate and what volume of carbon dioxide would be produced from 33.3 g of calcium carbonate?
-

9 What would be the total volume of gas produced by the action of heat on 33.1 g of lead(II) nitrate ?

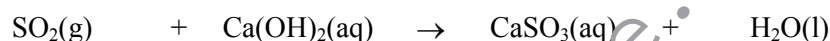
10 Magnesium reacts with sulphuric acid to produce a solution of magnesium sulphate. If this is allowed to crystallise out the solid produced has the formula $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$.

Write the equation for this reaction and calculate the mass of magnesium sulphate heptahydrate that could be produced from 4 g of magnesium.

11 Copper(II) oxide reacts with sulphuric acid to produce copper(II) sulphate. If this is allowed to crystallise the formula of the crystals is $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

What mass of copper oxide would be needed to produce 100 g of crystals?

12 Sulphur dioxide can be removed from the waste gases of a power station by passing it through a slurry of calcium hydroxide. The equation for this reaction is:



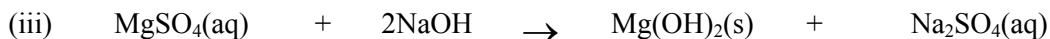
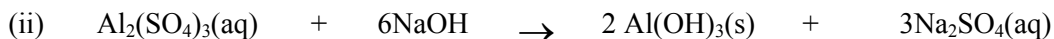
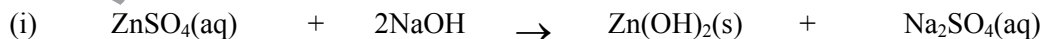
What mass of calcium hydroxide would be needed to deal with 1000 dm^3 of sulphur dioxide?

13 In a fermentation reaction glucose is converted to alcohol and carbon dioxide according to the following equation:



What mass of alcohol and what volume of carbon dioxide would be produced from 10 g of glucose?

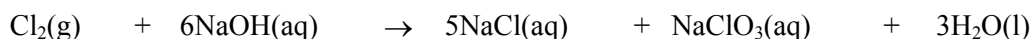
14 In the following reactions calculate the mass of precipitate formed from 20 g of the metal salt in each case.



15 What volume of hydrogen would be produced by 1 g of calcium in its reaction with water?

16 What mass of magnesium would be needed to produce 100 cm^3 of hydrogen?

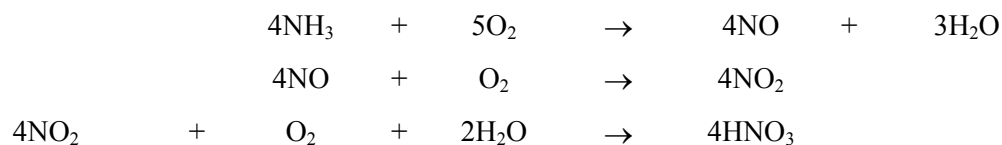
17 Chlorine reacts with sodium hydroxide as follows:



What mass of sodium chloride and what mass of sodium chlorate(V) be produced from 240 cm³ of chlorine gas?

18 When nitrogen reacts with hydrogen in the Haber Process only 17% of the nitrogen is converted to ammonia. What volume of nitrogen and what volume of hydrogen would be needed to produce 1 tonne of ammonia? (1 tonne = 1 x 10⁶ g)

19 Nitric acid is produced by the following series of reactions:



What mass of nitric acid would be produced from 17 tonnes of ammonia and what volume of oxygen would be needed in the reaction?

20 Hardness in water is caused by dissolved calcium compounds. When heated some of these break down and deposits calcium carbonate as follows:



This builds up as 'fur' on the inside of boilers. It can be removed by reaction with hydrochloric acid.

What mass of calcium carbonate would be produced from 10000 dm³ of water containing 0.356 g of calcium hydrogen carbonate per dm³ of water and what volume of 10 mol dm⁻³ hydrochloric acid solution would be needed to remove the solid calcium carbonate from the inside of the boiler?



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Section 9

Reactions involving gases

Whenever gases are involved in a reaction you need to remember that they have both mass and volume and that **1 mole of any gas has the same volume, 24 000 cm³, at room temperature and 1 atmosphere pressure as 1 mole of any other gas.** (See *Section 4* for more details)

This means:

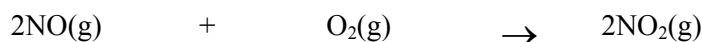
2 g of hydrogen, H₂, has a volume of 24 000 cm³

32 g of oxygen, O₂, has a volume of 24 000 cm³

81 g of hydrogen bromide, HBr, has a volume of 24 000 cm³

The effect of this is to make calculations involving gas volumes much easier than you might expect.

Consider the following reaction:



This says:

2 moles of NO(g) react with 1 mole of O₂(g) to give 2 moles of NO₂(g)

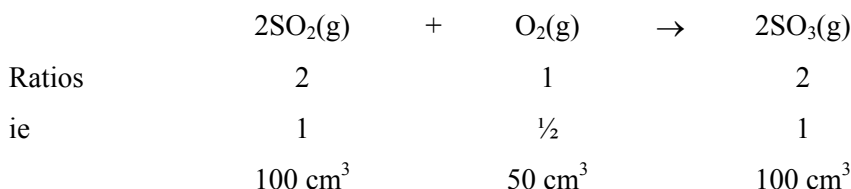
∴ (2 x 24 000) cm³ of NO react with (1 x 24 000) cm³ of oxygen to give (2 x 24 000) cm³ of NO₂

2 cm³ of NO react with 1 cm³ of oxygen to give 2 cm³ of NO₂

ie for gases only the reacting volume ratios are the same as the reacting mole ratios in the equation.

Example 1

What volume of sulphur trioxide would be produced by the complete reaction of 100 cm³ of sulphur dioxide with oxygen? What volume of oxygen would be needed to just react with the sulphur dioxide?



Thus 100 cm³ of sulphur dioxide will need 50 cm³ of oxygen and produce 100 cm³ of sulphur dioxide.



Example 2

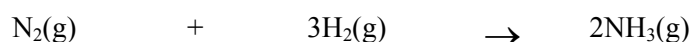
What would be the composition of the final product in Example 1 if 100 cm³ of oxygen had been used rather than 50 cm³?

Since 100 cm³ of the sulphur dioxide needs only 50 cm³ of oxygen there must be 50 cm³ of oxygen unused. Thus the final volume is:

$$100 \text{ cm}^3 \text{ of sulphur dioxide plus } 50 \text{ cm}^3 \text{ of excess oxygen} = 150 \text{ cm}^3$$

Example 3

What volume of ammonia would be produced if 10 cm³ of nitrogen was reacted with 20 cm³ of hydrogen?



You need to think before you start this question. The reacting volumes given in the question are not the same as those in the reaction. You must have excess of one of the gases.

From the equation:

10 cm³ of nitrogen needs 30 cm³ of hydrogen. You only have 20 cm³ of hydrogen so the nitrogen is in excess.

In this case you will need to use the hydrogen volume in the calculation.

	N ₂ (g)	+	3H ₂ (g)	→	2NH ₃ (g)
Ratios	1		3		2
	1/3		1		2/3
	1/3 x 20		20		2/3 x 20
	6.67 cm ³		20 cm ³		13.33 cm ³

Thus 20 cm³ of hydrogen will react to give **13.33 cm³ of ammonia** and there will be 3.33 cm³ of hydrogen left over.

