Chemistry A level

Course plan

This plan shows the structure of the course and gives an outline of the contents. Sections 1–5 cover the requirements of the AS and Part 1 of the A level; Sections 6–10 cover Part 2 of the A level. You need to do Sections 1–10 to prepare for the A level.

Getting Started

Introduction

Making the most of the course

A level Chemistry Course guide

Section 1 Working as a chemist

Data sheet

Chapter A1 (textbook)

Maths help 1

Topic 1 Moles

Prior knowledge (textbook)

Chapter 1 Atomic structure and the periodic table (textbook)

Core practical 2 Preparing a standard solution

Core practical 1 Measuring the molar volume of a gas

Topic 2 Monitoring the course of a reaction

Section 1 Key terms

Assignment 1

Section 2 Bonding and redox reactions

Maths help 2

Topic 1 Organic chemistry

Chapter 2 Bonding and structure (textbook)

Section 2 Key terms

Assignment 2

Chapter 3 Redox 1 (textbook)

Topic 2 Redox reactions

Section 2 Key terms

Assignment 2

Section 3 Inorganic chemistry, quantities and energy in reactions

Maths help 3

Topic 1 Electronic structures

Chapter 4 Inorganic chemistry and the periodic table (textbook)

Core practical 3 Determination of the concentration of hydrochloric acid solution

Chapter 5 Formulae, equations and amounts of substance (textbook)

Core practical 7A Identification of inorganic compounds

Chapter 8 Energetics 1 (textbook)

Core practical 8 Measuring enthalpy change using Hess's law

Section 3 Key terms

Assignment 3

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Section 4 Organic chemistry

Maths help 4

Topic 1 Naming organic compounds

Chapter 6.1 Introduction to organic chemistry (textbook)

Chapter 6.2 Hydrocarbons, alkanes and alkenes

Chapter 6.3 Halogenoalkanes and alcohols (textbook)

Core Practical 4 Rates of hydrolysis of halogenoalkanes

Core Practical 5 The oxidisation of ethanol

Core Practical 6 Synthesis of 2-chloro-2-methylbutane

Section 4 Key terms

Assignment 4

Section 5 Modern analytical techniques, kinetics and equilibrium

Equations for waves and the particle behaviour of light

Maths help 5

Chapter 7 Modern analytical techniques 1 (textbook)

Core practical 7B Identification of organic compounds

Chapter 9 Kinetics 1 (textbook)

Topic 1 A case study in kinetics

Chapter 10 Equilibrium 1

Topic 2 A case study in equilibrium

Section 5 Key terms

The data booklet

Assignment 5

Section 6 Equilibria and lattice energy

Maths help 6

Topic 1 Safety in the laboratory

Chapter 11 Equilibrium 2 (textbook)

Topic 2 Choosing a suitable indicator for acid-base titrations

Chapter 12 Acid-base equilibria (textbook)

Core Practical 9 Finding the K_a value for a weak acid

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Acid-base titration experiment video

Chapter 13.1 Lattice energy (textbook)

Section 6 Key terms

Assignment 6

Section 7 Entropy, further redox and transition metal chemistry

Maths help 7

Chapter 13.2 Entropy (textbook)

Topic 1 Entropy changes

Chapter 14 Redox II (textbook)

Core Practical 10 Electrochemical cells

Core Practical 11 Redox titration

Chapter 15 Transition metals (textbook)

Core Practical 12 Preparing a transition metal complex

Core Practical 15a Identification of inorganic unknowns

Section 7 Key terms

Electrochemical cells experiment video

Assignment 7

Section 8 Kinetics, chirality and carbonyl compounds

Section 8 Key terms

Maths help 8

Chapter 16 Kinetics II (textbook)

Core Practical 13A Rate of reaction measured using a 'clock' reaction

Core Practical 13B Reaction rate by monitoring evolution of gas

Core Practical 14 Measuring activation energy

Chapter 17.1 Chirality (textbook)

Chapter 17.2 Carbonyl compounds (textbook)

Assignment 8

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Section 9 Carboxylic acids, arenes and amines

