Transition Pack for A Level Chemistry

Dear all,
Welcome to the chemistry transition pack to support you as you move from KS4 to KS5.

It is important to ensure you have the basic GCSE knowledge as secure as you can get it before embarking on your courses at KS5.

Step 1: Sign into amazon and download the CGP transition pack- currently available free for the Kindle edition.

Step 2: Work through the pack and make notes as you go along. This work will be checked by your teachers in September.

This information aims to provide you with:
- Information about the course
- Preparation for a successful start to you're A level Chemistry course
- Relevant information about the course

OCR A Chemistry

We follow the **OCR Chemistry A H432** course, which has the following structure:

Year 12

Physical Chemistry	**Organic chemistry and analysis**
Module 1 – Development of practical skills in chemistry • Practical skills assessed in a written examination • Practical skills assessed in the practical endorsement • You will gain practical skills throughout the course	
Module 2 – Foundations in chemistry • Atoms, compounds, molecules and equations • Amount of substance ••Acid–base and redox reactions • Electrons, bonding and structure	**Module 4 – Core organic chemistry** • Basic concepts • Hydrocarbons • Alcohols and haloalkanes • Organic synthesis • Analytical techniques (IR and MS)
Module 3 – Periodic table and energy • The periodic table and periodicity • Group 2 and the halogens • Qualitative analysis • Enthalpy changes • Reaction rates and equilibrium (qualitative)	

Year 13:

Physical Chemistry	Organic Chemistry
Module 5 – Physical chemistry and transition elements • Reaction rates and equilibrium (quantitative) • pH and buffers • Enthalpy, entropy and free energy • Redox and electrode potentials •	**Module 6 – Organic chemistry and analysis** • Aromatic compounds • Carbonyl compounds • Carboxylic acids and esters • Nitrogen compounds • Polymers • Organic synthesis
• Transition elements • Chromatography and spectroscopy (NMR)	

Assessment all external exams are sat at the end of year 13 (hopefully).

Component	Marks	Duration	Weighting	
Periodic table, elements and physical chemistry (01)	100	2 hour 15 mins	37%	Assesses content from modules 1, 2, 3 and 5
Synthesis and analytical techniques (02)	100	2 hour 15 mins	37%	Assesses content from modules 1, 2, 4 and 6
Unified chemistry (03)	70	1 hour 30 mins	26%	Assesses content from all modules (1 to 6)
Practical endorsement in chemistry (04)	-	-	Pass or Fail	Non-exam assessment- teacher marked practical tasks throughout the course.

All components include synoptic assessment.

Students must complete all components (01, 02, 03, and 04) to be awarded the OCR A Level in Chemistry A

Equipment
You will need: Calculator, pen, pencil, 30 cm ruler, paper, folders with dividers to organise your work.
For practical's you also need a lab coat and goggles your lab coat must have poppers not buttons for safety.
You will be given a practical book which must be brought to every lesson.

Homework
This will be set weekly and will need to be handed in on time. If you struggle you need to email us in advance of the deadline to get help. Failure to hand in homework and we will contact home to ensure you do not fall behind in your studies.
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Independent study
Each week you are expected to work by yourself on extra work set by staff. This does not need to be handed in for marking. Your independent study will be checked during practical sessions. If this is not up to date you will be required to attend extra sessions to catch up.

This pack contains a programme of activities and resources to prepare you to start an A level in Chemistry in September. It is aimed to be used after you complete your GCSE, throughout the remainder of the summer term and over the Summer Holidays to ensure you are ready to start your course in September

Books:

We recommend AS/A level calculations in chemistry by Jim Clark to help with physical chemistry:
https://www.amazon.co.uk/Calculations-Level-Chemistry-Jim-Clark/dp/0582411270 there are many good secondhand versions available if you do not want to buy new.

A good revision guide: There are many out on the market and any will do- you will have an opportunity to purchase one at parents evening for reduced prices.

In terms of a course textbook- The OCR one has mistakes in, so we tend to use chemguide by Jim Clark- an online website. https://www.chemguide.co.uk/

If you prefer books, then wait until we can show you some of the ones available so you can see which ones you prefer.

Other useful sites:

RSC chemistry. http://www.rsc.org/learn-chemistry/resources/problem-solving-tutor/ helps with calculations

http://www.rsc.org/learn-chemistry/resources/gridlocks/ online games to unlock the grids- start with the GCSE ones and build up from there.

This booklet may seem like a lot of work, but you will be grateful for completing it once we start the course. If you are behind in your GCSE knowledge at the start you will find it hard to catch up. The grade boundaries for A level chemistry are a lot higher than those at GCSE. Last years A* was 88% overall compared with 75% at GCSE for a grade 9. Those students that succeed are the ones that put the work in from the start and keep asking questions to aid their understanding.

After completing the booklet summarise your strengths and weaknesses and have this ready for your first lesson.

If you need to ask any questions you can contact me at:

susmith@combertonvc.org always best to ask if you need help than sit at home and struggle or even worse give up!