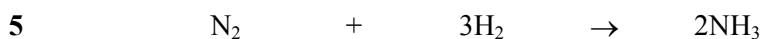
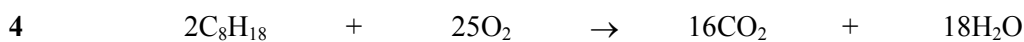
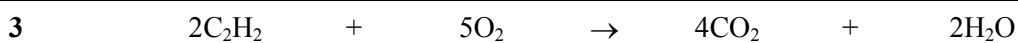
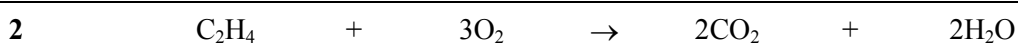
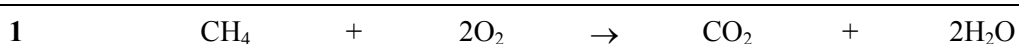


Exercise 9

Calculations based on equations involving only gases

Section a

In *Section a*, assume that you have 10 cm³ of the first named reactant and then calculate the volumes of all the gases involved in the equation. In these examples the reactions are being carried out at above 100°C and you should assume the water is present as a gas and therefore has a volume.





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Section b

In *Section b*, you are asked to find the total volume of gas produced at room temperature and pressure. You should ignore the volume of water produced as this will have condensed as a liquid. Be careful in some cases, as there is an excess of one of the reactants.

1 What volume of oxygen would be needed to convert 1000 cm^3 of nitrogen monoxide, NO, to nitrogen dioxide, NO_2 ? (Assume all volumes are measured at the same temperature and pressure.)

2 In the production of sulphuric acid sulphur dioxide is converted to sulphur trioxide by reaction with the oxygen in the air. What volume of air (assume 20% of the air is oxygen) would be needed to produce 150 cm^3 of sulphur trioxide? Assume complete conversion of sulphur dioxide to sulphur trioxide.

3 In equation for the oxidation of ammonia to nitrogen monoxide is:



What volume of ammonia would be required to produce 2500 cm^3 of nitrogen monoxide and what volume of air would be used in the conversion? Again assume that air is 20% oxygen by volume.

4 What volume of oxygen at room temperature and pressure would be needed to completely burn 1 mole of butane?

5 What volume of hydrogen at room temperature and pressure would be needed to convert 1 mole of ethene, C_2H_4 , to ethane, C_2H_6 ?

6 What is the final volume of gas produced at room temperature when 10 cm^3 of methane is burnt with 30 cm^3 of oxygen?

7 What is the final volume of gas produced at room temperature if 5 cm^3 of octane are burnt with 100 cm^3 of oxygen?



8 In a reaction between methane and oxygen 60 cm^3 of methane was burnt with 60 cm^3 of oxygen. What is the composition of the gas mixture produced?

9 What volume of ammonia would be produced if 10 cm^3 of nitrogen was reacted with 60 cm^3 of hydrogen?

10 What would be the final volume of gas produced in the reaction between 10 cm^3 of hydrogen and 10 cm^3 of oxygen?

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Section 10

Ions and ionic equations

Ionic theory

Many of the chemicals which you will use at GCE Advanced are ionic, that is the chemical bonds which hold the atoms together are ionic bonds. When you melt these compounds the ions are free to move and this gives them some special properties. Often, but not always, these chemicals are soluble in water and when they dissolve the ions separate to produce a solution containing positive and negative ions.

A few covalent substances also form ions when they dissolve in water. Some of these are extremely important: hydrogen chloride and sulphuric acid are examples.

Structures of ionic compounds

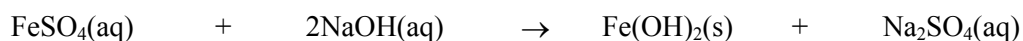
In your course you will study bonding and structure, but at a of the most important ideas are set out below.

- Ions are atoms or groups of atoms, which have a positive or negative electric charge.
- Positive ions are called cations (pronounced cat-ions) and negative ions are called anions (pronounced an-ions).
- Positive ions attract negative ions all around them and are firmly held in a rigid lattice; this is what makes ionic compounds solids.
- When an ionic compound is solid it is crystalline, but when it melts or is dissolved in water the ions become free and can move around.
- Ions have *completely* different properties from the atoms in them; chlorine is an extremely poisonous gas, but chloride ions are found in sodium chloride, which is essential to human life.

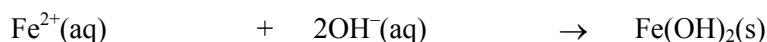
Ionic equations and spectator ions

Many of the chemicals, which you study are *ionic*, and in these cases it is the ions which react, not the molecules. For instance, copper(II) sulphate is usually written as CuSO_4 but it is more often the ion Cu^{2+} which reacts. When you write an ionic equation you include only the ions which **actually take part in the reaction**.

Let us look at molecular equation and see how it may be converted into an ionic equation. For example, look at the reaction between iron(II) sulphate solution and aqueous sodium hydroxide.

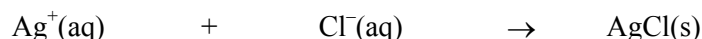


In water, the iron (II) sulphate and the sodium hydroxide are in the form of freely moving ions. When the two solutions are mixed together, we see a green precipitate of iron (II) hydroxide solid. Remaining in solution will be a mixture of sodium ions and sulphate ions.




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Also when silver nitrate solution reacts with sodium chloride solution the changes do not involve the nitrate ion from the silver nitrate or the sodium ion from the sodium chloride. These are referred to as 'spectator ions'. The equation for this reaction can be written



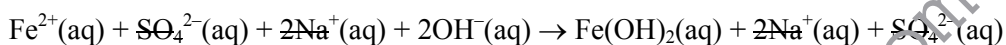
This equation represents the reaction between **any** aqueous solution containing silver ions and **any** aqueous solution containing chloride ions. This is the equation for the test for a chloride ion in solution.

You can work out an ionic equation as follows using the example of the reaction of iron(II) sulphate solution with excess sodium hydroxide solution.

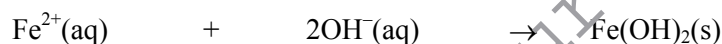
- 1 Write down the balanced molecular equation



- 2 Convert those chemicals that are ions **in solution** into their ions



- 3 Cross out those ions that appear on both sides of the equation as they have not changed during the reaction. They started in solution and they finished in the solution. To give the ionic equation:



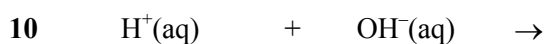
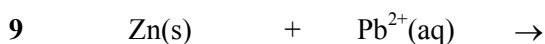
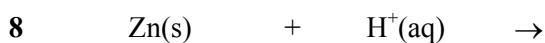
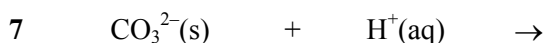
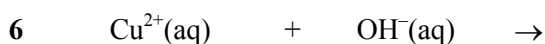
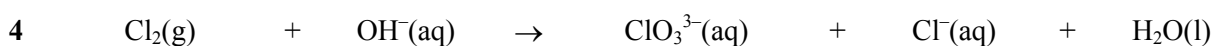
Check that the atoms **and** the charges balance.

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Exercise 10

Ionic equations

In questions 1–5 you are required to balance the equations, in questions 6–10 you are required to complete the equation and then balance it. For questions 1–17 you are required to write the full, balanced ionic equation. Questions 18–20 involve more complex ions again you are just asked to balance the equation.



-
- 11 Write an ionic equation for the reaction of magnesium with sulphuric acid.
-
- 12 Write an ionic equation for the reaction of sodium carbonate solution with nitric acid.
-
- 13 Write an ionic equation for the reaction of copper oxide with hydrochloric acid.
-
- 14 Write an ionic equation for the reaction of barium chloride solution with sodium sulphate solution.
-
- 15 Write an ionic equation for the reaction of silver nitrate solution with potassium chloride solution.
-
- 16 Write an ionic equation for the reaction of zinc with silver nitrate solution.
-
- 17 Write ionic equations for the reactions of sodium hydroxide and potassium hydroxide with hydrochloric acid.
-
- 18 Write ionic equations for the reactions of sodium hydroxide and potassium hydroxide with nitric acid.
-
- 19 Write ionic equations for the reactions of sodium hydroxide and potassium hydroxide with sulphuric acid.
-
- 20 What do you notice about the answers to questions 17, 18 and 19?
-



Section 11

Calculations involving chemicals in solution

These are often referred to as Volumetric Analysis. The name should not worry you, the basis of the calculations is the same as all the rest ie moles and equations.