Section 9 Key terms

Chapter 17.3 Carboxylic acids and their derivatives (textbook)

Chapter 18.1 Arenes - benzene compounds (textbook)

Chapter 18.2 Arenes, amides, amino acids and proteins (textbook)

Core Practical 15 Planning chemical analysis

Assignment 9

Section 10 Organic synthesis and modern analytical techniques

Section 10 Key terms

Chapter 18.3 Organic synthesis (textbook)

Core Practical 16 Synthesis of aspirin

Chapter 19 Modern analytical techniques II (textbook)

Data booklet

Assignment 10

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Sample of the Chemistry A Level course from Section 1

Maths help 1

Introduction

Mathematics is a tool which will help you to solve chemical problems. Throughout the course you will get hints and tips. It is assumed you have skills approximately at GCSE level Maths at the start, and most of the additional skills you need will be taught at the point where you need them.

You will probably need 2 hours to study this resource.

Objectives

When you have completed this resource you should be able to:

- recall the common SI units used in chemistry
- recall the adaptations that chemists have made in order to handle typical volumes used in the laboratory
- calculate percentage error
- measure the gradient of a graph
- use logarithmic scales in cases where data span a wide range

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SI units

SI stands for '*Système international*' which was an internationally agreed set of units adopted in the early 1970s. This system has only seven base units, from which others are derived. These seven are:

Table 1 The SI base units

SI unit	Unit symbol	Quantity name
metre	m	length
kilogram	kg	mass
second	s	time
ampere	a	electric current
kelvin	K	temperature
mole	mol	quantity of a substance
candela	cd	luminous intensity

Some common derived units which are used in chemistry are:

Table 2 Some derived units

Derived unit	Unit symbol	Used for
newton	N	force, weight
pascal	Pa	pressure, stress, force per unit area
joule	J	energy, work
watt	W	power
volt	V	potential difference, joule per coulomb
degree Celsius	°C	temperature, relative to 273.15 K

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The preferred multiples of these units are in thousands; for example, the units of measurement that are used are kg, g, mg, where 1000 mg = 1 g and 1000 g = 1 kg.

Non-SI units

Several non-SI units are still in common use. These are especially important where the SI units are not very user-friendly. The measurement of volume for laboratory work is a particular example.

The preferred SI units of length are metre and millimetre. A cubic metre looks like this:

Figure 1 A cubic metre

whereas a cubic millimetre looks like this:

Figure 2 A cubic millimetre

Neither of these is very appropriate as a measure for laboratory experiments. Chemists have therefore departed from the SI system slightly in their measures of volume. They have retained the

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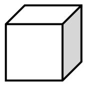
centimetre, cm, so continue to measure volumes in cubic centimetres, cm^3.

This is much more practical, because a cm^3 looks like this:

Figure 3 A cubic centimetre



Activity 1

(Allow 10 minutes)

There are 10 millimetres in a centimetre. How many cubic millimetres are there in a cubic centimetre?

- There are $10 \times 10 \times 10 = 1000$ cubic millimetres in a cubic centimetre.

If you have difficulty in imagining this, look at it this way:

- If you had a lot of millimetre cubes, you would need to lay out ten in a row to make a line 1 cm long, by 1 mm wide
- If you put 10 of these rows side by side you would have a 1-centimetre square. You would now have used 100 millimetre cubes
- You would need 10 of these layers on top of one another to form a cubic centimetre, which comes to $10 \times 100 = 1000$ cubic millimetres.

Typical everyday amounts are:

- a cup of tea – about 200 cm^3
- a bottle of wine – about 750 cm^3
- a teaspoon – about 5 cm^3.

When you reach a volume of 1000 cm^3 it is called 1 dm^3, which is '1 cubic decimetre'. This is not an SI unit or a derived unit.

- There are 1000 mm^3 in 1 cm^3
- There are 1000 cm^3 in 1 dm^3
- There are 1000 dm^3 in 1 m^3.

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Everyday units

It becomes even more complex when you compare these units to those in everyday use. The litre was the preferred unit of volume until the adoption of the SI system. The litre is the same size as a cubic decimetre. One millilitre, ml, is equivalent to 1 cm^3.

