

A-Level Chemistry Calculation Booklet

Name

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**You must complete this booklet and bring it with you
to your first class in September**

Section 1: Molar Mass

Exercise 1: calculation of the molar mass of compounds

Calculate the molar mass of the following compounds. You will find data concerning relative atomic masses on the periodic table of elements. When you have finished this set of calculations keep the answers for reference. You will find them useful for some of the other questions in this workbook.

1 H₂O

2 CO₂

3 NH₃

4 H₂SO₄

5 HNO₃

6 NaNO₃

7 Na₂CO₃

8 NaOH

9 Na₂SO₄

10 Al(NO₃)₃

11 Al₂(SO₄)₃

12 FeSO₄

13 Fe₂(SO₄)₃

14 PbO

15 PbO_2

16 Pb_3O_4

17 $\text{Pb}(\text{NO}_3)_2$

18 PbCl_2

19 PbSO_4

20 CuCl

21 CuCl_2

22 CuSO_4

23 ZnCl_2

24 AgNO_3

25 NH_4Cl

26 $(\text{NH}_4)_2\text{SO}_4$

27 $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

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, sulfur and oxygen
- how many atoms of each kind are in each molecule eg H_2SO_4 contains two atoms of hydrogen, one atom of sulfur 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 2g of hydrogen atoms and 16g 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 a large number of chemical formulae by heart. However, it is useful to know a few of them and then be able to work out the rest.

You can work out the formulae of compounds containing metals from the charges on the ions:

- metals in group 1 always have charge +1 in their compounds
- metals in group 2 always have charge +2 in their compounds
- metals in group 3 always have charge +3 in their compounds
- ions of group 7 elements have charge -1
- ions of group 6 elements have charge -2
- ions of group 5 elements have charge -3

In the compound, the number of positive and negative charges is equal so that the overall charge is zero.

Some metals form more than one ion, and this is shown by a roman numeral in the name. Iron(II) chloride contains Fe^{2+} ions so the compound is FeCl_2 . Iron(III) chloride contains Fe^{3+} ions so the compound is FeCl_3 .

Some ions have formulae which you cannot deduce from the periodic table, and you will need to learn these:

- OH^- hydroxide
- NO_3^- nitrate
- CO_3^{2-} carbonate
- SO_4^{2-} sulfate
- NH_4^+ ammonium

Compounds which do not contain metals have covalent bonds. The number of bonds a nonmetal can form depends on the number of electrons in its outer shell.

As a rule:

- carbon forms 4 bonds
- nitrogen forms 3 bonds
- phosphorus can form 3 or 5 bonds
- oxygen and sulfur form 2 bonds
- halogens form 1 bond

Here are a few examples:

- **Sodium sulfate**

The formula of a sodium ion is Na^+

The formula of a sulfate ion is SO_4^{2-}

There must be two sodium ions, each with charge 1+, to balance the two – charges on sulfate.

The formula with two Na^+ and one SO_4^{2-} is written Na_2SO_4

- **Calcium hydrogen carbonate** The formula of a calcium ion is Ca^{2+}

The formula of a hydrogen carbonate ion is HCO_3^-

There must be two hydrogen carbonate ions, each with charge 1–, to balance the two + charges on calcium.

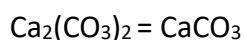
The formula with one Ca^{2+} and two HCO_3^- is written $\text{Ca}(\text{HCO}_3)_2$

Note: A bracket *must* be placed around a group or ion if it is multiplied by 2 or more *and/or* composed of more than one element. For example:

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 eg:



However, you should **not** do this for organic compounds. For example 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 with charge 1, ie Cu_2O . Lead(II) nitrate means use lead with charge 2, ie $\text{Pb}(\text{NO}_3)_2$

The periodic table can help you find the charge on an element and the number of bonds it can make, and hence the formula of its compounds.

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

On the next page you will find a table of the more common elements and ions that you may have met at GCSE level. Also included are some that you will meet in the first few weeks of your A-level course or that are mentioned in some of the calculations in this workbook.