Many reactions take place in solution involving solutions of known concentration.

Concentration in solution is generally measured as moles per 1000 cm³ **of solution**. For example the sodium chloride on the bench may be labelled as 1M NaCl. This means that each 1000 cm³ of the solution contains 1 Mole of NaCl (58.5 g).

It does not mean that 58.5 g of NaCl have been added to 1000 cm³ of water.

The solution will have been made up by measuring out 58.5 g of the solid, dissolving it in about 500 cm³ of water and then adding water to make the total volume of the mixture up to 1000 cm³. (1 dm³)

Concentration in mol dm⁻³ is called **molarity**.

$$\text{molarity} = \frac{\text{concentration in grams per 1000 cm}^3}{M_r \text{ for the material dissolved}}$$

$$\text{number of moles of material in a given volume} = \frac{\text{molarity} \times \text{volume (cm}^3\text{)}}{1000}$$

$$\text{mass of material in a given volume of solution} = \frac{\text{molarity} \times \text{volume (cm}^3\text{)} \times M_r}{1000}$$

In reactions in solution it is often more convenient to use molarity rather than g dm⁻³.

There are two ways you can approach calculations involving solutions. The first method (A) detailed below is really a short cut way of using the more detailed method B. Most of the straight forward calculations you will meet at the start of your course and the ones in this booklet can be carried through using method A.



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Method A

Consider the following reaction between two solutions



In this reaction a moles of substance A react with b moles of substance B

Let us suppose that $V_a \text{ cm}^3$ of the solution of A react with $V_b \text{ cm}^3$ of the solution of B. If this is an acid/alkali reaction we could find these volumes out using an indicator.

$$\therefore \text{Number of moles of A in } V_a \text{ cm}^3 \text{ of solution} = \frac{V_a M_a}{1000} = a$$

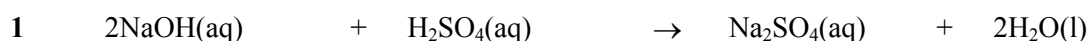
$$\therefore \text{Number of moles of B in } V_b \text{ cm}^3 \text{ of solution} = \frac{V_b M_b}{1000} = b$$

\therefore If we divide equation (i) by equation (ii) we get

$$\frac{V_a M_a}{V_b M_b} = \frac{a}{b}$$

This relationship will hold good for any reaction between two **solutions**.

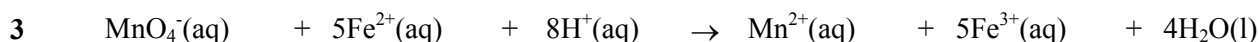
Examples



$$\frac{M_{\text{NaOH}} \times V_{\text{NaOH}}}{M_{\text{H}_2\text{SO}_4} \times V_{\text{H}_2\text{SO}_4}} = \frac{2}{1}$$



$$\frac{M_{\text{BaCl}_2} \times V_{\text{BaCl}_2}}{M_{\text{H}_2\text{SO}_4} \times V_{\text{H}_2\text{SO}_4}} = \frac{1}{1}$$



$$\frac{M_{\text{MnO}_4^-} \times V_{\text{MnO}_4^-}}{M_{\text{Fe}^{2+}} \times V_{\text{Fe}^{2+}}} = \frac{1}{5}$$

Calculation examples

- 1 What is the molarity of a solution of NaOH which contains 4 g of NaOH in 250 cm³ of solution?

$$M_r \text{ NaOH} = 40 \text{ g mol}^{-1}$$

$$4 \text{ g per } 250 \text{ cm}^3 = 16 \text{ g per } 1000 \text{ cm}^3$$

$$\therefore \text{ molarity} = \frac{16}{40} = \mathbf{0.040 \text{ mol dm}^{-3}}$$

This can also be written

$$\text{ molarity} = \frac{(4 \times 1000) \times 1}{250 \times 40} = \mathbf{0.040 \text{ mol dm}^{-3}}$$

- 2 What mass of KMnO₄ would be needed to prepare 250 cm³ of 0.020 mol dm⁻³ KMnO₄ solution? (M_r = 158)

$$1000 \text{ cm}^3 \text{ of } 0.020 \text{ mol dm}^{-3} \text{ KMnO}_4 \text{ will need } 158 \times 0.02 \text{ g}$$

$$\therefore 250 \text{ cm}^3 \text{ will need } \frac{158 \times 0.02 \times 250}{1000} = \mathbf{0.79 \text{ g}}$$

- 3 How many moles of H₂SO₄ will be contained in 25 cm³ of 0.10 mol dm⁻³ H₂SO₄?

$$\text{ number of moles} = \frac{0.10 \times 25}{1000} = \mathbf{0.0025 \text{ moles}}$$

- 4 25 cm³ of 0.10 mol dm⁻³ NaOH react with 50 cm³ of a solution of H₂SO₄. What is the molarity of the H₂SO₄?



$$\frac{M_{\text{NaOH}} \times V_{\text{NaOH}}}{M_{\text{H}_2\text{SO}_4} \times V_{\text{H}_2\text{SO}_4}} = \frac{2}{1}$$

$$\therefore \frac{0.1 \times 25}{M_{\text{H}_2\text{SO}_4} \times 50} = 2$$

$$\therefore M_{\text{H}_2\text{SO}_4} = \frac{0.1 \times 25}{2 \times 50} = \mathbf{0.025 \text{ mol dm}^{-3}}$$

NB If you are required to calculate the concentration in g dm⁻³ at this stage you need to multiply by the M_r of the material. In this case 98 g mol⁻¹



- 5 What volume of 0.02 mol dm^{-3} KMnO_4 solution will be needed to react with 25 cm^3 of 0.1 mol dm^{-3} Iron(II) ammonium sulphate ?

NB in Iron(II) ammonium sulphate only the iron(II) ions react with the manganate(VII) ions



$$\frac{M_{\text{MnO}_4^-} \times V_{\text{MnO}_4^-}}{M_{\text{Fe}^{2+}} \times V_{\text{Fe}^{2+}}} = \frac{1}{5}$$

$$\therefore \frac{0.02 \times V_{\text{MnO}_4^-}}{0.1 \times 25} = \frac{1}{5}$$

$$\therefore V_{\text{MnO}_4^-} = \frac{0.1 \times 25}{0.02 \times 5} = 25 \text{ cm}^3$$

- 6 25 cm^3 of a solution of 0.05 mol dm^{-3} silver nitrate react with 10 cm^3 of a solution of NaCl. What is the concentration of NaCl in g dm^{-3} in the solution?



$$\frac{M_{\text{NaCl}} \times V_{\text{NaCl}}}{M_{\text{AgNO}_3} \times V_{\text{AgNO}_3}} = \frac{1}{1}$$

$$\therefore \frac{10 \times M_{\text{NaCl}}}{25 \times 0.05} = \frac{1}{1}$$

$$\therefore M_{\text{NaCl}} = \frac{25 \times 0.05}{10} = 0.125 \text{ mol dm}^{-3}$$

$$\therefore \text{concentration of NaCl} = 0.125 \text{ mol dm}^{-3} \times 58.5 = 7.31 \text{ g dm}^{-3}$$


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- 7 In the reaction between an acid H_xA and 0.1 mol dm^{-3} NaOH solution. 25 cm^3 of a solution of 0.1 mol dm^{-3} H_xA react with 50 cm^3 of the 0.1 mol dm^{-3} NaOH. What is the value of x ?

This is not as difficult as it looks. You need to think what the equation for the reaction would be.



$$\therefore \frac{M_{H_xA} \times V_{H_xA}}{M_{NaOH} \times V_{NaOH}} = \frac{1}{x}$$

$$\therefore \frac{25 \times 0.1}{50 \times 0.1} = \frac{1}{x}$$

$$\therefore x = 2$$

Thus the acid is H_2A .