You will find that many pieces of laboratory equipment are marked in litres or ml.

For calculations for A level, the use of the dm^3 is preferred.

Uncertainty

When a chemist does an experiment, it is important to assess how precise the measurements have been, and the sources of error. Sometimes the precision is determined by the equipment you are using – for instance you cannot measure the thickness of a single human hair using a ruler. Sometimes it is determined by your own technique; this is particularly the case when assessing concentrations of solutions, and you will learn about this during the course.

If you have a ruler which has 1 mm markings, it is usual to express its precision as ± 0.5 mm. Therefore, for instance, a measurement of 50 mm could be anything from 49.5 mm to 50.5 mm. You would say the percentage error was:

$$\frac{0.5}{50.0} \times 100 = 1\%.$$

Gradients of graphs

We need to be able to find the gradients of graphs in order to work out the rates at which chemical reactions are taking place.

You are probably already familiar with the concept of finding the gradient of a straight-line graph:

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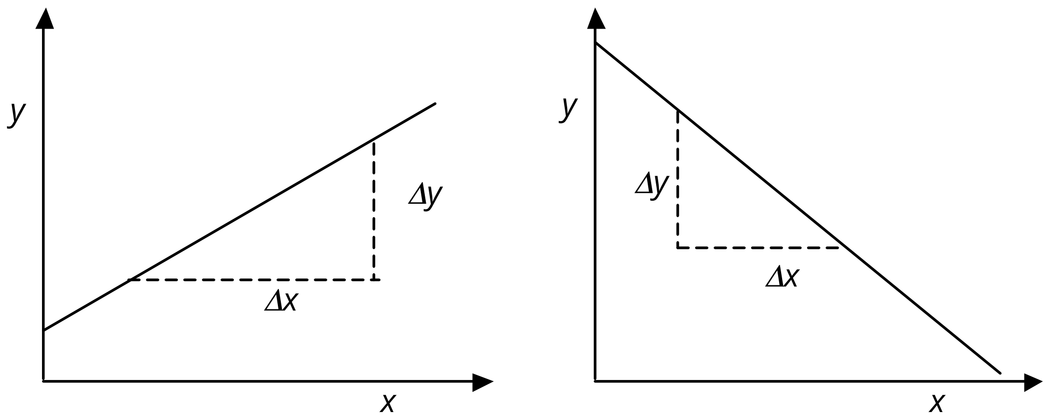
Figure 4 Two straight line graphs



In both of these examples, the gradient is simply $\frac{\text{the change in } y}{\text{the change in } x}$

or $\frac{\Delta y}{\Delta x}$. The symbol Δ is an upper-case Greek 'delta'.

In the first example in Figure 4 the gradient will be a positive number and in the second it will be negative.

The gradient of these graphs is equivalent to the **rate of change of *y* with respect to *x*.**

Activity 2

(Allow 20 minutes)

Find the gradients of the graphs in Figures 5 and 6.