Symbols and charges of common elements and ions

Elements	Symbol	Charge on ion	Ions	Symbol	Charge on ion
Aluminium	Al	+3	Ammonium	NH ₄	+1
Barium	Ba	+2	Carbonate	CO ₃	-2
Bromine	Br	-1	Hydrogencarbonate	HCO ₃	-1
Calcium	Ca	+2	Hydrogen-sulfate	HSO ₃	-1
Chlorine	Cl	-1	Hydroxide	OH	-1
Cobalt	Co	+2	Nitrate	NO ₃	-1
Copper	Cu	+1 and 2	Nitrite	NO ₂	-1
Hydrogen	H	+1	Sulfate	SO ₄	-2
Iodine	I	-1	Sulfite	SO ₃	-2
Iron	Fe	+2 and 3	Chlorate(I)	ClO	-1
Lead	Pb	+2 and 4	Chlorate(V)	ClO ₃	-1
Magnesium	Mg	+2	Vanadate(V)	VO ₃	-1
Manganese	Mn	+2 and 4	Manganate(VII)	MnO ₄	-1
Mercury	Hg	+1 and 2	Chromate(VI)	CrO ₄	-2
Nitrogen	N	3 and 5	Dichromate(VI)	Cr ₂ O ₇	-2
Oxygen	O	-2			
Potassium	K	+1			
Silver	Ag	+1			
Sodium	Na	+1			

The number of covalent bonds normally formed by an element

Element	Number of bonds
Hydrogen	1
Halogens (F, Cl, Br, I)	1
Oxygen	2
Sulfur	2 or more
Nitrogen	3
Phosphorus	3 or 5
Carbon	4
Silicon	4

Exercise 2: writing formulae from names

Use the data in the table *Symbols and charges of common elements and ions* to write the formulae of the following. Before you start this exercise, make sure you have read *Section 3: naming of compounds*.

1 Sodium chloride

2 Sodium hydroxide

3 Sodium carbonate

4 Sodium sulfate

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 sulfate

19 Calcium carbonate

20 Barium chloride

21 Barium sulfate

22 Aluminium chloride

23 Aluminium oxide

24 Aluminium hydroxide

25 Aluminium sulfate

26 Copper(II) sulfate

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 bromide

36 Silver iodide

37 Silver nitrate

38 Silver oxide

39 Lead(II) nitrate

40 Lead(II) carbonate

41 Lead(II) oxide

42 Lead(IV) oxide

43 Lead(II) chloride

44 Lead(IV) chloride

45 Lead(II) sulfide

46 Tin(II) chloride

47 Tin(IV) chloride

48 Iron(II) sulfate

49 Iron(III) chloride

50 Iron(III) hydroxide

51 Ammonium chloride

52 Ammonium carbonate

53 Ammonium hydroxide

54 Ammonium nitrate

55 Ammonium sulfate

56 Ammonium phosphate

57 Phosphorus trichloride

58 Phosphorus pentachloride

59 Phosphorus trioxide

60 Phosphorus pentoxide

61 Hydrogen phosphate (Phosphoric acid)

62 Hydrogen sulfate (Sulfuric acid)

63 Hydrogen nitrate (Nitric acid)

64 Hydrogen chloride (Hydrochloric acid)

65 Carbon tetrachloride

66 Silicon tetrachloride

67 Silicon dioxide

68 Sulfur dioxide

69 Sulfur trioxide

70 Hydrogen sulfide

71 Chlorine(I) oxide

72 Nitrogen dioxide

73 Nitrogen monoxide

74 Carbon dioxide

75 Carbon monoxide

76 Hydrogen hydroxide

Section 3: Naming of Compounds

At A-level you will meet many compounds that are new to you and a lot of these will be organic compounds. In this section you will look at the naming of compounds you may already have met at GCSE. 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 of an inorganic compound must show which elements are present and, where confusion is possible, the oxidation state (or charge) of the elements concerned.