Kind regards,

Dr Smith

Head of Chemistry.

Videos to watch online

Rough science – the Open University – 34 episodes available
Real scientists are 'stranded' on an island and are given scientific problems to solve using only what they can find on the island. Great fun if you like to see how science is used in solving problems. There are six series in total
http://www.dailymotion.com/playlist/x2igjq_Rough-Science_rough-science-full-series/1#video=xxw6pr; https://www.youtube.com/watch?v=lUoDWAt259I

A thread of quicksilver – The Open University
A brilliant history of the most mysterious of elements – mercury. This program shows you how a single substance led to empires and war, as well as showing you come of the cooler properties of mercury.
https://www.youtube.com/watch?v=t46lvTxHHTA

10 weird and wonderful chemical reactions
10 good demonstration reactions, can you work out the chemistry of …. any… of them?
https://www.youtube.com/watch?v=0Bt6RPP2ANI

Chemistry in the Movies
Dante's Peak 1997: Volcano disaster movie.
Use the link to look at the Science of acids and how this links to the movie.
http://www.open.edu/openlearn/science-maths-technology/science/chemistry/dantes-peak
http://www.flickclip.com/flicks/dantespeak1.html
http://www.flickclip.com/flicks/dantespeak5.html

Fantastic 4 2005 &2015: Superhero movie
Michio Kaku explains the "real" science behind fantastic four
http://nerdist.com/michio-kaku-explains-the-real-science-behind-fantastic-four/
http://www.flickclip.com/flicks/fantastic4.html

Research activities

Use your online searching abilities to see if you can find out as much about the topic as you can. You should aim to push **your** knowledge.

Task 1: The chemistry of fireworks
What are the component parts of fireworks? What chemical compounds cause fireworks to explode? What chemical compounds are responsible for the colour of fireworks?

Task 2: Why is copper sulfate blue?
Copper compounds like many of the transition metal compounds have got vivid and distinctive colours – but why?

Task 3: Aspirin
What was the history of the discovery of aspirin, how do we manufacture aspirin in a modern chemical process?

Task 4: The hole in the ozone layer

Why did we get a hole in the ozone layer? What chemicals were responsible for it? Why were we producing so many of these chemicals? What is the chemistry behind the ozone destruction?

Task 5: ITO and the future of touch screen devices
ITO – indium tin oxide is the main component of touch screen in phones and tablets. The element indium is a rare element and we are rapidly running out of it. Chemists are desperately trying to find a more readily available replacement for it. What advances have chemists made in finding a replacement for it?

Pre-Knowledge Topics

Chemistry topic 1 – Electronic structure, how electrons are arranged around the nucleus

A periodic table can give you the proton / atomic number of an element, this also tells you how many electrons are in the ***atom***.

You will have used the rule of electrons shell filling, where:
The first shell holds up to 2 electrons, the second up to 8, the third up to 8 and the fourth up to 18 (or you may have been told 8).
Atomic number =3, electrons = 3, arrangement 2 in the first shell and 1 in the second or
Li = 2,1

Q1.1 Draw out the labelled atomic model for

a) Ca b) Al c) S d) Cl e) Ar f) Na

Q1.2 Write out the electron configuration of:

a) Ca b) Al c) S d) Cl e) Ar f) Na

Q1.3 Write out the electron arrangement of the following ***ions***:
a) Ca^{2+} b) Al^{3+} c) Cl^{-} d) Na^{+}

Link to physics and maths tutor and BBC bitesize

https://www.physicsandmathstutor.com/chemistry-revision/gcse-aqa/atomic-structure-and-periodic-table

https://www.bbc.com/bitesize/guides/z3sg2nb/revision/1

At **A level** you will learn that the electron structure is more complex than this, but you need to be good at the basics first.

Chemistry topic 2 – Oxidation and reduction

At GCSE you know that oxidation is adding oxygen to an atom or molecule and that reduction is removing oxygen, or that oxidation is removing hydrogen and reduction is adding hydrogen. You may have also learned that oxidation is removing electrons and reduction is adding electrons (OILRIG).

At A level we use the idea of ***oxidation number*** a lot!
You know that the metals in group 1 react to form ions that are +1, i.e. Na+ and that group 7, the halogens, form -1 ions, i.e. Br -.

We say that sodium, when it has reacted has an oxidation number of +1 and that bromide has an oxidation number of -1.

All atoms that are involved in a reaction can be given an oxidation number.

An element, Na or O_2 is always given an oxidation state of zero (0), any element that has reacted has an oxidation state of + or -. As removing electrons is **reduction**, if, in a reaction the element becomes **more** negative it has been reduced, if it becomes more positive it has been oxidised.