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Figure 5

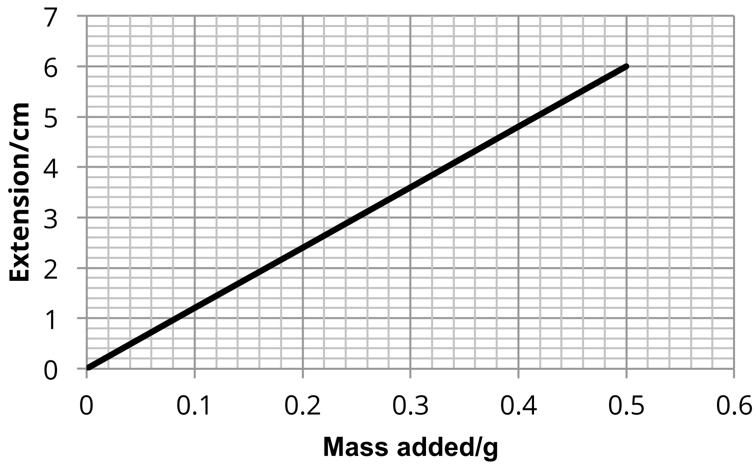
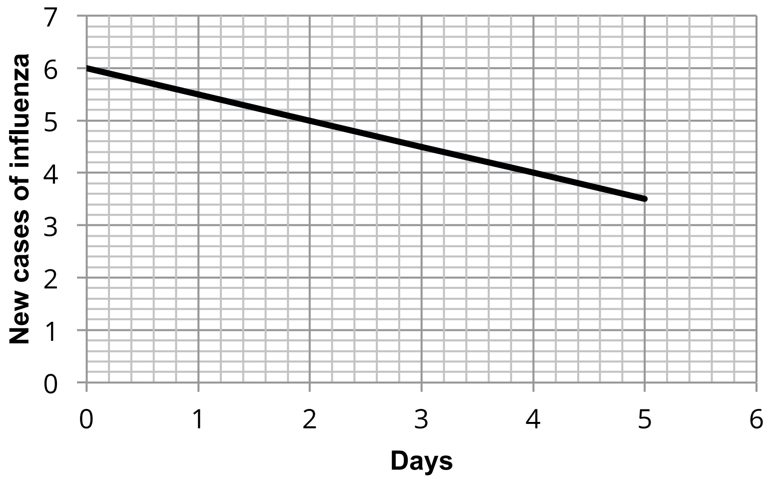


Figure 6



- The slope of graph in Figure 5 is:

$$\frac{\Delta y}{\Delta x} = \frac{6.0\text{ cm} - 0.0\text{ cm}}{0.5\text{ g} - 0.0\text{ g}} = 12\text{ cm g}^{-1}.$$

- The slope of graph in Figure 6 is:

$$\frac{\Delta y}{\Delta x} = \frac{4.0 - 6.0\text{ cases}}{4.0 - 0.0\text{ days}} = -0.5\text{ cases day}^{-1}.$$

Note the importance of including the units in your answer.

Gradients of curved graphs

Finding a gradient becomes more complex when the graph is a curve. The gradient has to be found in a slightly different way. The gradient is now different at every point on the curve, so you need

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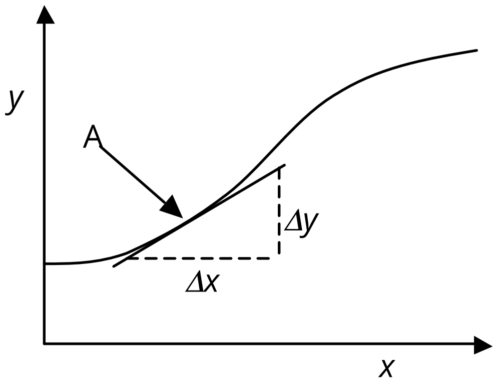
to define the place where you require the gradient. For the graph in Figure 7, this is the point marked A.

Figure 7 Finding the gradient of a curved line



The gradient at the point A marked with the arrow has the value of the slope of the line that has been drawn at a tangent to the curve.

Again it is calculated from $\frac{\Delta y}{\Delta x}$ in relation to this line.

When you are given a curve and asked to measure the gradient at a particular point, this is how you can do it. You lay your ruler next to the curve and draw a line that just touches the curve. Then measure the gradient of the straight line that you have drawn.

Activity 3

(Allow 10 minutes)

Find the gradient of the graph in Figure 8 at the point marked with the arrow, where x = 4.0.