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

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** which 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**

- 2 If the elements concerned have more than one oxidation state (or charge) this may need to be shown. For example as iron can have charge +2 or +3, the name **iron chloride** would not tell you which of the two possible compounds **FeCl_2** or **FeCl_3** is being considered. In this case the oxidation state (or charge) of the iron is indicated by the use of a roman II or III in brackets after the name of the metal. In this case **iron(II) chloride** for **FeCl_2** or **iron(III) chloride** for **FeCl_3** . Other examples are:

PbCl_2 is **lead(II) chloride**
 PbCl_4 is **lead(IV) chloride**
 Fe(OH)_2 is **iron(II) hydroxide**
 Mn(OH)_2 is **manganese(II) hydroxide**

3 For compounds containing two non-metal atoms the actual number of atoms of the element present are stated, eg:

CO is **carbon monoxide** where mon- means one

CO₂ is **carbon dioxide** where di- means two

SO₂ is **sulfur dioxide**. This could be called **sulfur(IV) oxide**

SO₃ is **sulfur trioxide**. This could be called **sulfur(VI) oxide**

PCl₃ is **phosphorus trichloride**. This could be called **phosphorus(III) chloride**

PCl₅ is **phosphorus pentachloride**. This could be called **phosphorus(V) chloride**

CCl₄ is **carbon tetrachloride**

SiCl₄ is **silicon tetrachloride**

4 Where a compound contains a **metal**, a **non-metal** and **oxygen** it has a name ending in **-ate** or **-ite**. You need to remember the names and formulae of the groups listed in the table *Symbols and charges of common elements and ions*. To cover the ideas we will look at the following groups:

Carbonate -CO₃

Sulfate -SO₄

Nitrate -NO₃

A compound of sodium, carbon and oxygen would be Na₂CO₃ and would be called **sodium carbonate**. For example:

NaNO₃ is **sodium nitrate**

Mg(NO₃)₂ is **magnesium nitrate**

Fe₂(SO₄)₃ is **iron(III) sulfate**

FeSO₄ is **iron(II) sulfate**

5 As most **non-metals** can have more than one oxidation state (or charge). For example sulfur can form **sulfates** and **sulfites**. 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 oxygen atoms. **Sulfate** can also be referred to as **sulfate(VI)** and **sulfite** can also be referred to as **sulfate(IV)**. In the case of nitrogen with oxygen the compounds would be nitrate and **nitrite** or **nitrate(V)** and **nitrate(III)**.

Other elements can form compounds involving oxygen in this way. These include **chlorate(V)**, **chromate(VI)**, **manganate(VII)** and **phosphate(V)**. For example:

KNO₂ is **potassium nitrite** or **potassium nitrate(III)**

Na₂SO₃ is **sodium sulfite** or **sodium sulfate(IV)**

K₂CrO₄ is **potassium chromate(VI)**

KMnO₄ is **potassium manganate(VII)**

KClO₃ is **potassium chlorate(V)**

In summary

Common name	Systematic name	Formulae
Sulfate	Sulfate(VI)	-SO ₄
Sulfite	Sulfate(IV)	-SO ₃
Nitrate	Nitrate(V)	-NO ₃
Nitrite	Nitrate(III)	-NO ₂
Chlorate	Chlorate(V)	-ClO ₃
Hypochlorite	Chlorate(I)	-ClO

Great care needs to be taken when using these systematic names, because the properties of the two groups of compounds will be very different. In some cases use of the wrong compound in a reaction can 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.

- 6 When a compound is being considered it is usual to write the metal down first, both in the name and the formula. The exceptions to this 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**. The **ammonium** group has the formula **NH₄⁺** and sits in the place generally taken by a metal in a formula.

NH₄Cl is **ammonium chloride**

(NH₄)₂SO₄ is **ammonium sulfate**

NH₄ClO₃ 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 H₂O

sulfuric acid which is H₂SO₄

nitric acid which is HNO₃

hydrochloric acid which is HCl

ammonia which is NH₃

methane which is CH₄

9 Organic compounds have their own set of naming and 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 ₂ H ₆	is	ethane
C ₃ H ₈	is	propane
C ₄ H ₁₀	is	butane

After butane the names are based on the prefix for the number of carbons: C₅-**pent**, C₆ - **hex** and so on.

Organic compounds with 2 carbons will either start with **eth-** or have **-eth-** in their name eg

C ₂ H ₄	is	<u>eth</u>ene
C ₂ H ₅ OH	is	<u>eth</u>anol
CH ₃ COOH	is	<u>eth</u>anoic acid
C ₂ H ₅ Cl	is	chloro<u>eth</u>ane

Exercise 3: names from formulae

Use the notes in this section, the data in the table *Symbols and charges of common elements and ions* and the periodic table to write the names of the following formulae. Before you start this exercise make sure you have read *Section 2: Chemical formulae*.