You can read about the rules for assigning oxidation numbers here:
http://www.dummies.com/how-to/content/rules-for-assigning-oxidation-numbers-to-elements.html

Elements that you expect to have a specific oxidation state actually have different states, so for example you would expect chlorine to be -1, it can have many oxidation states: NaClO, in this compound it has an oxidation state of +1

There are a few simple rules to remember:

- Metals have a + oxidation state when they react.
- Oxygen is 'king' it always has an oxidation state of -2
- Hydrogen has an oxidation state of +1 (except metal hydrides)
- The charges in a molecule must cancel.
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Examples: Sodium nitrate, $NaNO_3$
Where sodium has an oxidation state of +1 and O has an oxidation state of -2, but as there are 3 oxygen ions this is times 3 = -6. If the compound is stable, then the oxidation number of N must be +5

Q2. Work out the oxidation state of the **underlined** atom in the following:

a) $Mg\underline{C}O_3$ b) $\underline{S}O_3$ c) $Na\underline{Cl}O_3$ d) $\underline{Mn}O_2$ e) $\underline{Fe}_2O_3$ f) $\underline{V}_2O_5$ g) $K\underline{Mn}O_4$ h) $\underline{Cr}_2O_7^{2-}$ i) $\underline{Cl}_2O_4$

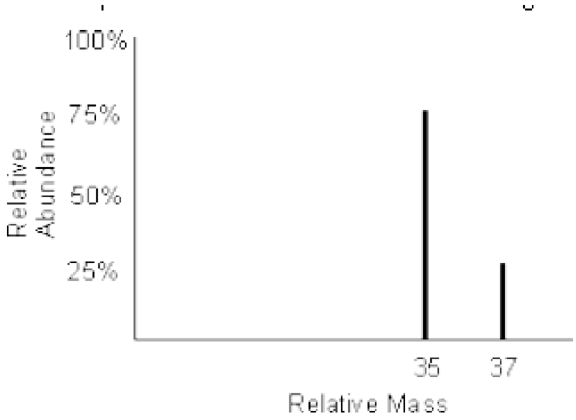
Chemistry topic 3 – Isotopes and mass

You will remember that isotopes are elements that have differing numbers of neutrons. Hydrogen has 3 isotopes.
Isotopes occur naturally, so in a sample of an element you will have a mixture of these isotopes. We can accurately measure the amount of an isotope using a **mass spectrometer**.

Q3.1 What must happen to the atoms before they are accelerated in the mass spectrometer?

Q3.2 Explain why the different isotopes travel at different speeds in a mass spectrometer.

A mass spectrum for the element chlorine will give a spectrum like this:
75% of the sample consist of chlorine-35, and 25% of the sample is chlorine-37. Given a sample of naturally occurring chlorine ¾ of it will be Cl-35 and ¼ of it is Cl-37. We can calculate what the **mean** mass of the sample will be:
Mean mass = (75 x 35)/100 + (25 x 37)/100 = 35.5



If you look at a periodic table (one attached at the end of this document) this is why chlorine has an atomic mass of 35.5.

An A level periodic table has the masses of elements recorded much more accurately than at GCSE. Most elements have isotopes, and these have been recorded using mass spectrometers.

Given the percentage of each isotope you can calculate the mean mass which is the accurate atomic mass for that element.

Q3.3 Use the percentages of each isotope to calculate the accurate atomic mass of the following elements.

a) Antimony has 2 isotopes: Sb-121 57.25% and Sb-123 42.75%

b) Gallium has 2 isotopes: Ga-69 60.2% and Ga-71 39.8%

c) Silver has 2 isotopes: Ag-107 51.35% and Ag-109 48.65%

d) Thallium has 2 isotopes: Tl-203 29.5% and Tl-205 70.5%

e) Strontium has **4** isotopes: Sr-84 0.56%, Sr-86 9.86%, Sr-87 7.02% and Sr-88 82.56%

Chemistry topic 4 – The shapes of molecules and bonding.

Have you ever wondered why your teacher drew a water molecule like this?
The lines represent a covalent bond, but why draw them at an unusual angle?

If you are unsure about covalent bonding, read about it here:

If you are unsure about ionic bonding, read about it here:

If you are unsure about metallic bonding, read about it here:

Q4.1 Draw a dot and cross diagram to show the bonding in a molecule of aluminium chloride ($AlCl_3$)

Q4.2 Draw a dot and cross diagram to show the bonding in a molecule of ammonia (NH_3)

Q4.3 Draw a dot and cross diagram to show the bonding in a molecule of methane (CH_4)

Q4.4 Draw a dot and cross diagram to show the bonding in the molecule of water (H_2O)

Q4.5 Draw a labelled model of metallic bonding in sodium

Chemistry topic 5 – Chemical equations

Balancing chemical equations is the stepping stone to using equations to calculate masses in chemistry.

There are loads of websites that give ways of balancing equations and lots of exercises in balancing.