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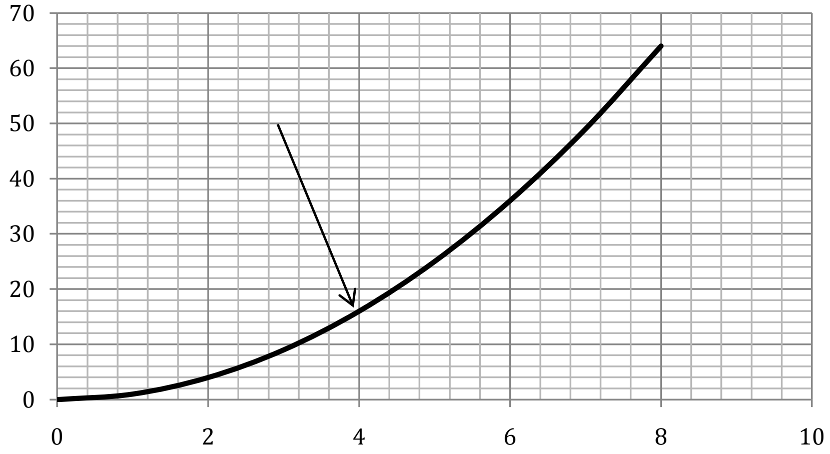
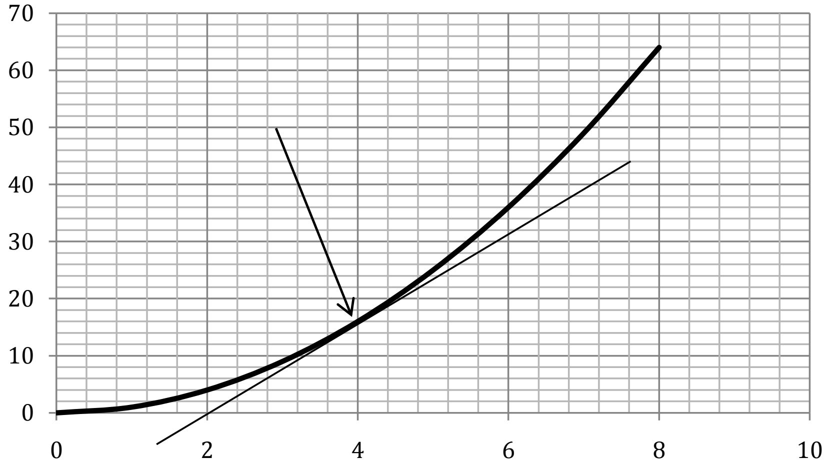
Figure 8



- We drew a tangent as shown in Figure 9. The gradient approximates to:

$$\frac{4.0 - 0.0}{6.6 - 2.05} = 8.8.$$

Figure 9 Construction for calculating the gradient of the curve



Difficult graph scales

Sometimes an experiment will give us data that span a very large range, but not evenly, and this can be difficult to display effectively on a graph.

To illustrate this, consider the data in the following table. They come from measurements of the cooling rates of solid bars, as a function of the bar diameter.

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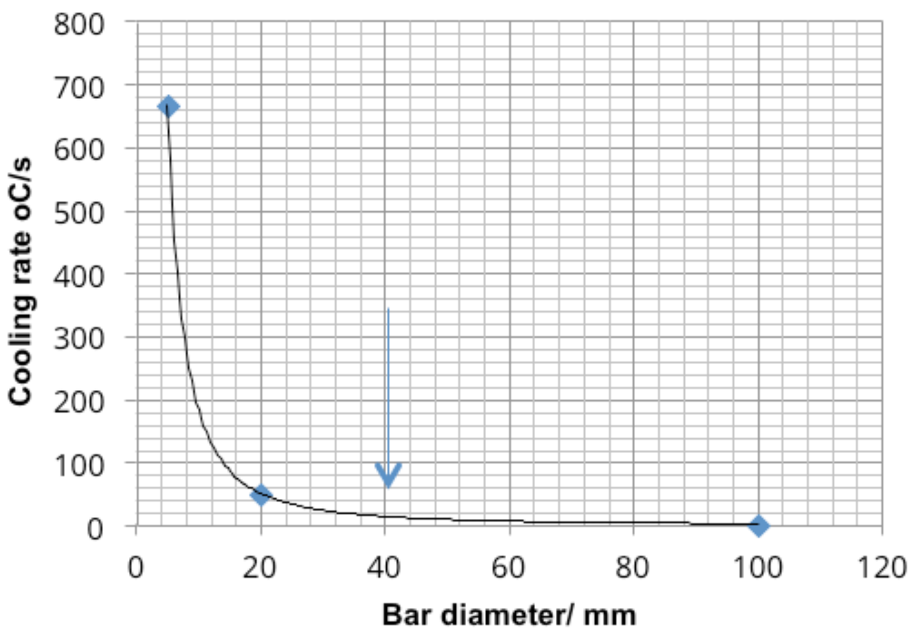
Table 3 Cooling data for solid bars

Bar diameter/mm	100	20	5
Cooling rate/°C s^{-1}	2.5	5	667

When we plot this data in the conventional way we get the graph in Figure 10.