1 H₂O

2 CO₂

3 NH₃

4 O₂

5 H₂

6 SO₂

7 SO₃

8 HCl

9 CH₄

10 H₂S

11 H₂SO₄

12 HNO₃ (aq)

13 NaCl

14 NaNO₃

15 Na₂CO₃

16 NaOH

17 Na₂SO₄

18 CaCl₂

19 Ca(NO₃)₂

20 CaSO₄

21 BaCl₂

22 AlCl₃

23 Al(NO₃)₃

24 Al₂(SO₄)₃

25 FeSO₄

26 FeCl₂

27 FeCl₃

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

29 PbO

30 PbO_2

31 $\text{Pb}(\text{NO}_3)_2$

32 AgNO_3

33 NH_4Cl

34 $(\text{NH}_4)_2\text{SO}_4$

35 NH_4VO_3 (V is Vanadium)

36 KClO_3

37 KIO_3

38 C_2H_6

39 C_4H_{10}

40 C_8H_{18}

41 $(\text{NH}_4)_2\text{CO}_3$

42 KMnO_4

43 K_2CrO_4

44 KHCO_3

Section 4: The Mole

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

- 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 atmospheric pressure contain the same number of moles or molecules. If the volume is 24 dm^3 , at room temperature and pressure, this number, 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 using moles

These calculations form the basis of many of the calculations you will meet in your A-level course.

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 and in this case it is 16 g mol^{-1}

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

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

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 $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} =$

$$[63.5 + 32.1 + (4 \times 16) + 5\{(2 \times 1) + 16\}] = 249.6 \text{ g mol}^{-1}$$

$$\begin{aligned} \therefore \text{number of moles of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O} &= \frac{100 \text{ g of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O}}{\text{molecular mass of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O of } 249.5 \text{ g mol}^{-1}} \\ &= \mathbf{0.4006 \text{ moles of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O molecules}} \end{aligned}$$

Example 2

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

$\text{The mass of a given number of moles} = \text{the mass of 1 mole} \times \text{the number of moles of material concerned}$
--

a Calculate the mass of 3 moles of sulfur dioxide SO_2

$$1 \text{ mole of sulfur dioxide has a mass} = 32.1 + (2 \times 16) = 64.1 \text{ g mol}^{-1}$$

$$\therefore 3 \text{ moles of } \text{SO}_2 = 3 \times 64.1 = \mathbf{192.3 \text{ g}}$$

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

$$1 \text{ mole of } \text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O} = [(23 \times 2) + (32.1 \times 2) + (16 \times 3)] + 5[(2 \times 1) + 16] = 248.2 \text{ g mol}^{-1}$$

$$\therefore 0.05 \text{ moles of } \text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O} = 0.05 \times 248.2 = \mathbf{12.41 \text{ g}}$$

Example 3

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

You will be given the information that **1 mole of any gas has a volume of 24 dm³ (24,000 cm³) at room temperature and pressure.**

$$\therefore \frac{\text{the volume of a given number of moles of gas}}{\text{moles of gas}} = \text{number of moles} \times 24\,000 \text{ cm}^3$$

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.

$$\therefore 2 \text{ moles of carbon dioxide} = 2 \times 24\,000 \text{ cm}^3 = 48\,000 \text{ cm}^3 = \mathbf{48 \text{ dm}^3}$$

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

$$\text{Volume of 0.0056 moles of chlorine} = 0.0056 \times 24\,000 \text{ cm}^3 = \mathbf{134.4 \text{ cm}^3}$$

Example 4

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

$$\text{number of moles of gas} = \frac{\text{volume of gas in cm}^3}{24\,000 \text{ cm}^3}$$

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

$$240 \text{ cm}^3$$

$$\text{Number of moles} = \frac{\quad}{24\,000 \text{ cm}^3} = 0.010 \text{ moles}$$

$$24\,000 \text{ cm}^3$$

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

$$1000 \text{ cm}^3$$

$$\text{Number of moles of gas} = \frac{\quad}{24\,000 \text{ cm}^3} = 0.0147 \text{ moles}$$

$$24\,000 \text{ cm}^3$$

Example 5

Calculation of the volume of a given mass of gas

For this calculation you need 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