Some of the equations to balance may involve strange chemicals, don't worry about that, the key idea is to get balancing right.
http://www.chemteam.info/Equations/Balance-Equation.html

This website has a download; it is safe to do so:
https://phet.colorado.edu/en/simulation/balancing-chemical-equations

Q5. Balance the following equations and name the type of reaction is i.e. neutralisation reaction

a. $H_2 + O_2 \rightarrow H_2O$
b. $S_8 + O_2 \rightarrow SO_3$
c. $HgO \rightarrow Hg + O_2$

d. $Zn + HCl \rightarrow ZnCl_2 + H_2$
e. $Na + H_2O \rightarrow NaOH + H_2$
f. $C_{10}H_{16} + Cl_2 \rightarrow C + HCl$
g. $Fe + O_2 \rightarrow Fe_2O_3$
h. $C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$
i. $Fe_2O_3 + H_2 \rightarrow Fe + H_2O$
j. $Al + FeO \rightarrow Al_2O_3 + Fe$

More practice on pages: 26-34- I suggest you learn the ionic equations to its very useful.

Chemistry topic 6 – Measuring chemicals – the mole

From this point on you need to be using an A level periodic table, not a GCSE one you can view one here:
https://www.ocr.org.uk/images/302739-units-h032-and-h432-data-sheet.pdf

Now that we have our chemical equations balanced, we need to be able to use them to work out masses of chemicals we need, or we can produce.
The ***mole*** is the chemists equivalent of a dozen, atoms are so small that we cannot count them out individually, we weigh out chemicals.

For example:
magnesium + sulfur → magnesium sulfide $Mg + S \rightarrow MgS$

We can see that one atom of magnesium will react with one atom of sulfur, if we had to weigh out the atoms, we need to know how heavy each atom is.
From the periodic table: Mg = 24.3 and S = 32.1

If I weigh out exactly 24.3g of magnesium this will be 1 mole of magnesium, if we counted how many atoms were present in this mass it would be a huge number (6.02×10^{23}), if I weigh out 32.1g of sulfur then I would have 1 mole of sulfur atoms.
So, 24.3g of Mg will react precisely with 32.1g of sulfur and will make 56.4g of magnesium sulfide.

Here is a comprehensive page on measuring moles, there are several descriptions, videos and practice problems.

Maybe BBC bitesize and physics and maths tutor
You will find the first 6 tutorials of most use here, and problem sets 1 to 3.
http://www.chemteam.info/Mole/Mole.html

Q6. Answer the following questions on moles.

a) How many moles of phosphorus pentoxide (P_4O_{10}) are in 85.2g?
b) How many moles of potassium in 73.56g of potassium chlorate (V) ($KClO_3$)?
c) How many moles of water are in 249.6g of hydrated copper sulfate(VI) ($CuSO_4.5H_2O$)? For this one, you need to be aware the dot followed by $5H_2O$ means that the molecule comes with 5 water molecules, so these must be counted in as part of the molecules mass.
d) What is the mass of 0.125 moles of tin sulfate ($SnSO_4$)?
e) If I have 2.4g of magnesium, how many g of oxygen(O_2) will I need to react completely with the magnesium? $2Mg + O_2 \rightarrow MgO$

Chemistry topic 7 – Solutions and concentrations

In chemistry a lot of the reactions we carry out involve mixing solutions rather than solids, gases or liquids.

You will have used bottles of acids in science that have labels e.g. 'Hydrochloric acid 1M', this is a solution of hydrochloric acid where 1 mole of HCl, hydrogen chloride (a gas) has been dissolved in 1dm3 of water.
The dm^3 is a cubic decimetre, it is actually 1 litre, but from this point on as an A level chemist you will use the dm^3 as your volume measurement.

http://www.docbrown.info/page04/4_73calcs11msc.htm

Q7.

a) What is the concentration (in mol dm^{-3}) of 9.53g of magnesium chloride ($MgCl_2$) dissolved in ^{100cm3} of water?

b) What is the concentration (in mol dm-3) of 13.248g of lead nitrate ($Pb(NO_3)_2$) dissolved in $2dm^3$ of water?

c) If I add $100cm^3$ of 1.00 mol dm^3 HCl to $1.9dm^3$ of water, what is the molarity of the new solution?

d) What mass of silver is present in $100cm^3$ of $1moldm^{-3}$ silver nitrate ($AgNO_3$)?

e) The Dead Sea, between Jordan and Israel, contains 0.0526 $moldm^{-3}$ of Bromide ions (Br -), what mass of bromine is in $1dm^3$ of Dead Sea water?

More help on pages 35-50

Chemistry topic 8 – Titrations

One key skill in A level chemistry is the ability to carry out accurate titrations, you may well have carried out a titration at GCSE, at A level you will have to carry them out very precisely **and** be able to describe in detail how to carry out a titration - there will be questions on the exam paper about how to carry out practical procedures.
You can read about how to carry out a titration here,

Remember for any titration calculation you need to have a balanced symbol equation; this will tell you the ratio in which the chemicals react.