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Method B

In this method the actual amounts of materials in the volumes involved are calculated rather than the ratios.

Example

25 cm³ of 0.10 mol dm⁻³ NaOH react with 50 cm³ of a solution of H₂SO₄.

What is the molarity of the H₂SO₄?



∴ 2 mol of NaOH react with 1 mol of H₂SO₄

In this case you know the concentration of the sodium hydroxide so

∴ 1 mol of NaOH reacts with 0.5 mol of H₂SO₄

always put the reactant you know as '1 mol'

In this reaction you have used 25 cm³ of 0.10 mol dm⁻³ NaOH

$$= \frac{25 \times 0.10}{1000} \text{ mol of NaOH}$$

$$= 2.5 \times 10^{-3} \text{ mol}$$

This will react with $0.5 \times 2.5 \times 10^{-3}$ moles of H₂SO₄

$$= 1.25 \times 10^{-3} \text{ moles of H}_2\text{SO}_4$$

∴ 1.25×10^{-3} moles of H₂SO₄ will be found in 50 cm³ of the solution

∴ In 1000 cm³ of the acid the same solution there will be

$$= \frac{1000 \times (1.25 \times 10^{-3})}{50} \text{ moles of H}_2\text{SO}_4$$

$$= 0.0250 \text{ moles}$$

∴ The concentration of the sulphuric acid is **0.025 mol dm⁻³**.



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Exercise 11a

Calculations based on concentrations in solution

Calculate the number of moles of the underlined species in the volume of solution stated.

1 25 cm³ of 1.0 mol dm⁻³ HCl

2 50 cm³ of 0.5 mol dm⁻³ HCl

3 250 cm³ of 0.25 mol dm⁻³ HCl

4 500 cm³ of 0.01 mol dm⁻³ HCl

5 25 cm³ of 1.0 mol dm⁻³ NaOH

6 50 cm³ of 0.5 mol dm⁻³ KOH

7 50 cm³ of 0.25 mol dm⁻³ HNO₃

8 100 cm³ of 0.1 mol dm⁻³ H₂SO₄

9 25 cm³ of 0.05 mol dm⁻³ KMnO₄

10 25 cm³ of 0.2 mol dm⁻³ FeSO₄

Calculate the mass of material in the given volume of solution

11 25 cm³ of 1 mol dm⁻³ HCl

12 50 cm³ of 0.5 mol dm⁻³ NaCl

13 100 cm³ of 0.25 mol dm⁻³ NH₄NO₃

14 100 cm³ of 0.1 mol dm⁻³ AgNO₃

15 25 cm³ of 1 mol dm⁻³ BaCl₂

16 50 cm³ of 0.2 mol dm⁻³ H₂SO₄

17 20 cm³ of 0.1 mol dm⁻³ NaOH

18 50 cm³ of 0.1 mol dm⁻³ K₂CrO₄

19 25 cm³ of 0.02 mol dm⁻³ KMnO₄

20 25 cm³ of 0.1 mol dm⁻³ Pb(NO₃)₂



What is the concentration in moles dm^{-3} of the following?

21 3.65 g of HCl in 1000 cm^3 of solution

22 3.65 g of HCl in 100 cm^3 of solution

23 6.62 g of $\text{Pb}(\text{NO}_3)_2$ in 250 cm^3 of solution

24 1.00 g of NaOH in 250 cm^3 of solution

25 1.96 g of H_2SO_4 in 250 cm^3 of solution

26 1.58 g of KMnO_4 in 250 cm^3 of solution

27 25.0 g of $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$ in 250 cm^3 of solution

28 25.0 g of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in 250 cm^3 of solution

29 4.80 g of $(\text{COOH})_2 \cdot 2\text{H}_2\text{O}$ in 250 cm^3 of solution

30 10.0 g of $\text{FeSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$ in 250 cm^3 of solution

31 240 cm^3 of $\text{NH}_3(\text{g})$ dissolved in 1000 cm^3 of solution

32 480 cm^3 of $\text{HCl}(\text{g})$ dissolved in 100 cm^3 of solution

33 120 cm^3 of $\text{SO}_2(\text{g})$ dissolved in 250 cm^3 of solution

34 24 cm^3 of $\text{HCl}(\text{g})$ dissolved in 200 cm^3 of solution

35 100 cm^3 of $\text{NH}_3(\text{g})$ dissolved in 10 cm^3 of solution



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Exercise 11b

Simple volumetric calculations

In this series of calculations you should start by writing the equation for the reaction taking place then generate the molarity/volume ratio. In some cases you will need to calculate the molarity of the solutions before you start the main part of the question.

For questions 1–10 calculate the molarity of the first named solution from the data below.

1	25 cm ³ of sodium hydroxide	reacts with	21.0 cm ³ of 0.2 mol dm ⁻³ HCl
2	25 cm ³ of sodium hydroxide	reacts with	17.0 cm ³ of 0.1 mol dm ⁻³ H ₂ SO ₄
3	20 cm ³ of hydrochloric acid	reacts with	23.6 cm ³ of 0.1 mol dm ⁻³ NaOH
4	20 cm ³ of hydrochloric acid	reacts with	20.0 cm ³ of a solution of NaOH containing 40 g dm ⁻³ of NaOH
5	25 cm ³ of nitric acid	reacts with	15 cm ³ of a solution of 0.2 mol dm ⁻³ NH ₄ OH
6	25 cm ³ of a solution of barium chloride	reacts with	20 cm ³ of a solution of 0.05 mol dm ⁻³ sulphuric acid
7	25 cm ³ of a solution of NaCl	reacts with	10 cm ³ of a 0.02 mol dm ⁻³ silver nitrate
8	10 cm ³ of a solution of AlCl ₃	reacts with	30 cm ³ of 0.01 mol dm ⁻³ silver nitrate
9	25 cm ³ of H _x A	reacts with	25 cm ³ of 0.2 mol dm ⁻³ NaOH to give Na ₂ A
10	25 cm ³ of H ₃ PO ₄	reacts with	100 cm ³ of 0.1 mol dm ⁻³ NaOH to give NaH ₂ PO ₄

-
- 11 25 cm³ of a solution of 0.1 mol dm⁻³ NaOH reacts with 50 cm³ of a solution of hydrochloric acid. What is the molarity of the acid?
-
- 12 25 cm³ of a solution of 0.2 mol dm⁻³ KOH reacts with 30 cm³ of a solution of nitric acid. What is the concentration of the acid in moles dm⁻³?
-
- 13 In a titration 25 cm³ of ammonia solution react with 33.30 cm³ of 0.1 mol dm⁻³ HCl. What is the concentration of the ammonia solution in g dm⁻³?
-
- 14 In the reaction between iron(II) ammonium sulphate and potassium manganate(VII) solution. 25 cm³ of the Fe²⁺ solution reacted with 24.8 cm³ of 0.020 mol dm⁻³ KMnO₄ solution. What is the molarity of the iron(II) ammonium sulphate solution?
-
- 15 10 cm³ of a solution of NaCl react with 15 cm³ of 0.02 mol dm⁻³ silver nitrate solution. What is the concentration of the NaCl solution in g dm⁻³?
-
- 16 25 cm³ of a solution of an acid H_xA containing 0.1 mol dm⁻³ of the acid in each 1000 cm³ of solution reacts with 75 cm³ of a solution of 0.1 mol dm⁻³ NaOH. What is the value of x?
-
- 17 25 cm³ of a solution of sodium carbonate react with 10 cm³ of a 0.1 mol dm⁻³ HCl. What is the concentration of the sodium carbonate?
-
- 18 What volume of 0.1 mol dm⁻³ HCl will be needed to react with 25 cm³ of 0.2 mol dm⁻³ NaOH?
-
- 19 What volume of 0.05 mol dm⁻³ H₂SO₄ will be needed to react with 25 cm³ of 0.2 mol dm⁻³ NaOH?
-
- 20 What volume of 0.02 mol dm⁻³ KMnO₄ will be needed to react with 25 cm³ of 0.1 mol dm⁻³ FeSO₄ solution?
-



The last five questions will require you to use the skills you have learnt in this section, together with those from other sections.