Figure 10 Plot of data from Table 3



Suppose that you wish to use the data to predict the cooling rate for a 40-mm bar. This point is marked with an arrow on the graph. Due to the very odd shape of this graph, it is difficult to read this point with any accuracy. Luckily there is a way that this can be improved.

Logarithms

There are two common types of logarithm, those to the base ten and those to the base e (which are also known as natural logarithms). If you look at your calculator you will see two buttons – one marked 'log' and another marked 'ln'. The 'log' button calculates logs to the base ten, and the 'ln' button calculates logs to the base e.

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Activity 4

(Allow 10 minutes)

For each number below in turn, press the 'log' button, then the number given, and press '=':

10, 100, 1000.

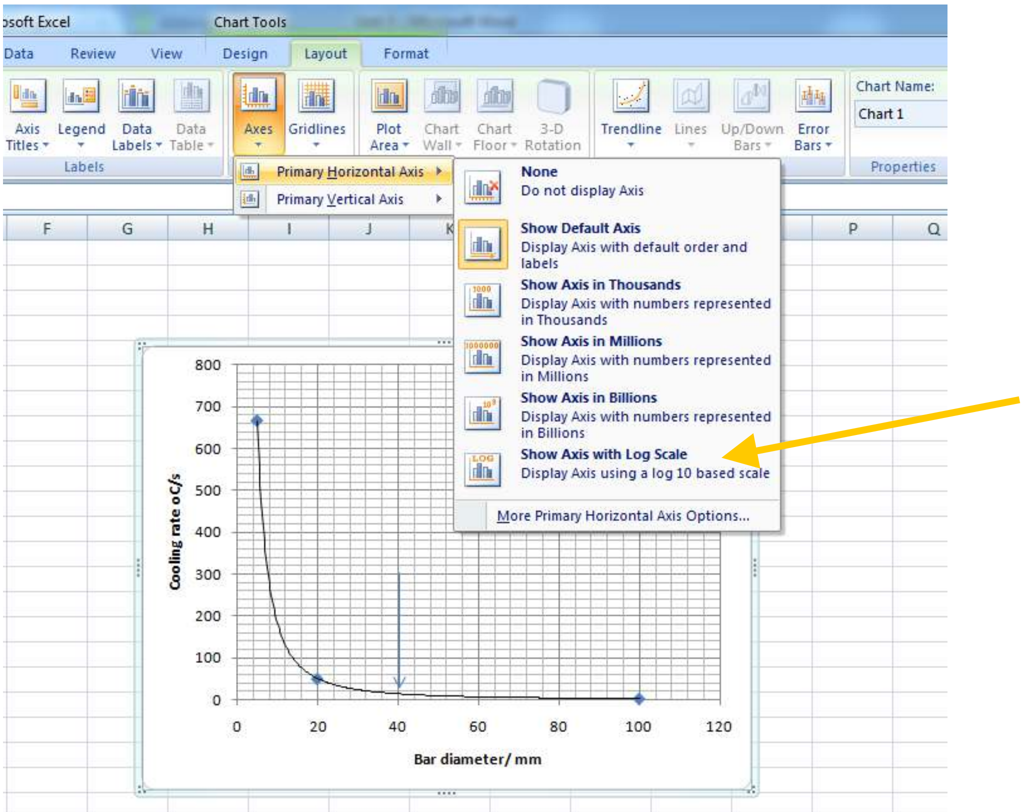
- Our answers were 1, 2, 3.

You will notice that when using the log function the sequence 10, 100, 1000 has logs that are 1, 2 and 3. This means that you can plot a really large range of numbers if you take logs.