E.g. a titration of an unknown sample of sulfuric acid with sodium hydroxide.
A 25.00cm^3 sample of the unknown sulfuric acid was titrated with 0.100moldm^{-3} sodium hydroxide and required exactly 27.40cm^3 for neutralisation. What is the concentration of the sulfuric acid?

Step 1: the equation $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

Step 2; the ratios 2:1

Step 3: how many moles of sodium hydroxide 27.40cm^3 = 0.0274dm^3

number of moles = c x v = 0.100 x 0.0274 = 0.00274 moles

Step 4: Using the ratio, how many moles of sulfuric acid
for every 2 NaOH there are 1 H_2SO_4 so, we must have 0.00274/2 =0.00137 moles of H_2SO_4

Step 5: Calculate concentration. concentration = moles/volume (in dm^3) = 0.00137/0.025 = ***0.0548 moldm^{-3}***

Here are some additional problems, which are harder, ignore the questions about colour changes of indicators. http://bit.ly/pixlchem12

Use the steps on the last page to help you

Q8 A solution of barium nitrate will react with a solution of sodium sulfate to produce a precipitate of barium sulfate.

$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$

What volume of 0.25moldm^{-3} sodium sulfate solution would be needed to precipitate all the barium from 12.5cm^3 of 0.15 moldm^{-3} barium nitrate?

Chemistry topic 9 – Organic chemistry – functional groups

At GCSE you would have come across **hydrocarbons** such as alkanes (ethane etc) and alkenes (ethene etc). You may have come across molecules such as alcohols and carboxylic acids. At A level you will learn about a wide range of molecules that have had atoms added to the carbon chain. These are called functional groups; they give the molecule certain physical and chemical properties that can make them incredibly useful to us.

Here you are going to meet a selection of the functional groups, learn a little about their properties and how we give them logical names.
Use the Bitesize revision pages to refresh your memories.
https://www.bbc.com/bitesize/topics/zx7mn39

Q9

Draw the following structures:

(a) Ethanol, (b) ethanoic acid (c) Ethyl Ethanoate (d) Propene (e) 2 methyl propanol (f) 1 chloro-2-methyl pentane.
(b) How do you make an ester- give a named example?

Chemistry topic 10 – Acids, bases, pH

At GCSE you will know that an acid can dissolve in water to produce H+ ions, at A level you will need a greater understanding of what an acid or a base is.

Q10.1 Find the A level definition of what an acid is?
Q10.2 How does ammonia (NH_3) act as a base?
Q10.3 Ethanoic acid (vinegar) is a weak acid, what does this mean?

Checkpoint Task
Amount of substance

Introduction

Often learners have very different levels of understanding when it comes to chemical calculations and the mole. This activity will probe your understanding of the words used to describe chemical quantities and amounts, and how they relate to the symbols (balanced equations), calculations and observations that happen in chemical reactions. You may find that your ideas change and evolve as you discuss the activity, so don't be concerned if you find some of the concepts difficult.

Task 1

Read through each of the ten statements below. With a partner or in a small group, discuss whether you think each statement is true or false and make a note of your answers.

1. The total number and type of atoms present are the same at the start and end of a reaction.

2. The amount of substance, measured in moles, is the same at the start and end of a reaction.

3. The total mass of reactants is equal to the total mass of products for any reaction.

4. The total volume of gas is the same at the start and the end of a reaction.

5. The amount in moles is proportional to the number of particles for that substance.

6. One mole of methane molecules (CH_4) contains 1/5 mole of carbon atoms and 4/5 mole of hydrogen atoms.

7. One mole of methane molecules (CH_4) contains 1 mole of carbon atoms and 4 moles of hydrogen atoms.

8. 100 cm^3 of methane gas contains the same number of molecules as 100 cm^3 hydrogen gas at room temperature and pressure.

9. 100 cm^3 of methane gas at room temperature and pressure has the same mass as 100 cm^3 of hydrogen gas under the same conditions.

10. If 0.1 mol of magnesium atoms reacts with a solution containing 0.1 mol of hydrochloric acid, 0.1 mol of hydrogen molecules will be produced. (Hint – you may need to look up or work out the balanced equation for this reaction.)

Task 2

Now for the difficult bit! For each of the statements you will need to justify your true/false answer with an explanation for example. If you have decided that a statement is true, try to give an explanation using the chemical concepts and definitions you know. If you have decided that a statement is false, you could find an example of a chemical process, reaction or balanced equation where it is not the case. You are free to look up information using whatever resources you have available to assist you with your explanations.

Extension

Read through the statements again and imagine you are trying to teach the concept of 'amount of substance' to a class of younger pupils who are having difficulty understanding. What practical demonstrations / activities / everyday examples can you think of that will help them understand?