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

$$\therefore 5 \text{ moles of hydrogen} = 5 \times 24\,000 \text{ cm}^3 = 120\,000 \text{ cm}^3 = \mathbf{120 \text{ dm}^3}$$

Example 6

Calculation of the mass of a given volume of gas

For this calculation you need 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

$$\begin{aligned} \therefore \text{number of moles of CO}_2 &= \frac{1000 \text{ cm}^3 \text{ of CO}_2 \text{ volume of 1 mole of CO}_2}{\text{of } 24\,000 \text{ cm}^3} \\ &= 0.0147 \text{ moles} \end{aligned}$$

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

Example 7

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

For calculations of this type you need to find the mass of 1 mole of the gas ie 24 000 cm³. This is the molar mass of the gas. For example, 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.

$$\begin{aligned} \therefore 24\,000 \text{ cm}^3 \text{ of the gas must have a mass of} &= \frac{0.0667 \text{ g} \times 24\,000 \text{ cm}^3}{100 \text{ cm}^3} \\ &= 16 \text{ g} \end{aligned}$$

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

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 0.641 g of SO₂

7 80.1 g of SO₃

8 18.20 g of HBr

9 0.0981 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 17.77 g of Na₂SO₄

16 3.16 g of KMnO₄

17 32.36 g of K_2CrO_4

18 100.1 g of KHCO_3

19 7.63 g of KI

20 3.90 g of CsNO_3

21 0.1111 g of CaCl_2

22 41.025 g of $\text{Ca}(\text{NO}_3)_2$

23 1.482 g of $\text{Ca}(\text{OH})_2$

24 3.405 mg of CaSO_4

25 41.66 kg of BaCl_2

26 14.96 μg of CuSO_4

27 13.64 g of ZnCl_2

28 1.434 mg of AgNO_3

29 13.76 kg of NH_4Cl

30 13.77 g of $(\text{NH}_4)_2\text{SO}_4$

31 23.4 g of NH_4VO_3

Exercise 4b: calculation of the mass of material in a given number of moles of that material

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

1	2 moles of H ₂ O
2	3 moles of CO ₂
3	2.8 moles of NH ₃
4	0.50 moles of C ₂ H ₅ OH
5	1.2 moles of C ₂ H ₄
6	0.64 moles of SO ₂
7	3 moles of SO ₃
8	1 mole of HBr
9	0.012 moles of H ₂ SO ₄
10	0.15 moles of HNO ₃
11	0.45 moles of NaCl
12	0.70 moles of NaNO ₃
13	0.90 moles of Na ₂ SO ₄

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 cm³ under these conditions.)

1	1 mole of CO ₂
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2	0.1 moles of NH ₃
---	------------------------------

3	0.5 moles of C ₂ H 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 ₂ H ₆
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12	0.0090 moles of C ₃ H ₈
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13	0.040 moles of C ₂ H ₂
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14	0.123 moles of NO
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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 cm³ under these conditions.)

1	200 cm ³ of CO ₂
2	500 cm ³ of NH ₃
3	1000 cm ³ of C ₂ H 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

Exercise 4e: 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

Exercise 4f: 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 cm ³ of CO ₂
2	500 cm ³ of NH ₃
3	1000 cm ³ of C ₂ H 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 ₂ H ₆
12	5600 cm ³ of C ₃ H ₈
13	2200 cm ³ of C ₂ H ₂

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

Section 5: calculations involving chemicals in solution

Experiments measuring concentrations of chemicals in solution 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 solutions of known concentration.

Concentration in solution is generally measured as moles per 1000 cm³ **of solution**. For example, sodium chloride may be labelled as 1M NaCl. This means that each 1000 cm³ of the solution contains 1 mole of NaCl (58.5 g), or its concentration is 1 mol dm⁻³.

It does not mean that 58.5 g of NaCl have been added to 1000 cm³ of water as the volume of the mixture may no longer be 1000 cm³.

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 more water to make the total volume of the mixture up to 1000 cm³. (1 dm³)

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

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

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

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

In reactions in solution it is often more convenient to use molarity (number of mol dm⁻³) rather than g dm⁻³.