Plotting a graph with logarithmic scales

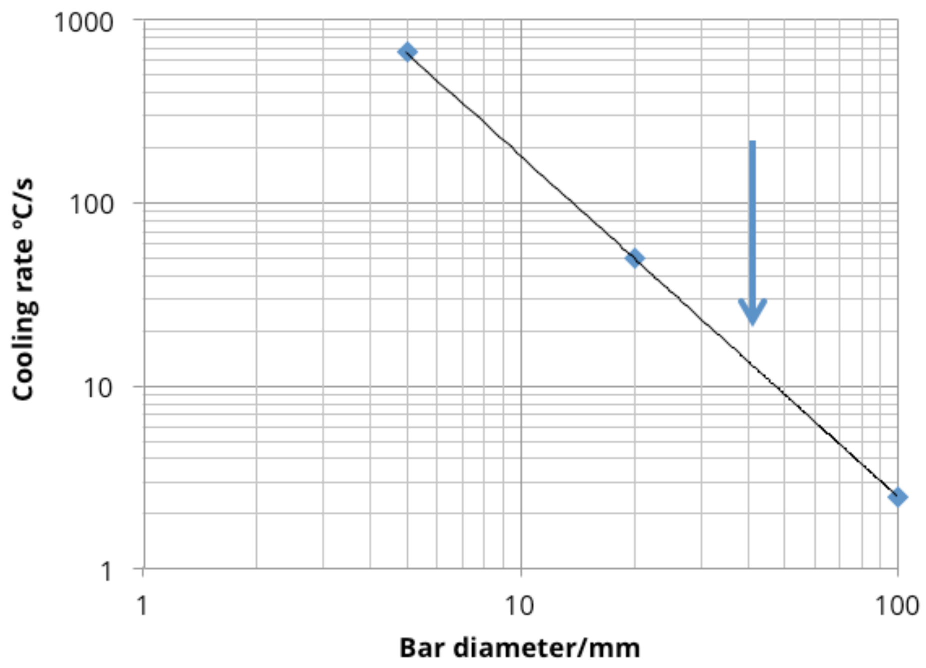
The package Excel and others that are similar offer an option for using logarithmic scales on the axes. Look at the screen shot in Figure 11.

Figure 11 Excel axes options



By applying logarithmic scales to both horizontal and vertical axes for the data on cooling the bar we obtained this graph:

Figure 12 Data plotted on logarithmic scales on both axes



This is much better and we can now read off the cooling rate for a bar 40 mm diameter – it is approximately 12 °C s^{-1}. Notice that the divisions on the horizontal axis are now: 1, 2, 3, 4, 5, 6, 7, 8, 9, 10, 20, 30, 40, 50, 60, 70, 80, 90, 100.

As a consequence, reading off values from a logarithmic scale takes a little practice.

Summary

You are now aware of SI units and the preferred units for volume in laboratory chemistry. You can calculate a percentage error, a slope of either a straight line or a curve, and know how to use logarithms to deal with data which span a very large range.

References

Any figures for which there is no reference are the author's own work.

Figure 1: By KVDP (Own work) [Public domain], via Wikimedia Commons.

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Topic 1

Moles

Introduction

The concept of the mole is central to chemical calculations. You may have already met it in your GCSE studies, and this topic will remind you what it is and give you further practice in using the concept.

You will probably need 1 hour to complete this topic.

Objectives

When you have completed this topic you should be able to:

- use the concept of the mole in describing chemical reactions
- calculate molar masses of substances`

The meaning of chemical equations

The reaction of hydrogen gas and oxygen gas results in the formation of water. The equation for this is:

$2\ H_2 + O_2 = 2\ H_2O$

This chemical equation is balanced because there are the same number of atoms of each kind on the left-hand side as there are on the right-hand side. This balancing is always a prerequisite of any chemical calculation.

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This balanced equation can be interpreted as '2 molecules of hydrogen gas react with 1 molecule of oxygen gas to form 2 molecules of water'

You can scale this up, and it would be equally true to say '100 molecules of hydrogen gas react with 50 molecules of oxygen gas to form 100 molecules of water'.