If you found this hard go to the additional tasks on pages 35-50

Enthalpy Change:
Watch:

The Whole of AQA - ENERGY CHANGES. GCSE 9-1 Chemistry or Combined Science Revision Topic 5 for C1

And

The Whole of AQA-THE RATE AND EXTENT OF CHEMICAL CHANGE. GCSE Chemistry Combined Science Revision C2

Energy and fuels

1. Liquefied petroleum gas (LPG) is a type of fuel that can be used in heating appliances, cooking equipment and vehicles. One type of LPG consists mainly of propane (C_3H_8).

   a) Write an equation to represent the standard enthalpy change of formation of propane.

   b) Write an equation to represent the standard enthalpy change of combustion of propane.

   c) Use your equations and the enthalpy change of formation values below to calculate the standard enthalpy change of combustion of propane.

Substance	$\Delta_f H^{\circ}$ / kJ mol^{-1}
$C_3H_8(g)$	−104.5
$CO_2(g)$	−394
$H_2O(l)$	−286

2. Butanol ($C_4H_{10}O$) is a fuel that can be used in standard petrol engines. It is being investigated as a potential biofuel, though it is currently not efficient to produce.

a) Write an equation to represent the standard enthalpy change of combustion of butanol.

b) Write an equation to represent the standard enthalpy change of formation of butanol.

c) Use the standard enthalpy change of combustion values below to calculate the enthalpy change of formation of butanol.

Substance	$\Delta_c H^\circ$ **/ kJ mol^{-1}**
C(s)	−394
H_2(g)	−286
$C_4H_{10}O$(l)	−2676

3. Kerosene has many uses, including as cooking fuel and jet fuel.

A kerosene burner was used to heat 100 g of water. 0.225 g of fuel was used to increase the temperature of the water by 20.0 °C.

a) Calculate the amount of heat, in kJ, transferred to the water.

The specific heat capacity of water, *c*, is 4.18 $J\,K^{-1}\,g^{-1}$.

b) Assume the chemical formula for kerosene is $C_{12}H_{26}$. Calculate the amount of kerosene, in moles, used in the experiment.

c) Calculate the enthalpy change of combustion of kerosene.

d)The data book value for the standard enthalpy change of combustion of kerosene is −7153 $kJ\,mol^{-1}$.

Suggest two reasons why the experimental value is different from the

4. A mixture of dinitrogen tetroxide (N_2O_4) and hydrazine (N_2H_4) is used as a propellant for deep space rockets. When they react, they produce steam and nitrogen gas.

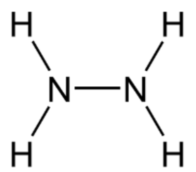
a) Construct a balanced equation for the reaction.

b) Use oxidation numbers to explain why this is a redox reaction.

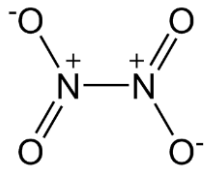
c) Use the mean bond enthalpies in the table below to calculate the enthalpy change for this reaction.

Bond	N–H	N–N	N–O (in N_2O_4)	O–H	N≡N
Average bond enthalpy / $kJ\,mol^{-1}$	391	163	404	463	941

Use the following structures to help you. You can assume that all N–O bonds in N_2O_4 have the bond enthalpy value show in the table.



N_2H_4



N_2O_4

d) Suggest why this combination makes a good rocket propellant.

5. The recommended daily intake of an average man is 2500 dietary calories per day.

1 kJ = 0.239 dietary calories.

a) A small bag of potato crisps contains 25 g of the snack.

A 2.0 g sample of the snack was burnt completely, and the heat energy was used to increase the temperature of 500 g of water in a calorimeter. The temperature of the water in the calorimeter increased by 20.9 °C.

(i) Calculate the energy, in kJ, that would be released if the contents of the entire bag were burnt.
The specific heat capacity of water, *c*, is $4.18\,J\,K^{-1}\,g^{-1}$.

(ii) Assume that the same amount of energy is released from crisps when you eat them, as when you burn them.

How many bags of crisps would an average man need to provide the recommended daily intake of dietary calories?
(This is not a good way to meet your calorie needs!)

b) The main component of boiled sweets is sucrose, $C_{12}H_{22}O_{11}$.

(i) Write an equation for the complete combustion of sucrose.

(ii) The standard enthalpy change of combustion of sucrose is –5644 kJ mol^{-1}.

Calculate the energy released when one sweet containing 6.70 g of sucrose is completely burnt.

(iii) How many sweets would an average man need to provide the recommended daily intake of dietary calories?
(This is also not a good way to meet your calorie needs!)

Bonding:

Watch:

[The Whole of AQA - BONDING, STRUCTURE AND PROPERTIES. GCSE Chemistry or Combined Science Revision.]

[The CHEMICAL BONDS Song - NOW WITH CLOSED CAPTION SO YOU CAN SING ALONG! Mr. Edmonds -]

[Atomic Bonding Song]

Checkpoint Task

Bonding and structure

Student Activity

Introduction

In your study of Bonding and Structure at A Level, you will be building a lot on ideas that you have already covered previously. Because bonding is a complex subject that is often simplified at GCSE, many learners can have unclear ideas or misconceptions about the topic. This activity will encourage you to explore what you already understand about chemical bonding, and to identify those areas that you still struggle with or require refinement at A Level.