21 What weight of silver chloride will be produced if 25 cm³ of 0.1 mol dm⁻³ silver nitrate is added to excess sodium chloride solution?

22 What weight of calcium carbonate will dissolve in 100 cm³ of 0.2 mol dm⁻³ HCl?

23 What volume of carbon dioxide will be produced if 100 cm³ of 0.2 mol dm⁻³ HNO₃ is added to excess sodium carbonate solution?

24 What weight of magnesium will dissolve in 10 cm³ of 1 mol dm⁻³ HCl and what volume of hydrogen will be produced?

25 What volume of ammonia gas will be produced in the following reaction if 50 cm³ of 0.5 mol dm⁻³ sodium hydroxide is boiled with 50 cm³ of 0.4 mol dm⁻³ ammonium chloride solution? (Care: one of these is in excess.)





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Section 12

Data – The periodic table

THE PERIODIC TABLE

Period 1 2 3 4 5 6 7 8

Period	Group							
	1	2	3	4	5	6	7	8
1	1 H Hydrogen 1							2 He Helium 4
2	3 Li Lithium 7	4 Be Beryllium 9						9 F Fluorine 19
3	11 Na Sodium 23	12 Mg Magnesium 24						16 O Oxygen 16
4	19 K Potassium 39	20 Ca Calcium 40	21 Sc Scandium 45	22 Ti Titanium 48	23 V Vanadium 51	24 Cr Chromium 52	25 Mn Manganese 55	26 Fe Iron 56
5	37 Rb Rubidium 85	38 Sr Strontium 88	39 Y Yttrium 89	40 Zr Zirconium 91	41 Nb Niobium 93	42 Mo Molybdenum 96	43 Tc Technetium (99)	44 Ru Ruthenium 101
6	55 Cs Caesium 133	56 Ba Barium 137	57 La Lanthanum 139	72 Hf Hafnium 178	73 Ta Tantalum 181	74 W Tungsten 184	75 Re Rhenium 186	76 Os Osmium 190
7	87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Unq Unnilquadium (261)	105 Unp Unnilpentium (262)	106 Unh Unnilhexium (263)	107 Uuh Unnilheptium (264)	108 Uuo Unniloctium (265)

key

1	Atomic Number
Symbol	
Name	
Molar mass in g mol ⁻¹	

58 Ce Cerium 140	59 Pr Praseodymium 141	60 Nd Neodymium 144	61 Pm Promethium (147)	62 Sm Samarium 150	63 Eu Europium 152	64 Gd Gadolinium 157	65 Tb Terbium 159	66 Dy Dysprosium 163	67 Ho Holmium 165	68 Er Erbium 167	69 Tm Thulium 169	70 Yb Ytterbium 173	71 Lu Lutetium 175
90 Th Thorium 232	91 Pa Protactinium (231)	92 U Uranium 238	93 Np Neptunium (237)	94 Pu Plutonium (242)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (245)	98 Cf Californium (251)	99 Es Einsteinium (254)	100 Fm Fermium (253)	101 Md Mendelevium (256)	102 No Nobelium (254)	103 Lr Lawrencium (257)



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Answers



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Exercise 1

1	18	21	11	41	159.5
2	44	22	164	42	161.4
3	17	23	74	43	170
4	46	24	136	44	53.5
5	28	25	208	45	132
6	64	26	1335	46	117.0
7	80	27	213	47	122.5
8	81	28	342	48	166.0
9	98	29	152	49	74.5
10	63	30	127	50	69.0
11	58.5	31	162.5	51	249.5
12	85	32	400	52	278
13	106	33	223	53	964
14	40	34	239	54	248
15	142	35	685	55	126
16	158	36	331	56	246
17	194	37	278	57	2635
18	100	38	303	58	60
19	166	39	99.0	59	58
20	195	40	134.5	60	122



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Exercise 2

1 NaCl	21 BaSO ₄	41 PbCO ₃	61 PCl ₃
2 NaOH	22 AlCl ₃	42 PbO	62 PCl ₅
3 Na ₂ CO ₃	23 Al ₂ O ₃	43 PbO ₂	63 P ₂ O ₃
4 Na ₂ SO ₄	24 Al(OH) ₃	44 PbCl ₂	64 P ₂ O ₅
5 NO ₃ PO ₄	25 Al ₂ (SO ₄) ₃	45 PbCl ₄	65 H ₃ PO ₄
6 KCl	26 CuSO ₄	46 PbS	66 H ₂ SO ₄
7 KBr	27 CuO	47 SnCl ₂	67 HNO ₃
8 KI	28 CuCl ₂	48 SnCl ₄	68 HCl
9 KHCO ₃	29 Cu(NO ₃) ₂	49 FeSO ₄	69 CCl ₄
10 KNO ₂	30 Cu ₂ O	50 FeCl ₂	70 SiCl ₄
11 MgCl ₂	31 CuCl	51 Fe ₂ (SO ₄) ₃	71 SiO ₂
12 Mg(NO ₃) ₂	32 Zn(NO ₃) ₂	52 FeCl ₃	72 SO ₂
13 Mg(OH) ₂	33 ZnCO ₃	53 Fe(OH) ₃	73 SO ₃
14 MgO	34 ZnO	54 Fe(OH) ₂	74 H ₂ S
15 MgCO ₃	35 AgCl	55 NH ₄ Cl	75 Cl ₂ O
16 CaO	36 AgBr	56 (NH ₄) ₂ CO ₃	76 NO ₂
17 CaCl ₂	37 AgI	57 NH ₄ OH	77 NO
18 CaSO ₄	38 AgNO ₃	58 NH ₄ NO ₃	78 CO ₂
19 CaCO ₃	39 Ag ₂ O	59 (NH ₄) ₂ SO ₄	79 CO
20 BaCl ₂	40 Pb(NO ₃) ₂	60 (NH ₄) ₃ PO ₄	80 HOH/H ₂ O



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Exercise 3

- 1 Water
- 2 Carbon dioxide
- 3 Ammonia
- 4 Oxygen
- 5 Hydrogen
- 6 Sulphur dioxide (or IV oxide)
- 7 Sulphur trioxide (or VI oxide)
- 8 Hydrogen chloride
- 9 Hydrogen iodide
- 10 Hydrogen fluoride
- 11 Methane
- 12 Hydrogen sulphide
- 13 Hydrogen bromide
- 14 Sulphuric acid
- 15 Nitric acid
- 16 Sodium chloride
- 17 Sodium nitrate
- 18 Sodium carbonate
- 19 Sodium hydroxide
- 20 Sodium sulphate
- 21 Calcium chloride
- 22 Calcium nitrate
- 23 Calcium hydroxide
- 24 Calcium sulphate
- 25 Barium chloride
- 26 Aluminium chloride
- 27 Aluminium nitrate
- 28 Aluminium sulphate
- 29 Iron(II) sulphate
- 30 Iron(II)chloride
- 31 Iron(III) chloride
- 32 Iron(III) sulphate
- 33 Lead(II) oxide
- 34 Lead(IV) oxide
- 35 Lead(II) nitrate
- 36 Lead(II) chloride
- 37 Lead (II) sulphate
- 38 Copper(II) nitrate
- 39 Copper(I) chloride
- 40 Copper(II) chloride
- 41 Copper(II) sulphate
- 42 Zinc chloride
- 43 Silver nitrate
- 44 Ammonium chloride
- 45 Ammonium sulphate
- 46 Ammonium vanadate(V)
- 47 Potassium chlorate(V)
- 48 Potassium iodate
- 49 Sodium chlorate(I)
- 50 Sodium nitrite
- 51 Ethane
- 52 Butane
- 53 Octane
- 54 Ammonium carbonate
- 55 Potassium manganate(VII)
- 56 Potassium chromate(VI)
- 57 Potassium hydrogencarbonate
- 58 Potassium iodide
- 59 Cobalt(II) nitrate
- 60 Potassium astatide