In other words, provided you keep the ratios the same, you can scale it up to any number you like.

The mole

The **mole** is one of the base units for the International System of Units, and its abbreviation is **mol**. Please don't confuse 'mol' with a possible abbreviation of 'molecule', because they are totally different.

The definition of the mole is:

> 'The mole is a term that defines the quantity of a substance that has the same number of elementary entities (such as atoms, molecules, ions or electrons) as 12 g of pure 12C.'

This is a bit of a mouthful – the important thing to remember is that a mole is a number (like 100 in the example above).

12 g of pure 12C contains 6.02 × 1023 atoms (this number is known as the Avogadro constant), so a mole of 12C is 6.02 × 1023 atoms.

The relative atomic mass in grams of any element also contains 6.02 × 1023 of its atoms. The relative molecular mass of a molecular compound also contains 6.02 × 1023 of the molecule. A similar statement can be made about the formula mass of an ionic compound. (The subtle differences between the terms 'molecular mass' and 'formula mass' will be defined in Chapter 1 of the textbook.)

Therefore a mole of anything is its relative mass (atomic, molecular or formula) in grams.

Activity 1

(Allow 10 minutes)

What is the mass of a mole of:

1 hydrogen molecules, H_2

1 sodium chloride, NaCl

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1 hydrogen atoms, H?

- The relative atomic mass of hydrogen is 1.0, so a mole of hydrogen molecules has a mass of 2.0 g.
- The relative atomic mass of sodium is 23.0, and of chlorine is 35.5, so a mole of NaCl has a mass of 58.5 g.
- The relative atomic mass of hydrogen is 1.0, so a mole of hydrogen atoms has a mass of 1.0 g.

The mole makes chemical calculations a great deal easier. We can use moles to describe the overall outcome of a reaction.

Thus the equation:

$2\ H_2 + O_2 = 2\ H_2O$

- can be stated as '2 moles of hydrogen molecules react with 1 mole of oxygen molecules to form 2 moles of water'.

You will notice that there are three moles of the reactants on the left-hand side of this example, and only two moles of products. This is not a problem – the atoms are conserved, but the reorganisation can cause the total number of moles to change.

Moles of electrons

You can also have a mole of electrons. One mole of electrons is 6.02×10^{23} electrons.

In the discharge reaction that converts aluminium ions into aluminium metal in the electrochemical cell, 1 mole of aluminium ions reacts with 3 moles of electrons to give 1 mole of aluminium metal:

$Al^{3+} + 3e^- = Al$.

One ion of Al^{3+} needed 3 electrons to become aluminium metal. It follows that 1 mole of Al^{3+} ions, which is 6.02×10^{23} ions, required 3 moles, or $3 \times 6.02 \times 10^{23}$ electrons.

Converting masses into moles

It is useful to convert masses of substances into moles or fractions of a mole. For example:

You have 10 g of potassium chloride, KCl. How many moles is this?

Answer: 1 mole of potassium chloride has a mass of (39.1 + 35.5) g = 74.6 g.

Hence 10 g is 10 / 74.6 = 0.134 mol.

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Activity 2

(Allow 10 minutes)

How many moles are the following?

1 5 g of HCl

1 40 g of $Mg(NO_3)_2$

1 10^{26} electrons

- The formula mass of HCl is (1 + 35.5) = 36.5 g. Hence 5 g is 5 / 36.5 = 0.137 mol.
- The formula mass of Mg(NO3)2 is (24.3 + 2 × 14.0 + (6 × 16.0) = 148.3 g, hence 40 g is 40 / 148.3 = 0.270 mol.
- There are 6.02 × 1023 electrons in a mole, so 1026 electrons is 1026 / (6.02 × 1023) = 166 mol.

Summary

You can now calculate how many moles a given mass of a given substance is, and use the term in relation to chemical reactions.

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What next?

We hope this sample has helped you to decide whether this course is right for you.

If you have any further questions, please do not hesitate to contact us using the details below.

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