Task 1

Here are twenty statements about chemical bonding. Discuss each statement within your pair or group and decide whether the statement is always true (unbreakable rule) or usually true (rule of thumb).

Separate the statements into two piles – always true or usually true. Make sure that you discuss the statements with your partner or in a group as you may have different ideas. It is more important to think carefully about each statement than to get to the end of the activity.

Task 2

For the statements that you think are not always true, try to think up some exceptions to the rule. You could use an equation or example element or compound to illustrate the 'exception to the rule'. Feel free to consult textbooks or other resources to help you with this.

Extension

If you have confidently identified all of the 'rules of thumb' and provided each with an exception to the rule, now see if you can think of any other parts of chemistry where we use 'rules of thumb', analogies or simplifications that are not strictly true. Some topics you could think about include atomic structure, reactivity, solubility, or acid–base theory.

Statements for use in activity

A. The atoms of Group 2 elements have two electrons in their outer shell.	**B. Noble gases do not form any types of bonds because they have full outer shells.**
C. Ionic substances have higher melting points than covalent substances.	**D. Oppositely charged ions attract.**
E. Delocalised electrons are more stable than electrons in fixed atomic orbitals.	**F. Energy is released when ionic bonds form.**
G. In an ionic compound, ions are combined in proportions which balance out the electrical charges.	**H. Energy is needed to break covalent bonds.**
I. Energy is required to form positive ions from atoms.	**J. Energy is released when negative ions are formed from atoms.**

Other bonding activities:

http://www.rsc.org/learn-chemistry/resource/res00000617/atoms-elements-molecules-compounds-and-mixtures

http://www.rsc.org/learn-chemistry/resources/gridlocks/puzzles/level-2/types-of-bonding.html

Allotropes of carbon, RSC

http://www.rsc.org/learn-chemistry/resource/res00001121/ri-christmas-lectures-2012- allotropes-of-carbon?cmpid=CMP00002102

Bonding Bingo, RSC

http://www.rsc.org/learn-chemistry/resource/res00000093/afl-bonding-bingo

We will go over this when you return to sixth form.

All question boxes need to be completed and handed in for the first chemistry lesson in September. You can just write the answers on paper (no need to print the document). There will be a transition test in week 3 to test prior knowledge and knowledge covered in the first 2 weeks of term. Additional task to complete have been added to the end of this document to strengthen your understanding.

Extra support booklets:

Maths skills:
https://www.ocr.org.uk/Images/295468-chemistry-mathematical-skills-handbook.pdf

Practical skills: written for teachers but has some useful examples on how to do the calculation's in practical's and the skills being looked at in each PAG next year.
https://www.ocr.org.uk/Images/208932-chemistry-practical-skills-handbook.pdf

Research: Careers in chemistry find out where this subject can take you.

https://edu.rsc.org/future-in-chemistry/career-options

Additional tasks: Chemical Equations

Watch: https://www.youtube.com/watch?v=T0wb4z-_kmY

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:

hydrogen + oxygen → water

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 H_2
- oxygen is represented as O_2
- water is

H_2O. So, we get:

$$H_2 + O_2 \rightarrow H_2O$$

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 + oxygen → carbon dioxide

$$C + O_2 \rightarrow CO_2$$

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 O_2.

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 + oxygen $\rightarrow$ magnesium oxide

Magnesium is written as Mg (one atom just like carbon) and oxygen is, O_2, but magnesium oxide has just one atom of oxygen per molecule and is therefore written as MgO.

$$Mg + O_2 \rightarrow MgO$$

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:

$$Mg + O_2 \rightarrow 2MgO$$

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:

$$2Mg + O_2 \rightarrow 2MgO$$

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:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Exercise: Balancing equations

Balance the following equations. 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!

1. $CH_4 + O_2 \rightarrow CO_2 + H_2O$

---

2. $C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$

---

3. $PCl_3 + H_2O \rightarrow H_3PO_3 + HCl$

---

4. $8HNO_3 + 3Cu \rightarrow Cu\ (NO_3)_2 + NO + H_2O$

---

5. $4HNO_3 + Cu \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

---

6. $H_3PO_4 + NaOH \rightarrow NaH_2PO_4 + H_2O$

---

7. $H_3PO_4 + NaOH \rightarrow Na_3PO_4 + H_2O$

---

8. $N_2 + H_2 \rightarrow NH_3$

9. $NaBr + H_2SO_4 \rightarrow Na_2SO_4 + HBr$

10. $HBr + H_2SO_4 \rightarrow H_2O + SO_2 + Br_2$

11. $C_2H_5OH + PCl_3 \rightarrow C_2H_5Cl + H_3PO_3$

12. $Fe_3O_4 + H_2 \rightarrow Fe + H_2O$

13. $Fe_2O_3 + CO \rightarrow Fe + CO_2$

14. $C_2H_5OH + CH_3CO_2H \rightarrow CH_3CO_2C_2H_5 + H_2O$

15. $2KMnO_4 + HCl \rightarrow MnCl_2 + Cl_2 + 8H_2O + KCl$

Exercise: What's wrong here?