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Exercise 4a

1	0.50	26	0.10
2	2.0	27	0.10
3	0.10	28	0.0085
4	5.0	29	0.26
5	20	30	0.104
6	0.010	31	0.20
7	1.0	32	0.082
8	0.22	33	0.050
9	0.0010	34	1.34
10	0.050	35	0.025
11	0.33	36	0.204
12	0.25	37	0.071
13	0.021	38	0.010
14	0.020	39	0.050
15	0.125	40	0.254
16	0.020	41	0.0125
17	0.167	42	0.152
18	1.0	43	0.10
19	0.046	44	0.053
20	0.020	45	0.0043
21	0.0010	46	0.036
22	0.25	47	0.266
23	0.02	48	0.024
24	0.0025	49	0.025
25	0.20	50	1.574



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Exercise 4b

- | | | | |
|----|---------|----|---------|
| 1 | 36 g | 26 | 14.95 g |
| 2 | 132 g | 27 | 76.2 g |
| 3 | 47.6 g | 28 | 10.03 g |
| 4 | 23 g | 29 | 17.82 g |
| 5 | 33.6 g | 30 | 145.2 g |
| 6 | 40.96 g | 31 | 2.925 g |
| 7 | 240 g | 32 | 12.25 g |
| 8 | 81 g | 33 | 21.4 g |
| 9 | 1.152 g | 34 | 745 g |
| 10 | 9.45 g | 35 | 0.069 g |
| 11 | 26.3 g | 36 | 49.9 g |
| 12 | 59.5 g | 37 | 27.8 g |
| 13 | 11.66 g | 38 | 4.82 g |
| 14 | 80.0 g | 39 | 9.92 g |
| 15 | 127.8 g | 40 | 302.4 g |
| 16 | 7.9 g | 41 | 756.5 g |
| 17 | 34.92 g | 42 | 39.53 g |
| 18 | 90 g | 43 | 10.2 g |
| 19 | 249 g | 44 | 11.6 g |
| 20 | 23.4 g | 45 | 9.76 g |
| 21 | 12.2 g | 46 | 4.34 g |
| 22 | 672.4 g | 47 | 9.59 g |
| 23 | 0.296 g | 48 | 41.0 g |
| 24 | 13.6 g | 49 | 304 g |
| 25 | 43.68 g | 50 | 1397 g |

Exercise 4c

- | | | | |
|----|------------------------|----|------------------------|
| 1 | 24000 cm ³ | 11 | 134.4 cm ³ |
| 2 | 2400 cm ³ | 12 | 216 cm ³ |
| 3 | 12000 cm ³ | 13 | 960 cm ³ |
| 4 | 48000 cm ³ | 14 | 2952 cm ³ |
| 5 | 2880 cm ³ | 15 | 55.2 cm ³ |
| 6 | 81600 cm ³ | 16 | 192000 cm ³ |
| 7 | 2640 cm ³ | 17 | 0.24 cm ³ |
| 8 | 96 cm ³ | 18 | 144000 cm ³ |
| 9 | 240000 cm ³ | 19 | 182.4 cm ³ |
| 10 | 10800 cm ³ | 20 | 72000 cm ³ |



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Exercise 4d

- | | | | |
|----|------------|----|--------------|
| 1 | 0.0083 mol | 11 | 0.0292 mol |
| 2 | 0.0208 mol | 12 | 0.2333 mol |
| 3 | 0.0416 mol | 13 | 0.0917 mol |
| 4 | 0.0533 mol | 14 | 0.0088 mol |
| 5 | 0.0098 mol | 15 | 0.0333 mol |
| 6 | 0.0094 mol | 16 | 0.0033 mol |
| 7 | 0.0106 mol | 17 | 0.000080 mol |
| 8 | 0.0033 mol | 18 | 0.8333 mol |
| 9 | 0.0833 mol | 19 | 0.0175 mol |
| 10 | 0.10 mol | 20 | 0.0375 mol |

Exercise 4e

- | | | | |
|----|---------|----|---------|
| 1 | 0.367 g | 11 | 0.875 g |
| 2 | 0.354 g | 12 | 10.27 g |
| 3 | 1.166 g | 13 | 2.38 g |
| 4 | 5.333 g | 14 | 0.263 g |
| 5 | 0.78 g | 15 | 1.217 g |
| 6 | 0.763 g | 16 | 0.270 g |
| 7 | 0.757 g | 17 | 0.011 g |
| 8 | 0.233 g | 18 | 38.33 g |
| 9 | 0.167 g | 19 | 0.683 g |
| 10 | 3.20 g | 20 | 1.05 g |

Exercise 4f

- | | | | |
|----|-------------------------|----|-----------------------|
| 1 | 1091 cm ³ | 11 | 56000 cm ³ |
| 2 | 7059 cm ³ | 12 | 30545 cm ³ |
| 3 | 8571 cm ³ | 13 | 20308 cm ³ |
| 4 | 7500 cm ³ | 14 | 16000 cm ³ |
| 5 | 702 cm ³ | 15 | 5260 cm ³ |
| 6 | 670 cm ³ | 16 | 2370 cm ³ |
| 7 | 3380 cm ³ | 17 | 375 cm ³ |
| 8 | 30000 cm ³ | 18 | 12000 cm ³ |
| 9 | 2400000 cm ³ | 19 | 26526 cm ³ |
| 10 | 180000 cm ³ | 20 | 77143 cm ³ |



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Exercise 4g

1	160	11	34
2	64	12	17
3	80	13	38
4	71	14	28
5	2.0	15	44
6	28	16	32
7	30	17	211
8	58	18	36.5
9	32	19	81
10	28	20	128

Exercise 5

Section (a)

1	CaCO_3	4	$\text{N}_2\text{H}_4\text{S}_2\text{O}_8$
2	Na_2SO_4	5	P_4O_{10}
3	$\text{Na}_2\text{S}_2\text{O}_3$	6	$\text{C}_2\text{H}_4\text{O}_2 - \text{CH}_3\text{COOH}$
4	PbO	7	C_4H_{10}
5	Pb_3O_4	8	Fe_2O_3
6	H_3PO_3	9	$\text{H}_2\text{S}_2\text{O}_8$
7	H_2SO_3	10	C_6H_6
8	CH_4		
9	C_3H_8		
10	HO (giving H_2O_2)		
11	$\text{H}_4\text{N}_2\text{O}_3 (\text{NH}_4\text{NO}_3)$		
12	$\text{FeSO}_{11}\text{H}_{14} (\text{FeSO}_4 - 7\text{H}_2\text{O})$		

Section (b)