Exercise 5a: calculations based on concentrations in solution

Calculate the number of moles of the underlined species in the given volume of solution

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

Exercise 5b: calculate the mass of material in the given volume of solution

1 25 cm³ of 1 mol dm⁻³ HCl

2 50 cm³ of 0.5 mol dm⁻³ NaCl

3 100 cm³ of 0.25 mol dm⁻³ NH₄NO₃

4 100 cm³ of 0.1 mol dm⁻³ AgNO₃

5 25 cm³ of 1 mol dm⁻³ BaCl₂

6 50 cm³ of 0.2 mol dm⁻³ H₂SO₄

7 20 cm³ of 0.1 mol dm⁻³ NaOH

8 50 cm³ of 0.1 mol dm⁻³ K₂CrO₄

9 25 cm³ of 0.02 mol dm⁻³ KMnO₄

10 25 cm³ of 0.1 mol dm⁻³ Pb(NO₃)₂

Exercise 5c: what is the concentration in moles dm^{-3} of the following?

1 3.65 g of HCl in 1000 cm^3 of solution

2 3.65 g of HCl in 100 cm^3 of solution

3 6.624 g of $\text{Pb}(\text{NO}_3)_2$ in 250 cm^3 of solution

4 1.00 g of NaOH in 250 cm^3 of solution

5 1.962 g of H_2SO_4 in 250 cm^3 of solution

6 1.58 g of KMnO_4 in 250 cm^3 of solution

7 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

8 25.0 g of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in 250 cm^3 of solution

9 4.80 g of $(\text{COOH})_2 \cdot 2\text{H}_2\text{O}$ in 250 cm^3 of solution

10 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

11 240 cm^3 of $\text{NH}_3(\text{g})$ dissolved in 1000 cm^3 of solution

12 480 cm^3 of $\text{HCl}(\text{g})$ dissolved in 100 cm^3 of solution

Section 6: Chemical Equations

Chemical equations do much more than tell us what reacts with what in a chemical reaction. They tell us how many of each type of molecule are needed and produced, so they also tell us 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:

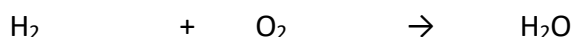


Writing formulae

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

- hydrogen is represented as H₂
- oxygen is represented as O₂
- water is H₂O

So we get:

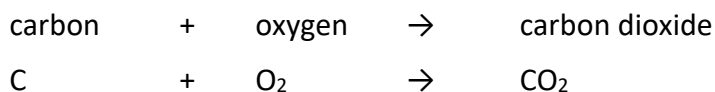


However, this will not suffice as a full equation as you will discover if you read on!

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 a chemical equation you must have the same number of each kind of atoms 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:



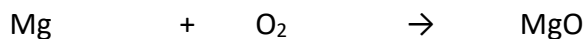
Carbon dioxide has one atom of carbon and two atoms of oxygen in one molecule. Carbon is written as C (one atom) and oxygen molecules have two atoms each, written as O₂.

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) – therefore it is already balanced.

Now let us try one that does not work out.



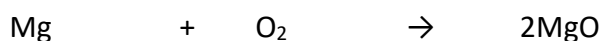
Magnesium is written as Mg (one atom just like carbon) and oxygen is O₂, but magnesium oxide has just one atom of oxygen per molecule and is therefore written as MgO.



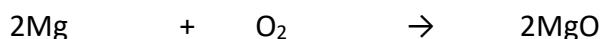
The magnesium balances, one atom on the left and one on the right, but the oxygen does not as there are two atoms on the left-hand side of the equation and only one on the right-hand side.

You cannot change the formulae of the reactants or products.

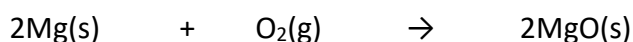
Each 'formula' of magnesium oxide has only one atom of oxygen but each molecule of oxygen has two atoms of oxygen, so you can make *two* formulae of magnesium oxide for each molecule of oxygen. So we get:



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 adding 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 in the first six questions where numbers need to be inserted to achieve the balance. In one or two difficult cases some of the numbers have been added. You will not need to change these. Remember all the formulae are correct!