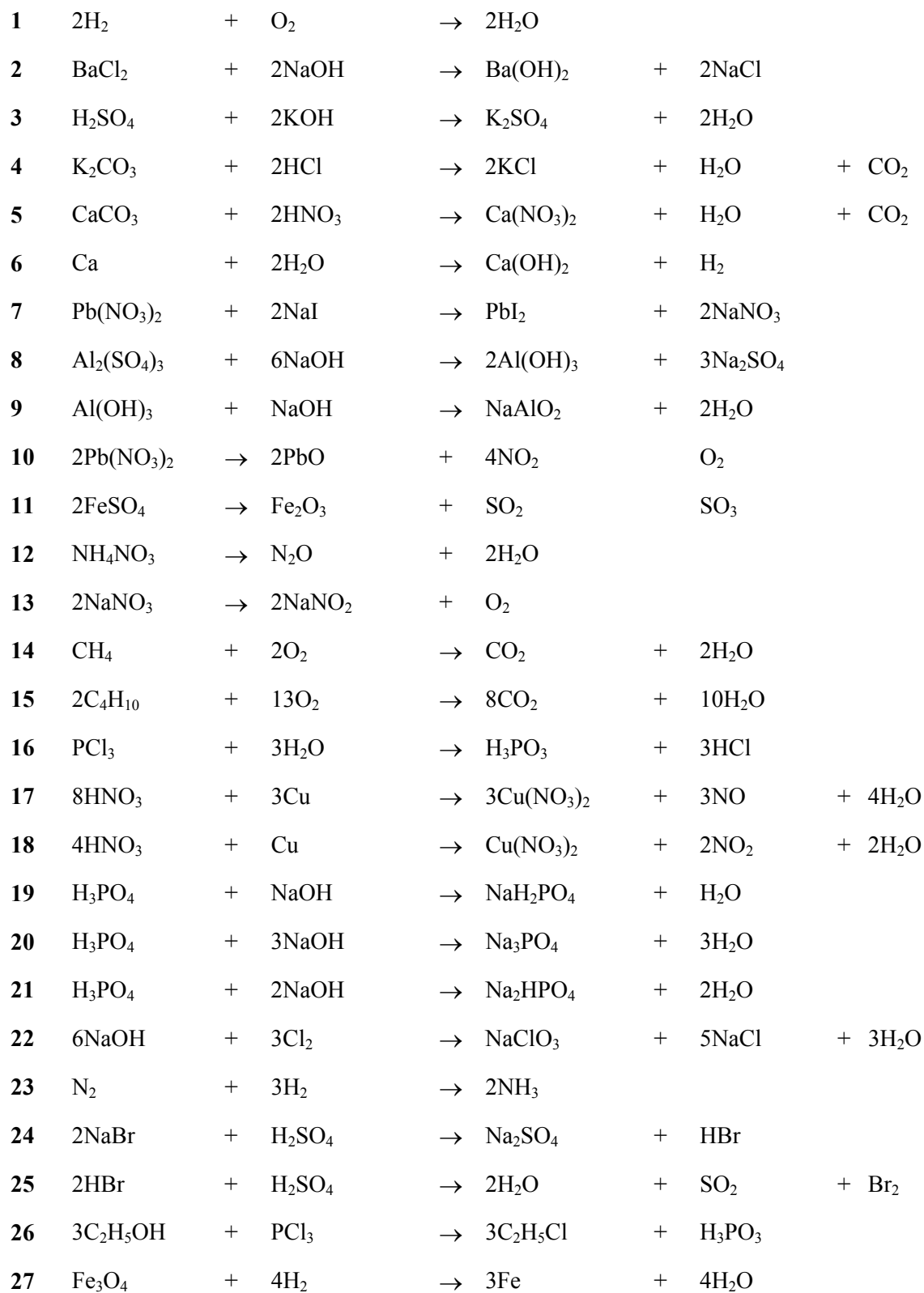
1	C_2H_4		
2	C_3H_6		
3	P_2I_4		

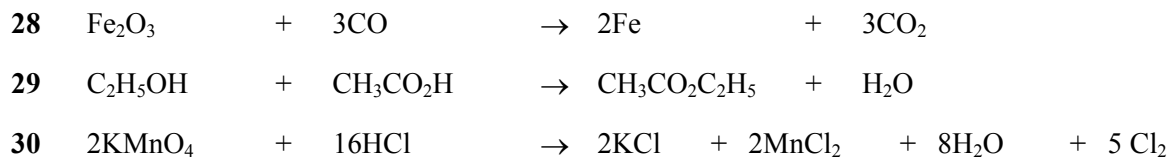
Section (c)

1	PbO
2	KO_2
3	C_2H_2
4	AlCl_3
5	CH_4
6	yes
7	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
8	$\text{PCl}_5, \text{PCl}_3, \text{Cl}_2$



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Exercise 6a



Exercise 6b

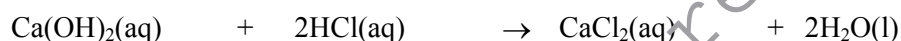
1 Hydrogen is not H but H₂, which gives



2 Since the valency of lead is 2 not 1 lead nitrate is not PbNO₃ but Pb(NO₃)₂ and also lead chloride is PbCl₂



3 Calcium hydroxide is Ca(OH)₂



4 This does not balance.



5 A magnesium compound cannot give a calcium compound!

6 Ozone O₃ is not produced by heating a nitrate O₂ is.

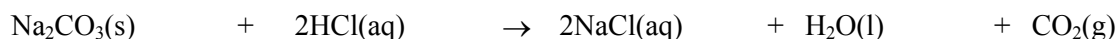


7 This reaction does not take place and so no equation can be written.

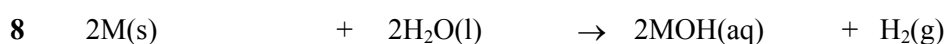
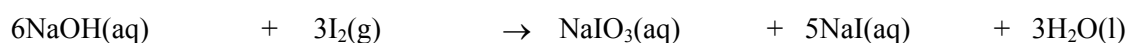
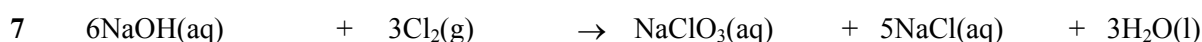
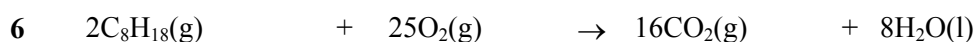
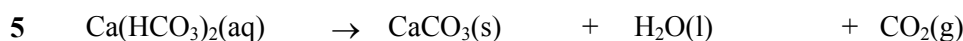
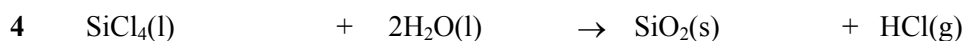
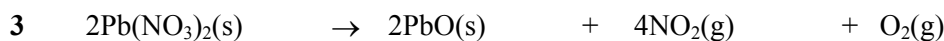
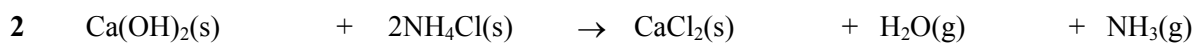
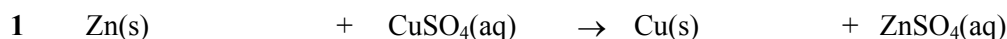
8 Aluminium has a valency of 3 not 2 as in this equation.



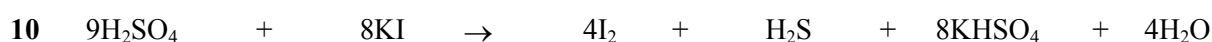
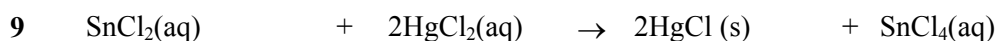
9 Sodium has a valency of 1 not 2 as in this equation



10 Silver chloride is not soluble in water. Thus the AgCl needs a (s) symbol

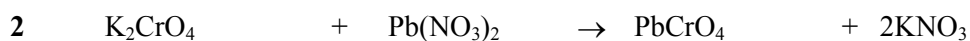
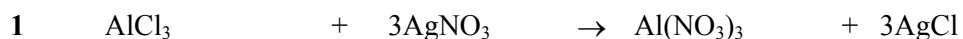
Exercise 6c

Where M = Li, Na, K, Rb or Cs





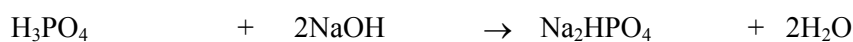
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Exercise 7

4 i) 1 mole

ii) 2 moles

iii)



iv) 75 cm^3

5 $x = 3$

6 $x = 5$



9 It is





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Exercise 8

- 1 11.2 g
- 2 21.6 g
- 3 0.682 g of ethanoic acid and 0.523 g of ethanol
- 4 143 tonnes
- 5 14.5 g
- 6 8.0 g of sodium hydroxide, 9.75 g of copper hydroxide
- 7 12000 cm³
- 8 54.7 g of calcium nitrate, 8.0 dm³ of carbon dioxide
- 9 6 dm³ total (4.8 dm³ of nitrogen dioxide and 1.2 dm³ of oxygen)
- 10 $\text{Mg} + \text{H}_2\text{SO}_4 + 7\text{H}_2\text{O} \rightarrow \text{Mg SO}_4 \cdot 7\text{H}_2\text{O} + \text{H}_2$
41.0 g
- 11 31.9 g
- 12 324.3 g
- 13 5.11 g of ethanol, 2.67 dm³ of carbon dioxide
- 14 (i) 12.30 g of zinc hydroxide
(ii) 9.12 g of aluminium hydroxide
(iii) 9.67 g of magnesium hydroxide
- 15 0.600 dm³
- 16 0.100 g
- 17 2.94 g of sodium chloride, 1.065 g of sodium chlorate(v)
- 18 4.15×10^6 dm³ of nitrogen, 12.5×10^6 dm³ of hydrogen
- 19 63 tonnes of nitric acid, 4.8×10^7 dm³ of oxygen
- 20 2198 g of calcium carbonate, 4.395 dm³ of 10M HCl



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Exercise 9

Section (a)

- | | | | |
|---|-----------------------------------|-----------------------------------|--|
| 1 | 20cm ³ O ₂ | 10cm ³ CO ₂ | 20cm ³ H ₂ O (g) |
| 2 | 30cm ³ O ₂ | 20cm ³ CO ₂ | 20cm ³ H ₂ O (g) |
| 3 | 25cm ³ O ₂ | 20cm ³ CO ₂ | 10cm ³ H ₂ O (g) |
| 4 | 125cm ³ O ₂ | 80cm ³ CO ₂ | 90cm ³ H ₂ O (g) |
| 5 | 30cm ³ H ₂ | 20cm ³ NH ₃ | |

Section (b)

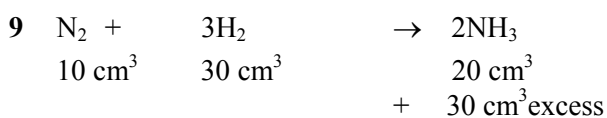
- 500cm³ O₂ (2NO + O₂ → 2NO₂)
- 375cm³ air (2SO₂ + O₂ → 2SO₃)
- 2500cm³ NH₃ needed $\frac{5}{4} \times 2500 = 3125\text{cm}^3 \text{ O}_2 \rightarrow 15625\text{cm}^3 \text{ air}$
- 6.5 x 24000cm³ = 156m³
- 24000cm³

6 Final volume = 20cm³ (10cm³ CO₂ + 10 cm³ unused O₂)

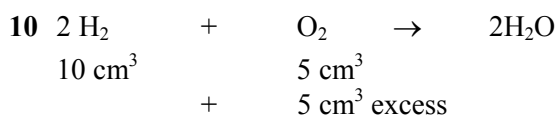
7 Final volume = 77.5cm³ (40cm³ CO₂ + 37.5cm³ used O₂)

- 8 This time the CH₄ is in excess. We must assume that CO₂ is produced (not CO or C)!

Final volume = 60 cm³ (30 cm³ CO₂ + 30 cm³ CH₄)



20 cm³ NH₃ produced + 30 cm³ excess H₂



Final volume = 5 cm³ (all excess O₂)



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Exercise 10

- 1 $\text{Pb}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Pb}(\text{OH})_2(\text{s})$
- 2 $\text{Al}^{3+}(\text{aq}) + 3\text{OH}^{-}(\text{aq}) \rightarrow \text{Al}(\text{OH})_3(\text{s})$
- 3 $\text{Al}(\text{OH})_3(\text{s}) + \text{OH}^{-}(\text{aq}) \rightarrow \text{AlO}_2^{-}(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
- 4 $\text{Cl}_2(\text{g}) + 6\text{OH}^{-}(\text{aq}) \rightarrow \text{ClO}_3^{-}(\text{aq}) + 5\text{Cl}^{-}(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$
- 5 $2\text{S}_2\text{O}_3^{2-}(\text{aq}) + \text{I}_2(\text{s}) \rightarrow \text{S}_4\text{O}_6^{2-}(\text{aq}) + 2\text{I}^{-}(\text{aq})$
- 6 $\text{Cu}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s})$
- 7 $\text{CO}_3^{2-}(\text{s}) + 2\text{H}^{+}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
- 8 $\text{Zn}(\text{s}) + 2\text{H}^{+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g})$
- 9 $\text{Zn}(\text{s}) + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Pb}(\text{s}) + \text{Zn}^{2+}(\text{aq})$
- 10 $\text{H}^{+}(\text{aq}) + \text{OH}^{-}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$
- 11 $\text{Mg}(\text{s}) + 2\text{H}^{+}(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{H}_2(\text{g})$
- 12 $\text{CO}_3^{2-}(\text{s}) + 2\text{H}^{+}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
- 13 $\text{CuO}(\text{s}) + 2\text{H}^{+}(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- 14 $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$
- 15 $\text{Ag}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \rightarrow \text{AgCl}(\text{s})$
- 16 $\text{Zn}(\text{s}) + 2\text{Ag}^{+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$
- 17–20 $2\text{OH}^{-}(\text{aq}) + \text{OH}^{-}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$

In every case the reaction is the same



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Exercise 11a

- | | |
|-----------------|---------------------------------|
| 1 0.025 moles | 19 0.079 g |
| 2 0.025 moles | 20 0.828 g |
| 3 0.0625 | 21 0.1 mol dm^{-3} |
| 4 0.005 moles | 22 1.0 mol dm^{-3} |
| 5 0.025 moles | 23 0.03 mol dm^{-3} |
| 6 0.025 moles | 24 0.1 mol dm^{-3} |
| 7 0.0125 moles | 25 0.03 mol dm^{-3} |
| 8 0.01 moles | 26 0.04 mol dm^{-3} |
| 9 0.00125 moles | 27 0.40 mol dm^{-3} |
| 10 0.005 moles | 28 0.40 mol dm^{-3} |
| 11 0.9125 g | 29 $0.152 \text{ mol dm}^{-3}$ |
| 12 1.463 g | 30 $0.0102 \text{ mol dm}^{-3}$ |
| 13 2 g | 31 0.01 mol dm^{-3} |
| 14 1.70 g | 32 0.2 mol dm^{-3} |
| 15 5.2 g | 33 0.02 mol dm^{-3} |
| 16 0.98 g | 34 $0.005 \text{ mol dm}^{-3}$ |
| 17 0.08 g | 35 $0.417 \text{ mol dm}^{-3}$ |
| 18 0.97 g | |

Exercise 11b

- | | |
|--------------------------------|--|
| 1 $0.168 \text{ mol dm}^{-3}$ | 16 3.0 |
| 2 $0.136 \text{ mol dm}^{-3}$ | 17 0.02 mol dm^{-3} |
| 3 $0.118 \text{ mol dm}^{-3}$ | 18 50 cm^3 |
| 4 1.0 mol dm^{-3} | 19 50 cm^3 |
| 5 0.12 mol dm^{-3} | 20 25 cm^3 |
| 6 $0.040 \text{ mol dm}^{-3}$ | 21 0.359 g |
| 7 $0.0080 \text{ mol dm}^{-3}$ | 22 1.0 g |
| 8 $0.010 \text{ mol dm}^{-3}$ | 23 240 cm^3 |
| 9 0.10 mol dm^{-3} | 24 0.12 g Mg
$120 \text{ cm}^3 \text{ H}_2$ |
| 10 0.40 mol dm^{-3} | 25 480 cm^3 |
| 11 $0.050 \text{ mol dm}^{-3}$ | |
| 12 $0.167 \text{ mol dm}^{-3}$ | |
| 13 2.26 g dm^{-3} | |
| 14 $0.099 \text{ mol dm}^{-3}$ | |
| 15 1.755 g dm^{-3} | |