The following equations all contain one or more mistakes. These may be in a formula, in the balancing, in the state symbols or even in the chemistry. You need to identify the error and then write out a correct equation.

1	$Na(s)$	+	$H_2O(l)$	→	$NaOH(aq)$	+	$H(g)$	
2	$PbNO_3(aq)$	+	$NaCl(aq)$	→	$PbCl(s)$	+	$NaNO_3(aq)$	
3	$CaOH_2(aq)$	+	$2HCl(aq)$	→	$CaCl_2(aq)$	+	$2H_2O(l)$	
4	$C_2H_4(g)$	+	$2O_2(g)$	→	$2CO_2(g)$	+	$2H_2(g)$	
5	$MgSO_4(aq)$	+	$2NaOH$	→	$Ca\ (OH)_2(s)$	+	$Na_2SO_4(aq)$	
6	$Cu\ (NO_3)_2(s)$	+	$CuO(s)$	→	$2NO(g)$	+	$O_3(g)$	
7	$Cu(s)$	+	$H_2SO_4(aq)$	→	$CuSO_4(aq)$	+	$H_2(g)$	

Ionic equations- very important to understand with A level Chem

Watch:

[Formula of common ions you MUST LEARN!! Flashcards for new 9-1 GCSE Chemistry or combined science]

Ionic theory

Many of the chemicals you will use at GCE Advanced Level 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 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 for example hydrogen chloride and sulfuric acid.

Structures of ionic compounds

During your course you will study bonding and structure, and some 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 the 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 found in them. For example, 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

For *ionic* chemicals it is the ions which react, not the molecules. For instance, copper (II) sulfate is usually written as $CuSO_4$ but it is more often the ion Cu^{2+} which reacts. When you write out an ionic equation you should include only the ions which **actually take part in the reaction**.

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

$$FeSO_4(aq) \quad + \quad 2NaOH(aq) \quad \rightarrow \quad Fe(OH)_2(s) \quad + \quad Na_2SO_4(aq)$$

In water, the iron (II) sulfate 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 sulfate ions.

$$Fe^{2+}(aq) \quad + \quad 2OH^{-}(aq) \quad \longrightarrow \quad Fe\,(OH)_2(s)$$

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 as:

$$Ag^{+}(aq) \quad + \quad Cl^{-}(aq) \quad \longrightarrow \quad AgCl(s)$$

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) sulfate solution with excess sodium hydroxide solution.

1 Write down the balanced molecular equation.

$$FeSO_4(aq) \quad + \quad 2NaOH(aq) \quad \longrightarrow \quad Fe\,(OH)_2(s) \quad + \quad Na_2SO_4(aq)$$

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

$$Fe^{2+}(aq) + \cancel{SO_4}^{2-}(aq) + \cancel{2Na}^{+}(aq) + 2OH^{-}(aq) \rightarrow Fe\,(OH)_2(aq) + \cancel{2Na}^{+}(aq) + \cancel{SO_4}^{2-}(aq)$$

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

$$Fe^{2+}(aq) \quad + \quad 2OH^{-}(aq) \quad \longrightarrow \quad Fe\,(OH)_2(s)$$

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

Exercise Ionic equations

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

1 Pb^{2+}(aq) + OH^-(aq) $\rightarrow$ $Pb(OH)_2$(s)

2 Al^{3+}(aq) + OH^-(aq) $\rightarrow$ $Al(OH)_3$(s)

3 $Al(OH)_3$(s) + OH^-(aq) $\rightarrow$ AlO_2^{2-}(aq) + H_2O(l)

4 Cl_2(g) + OH^-(aq) $\rightarrow$ ClO_3^{3-}(aq) + Cl^-(aq) + H_2O(l)

5 $S_2O_3^{2-}$(aq) + I_2(s) $\rightarrow$ S_4O^{2-}(aq) + $2I^-$(aq)

6 $Cu^{2+}(aq)$ + $OH^{-}(aq)$ $\longrightarrow$

---

7 $CO_3^{2-}(s)$ + $H^{+}(aq)$ $\longrightarrow$

---

8 $Zn(s)$ + $H^{+}(aq)$ $\longrightarrow$

---

9 $Zn(s)$ + $Pb^{2+}(aq)$ $\longrightarrow$

---

10 $H^{+}(aq)$ + $OH^{-}(aq)$ $\longrightarrow$

Chemical Calculations:

Watch: The Whole of AQA -QUANTITATIVE CHEMISTRY. GCSE Chemistry or Combined Science Revision Topic 3 for C1

Mass Calculations:

Examples: Calculation of Molar Mass from relative atomic mass data

Before you start these questions make sure you read *Section 4*: *The mole* of this workbook. When you carry out experiments you will weigh chemicals in grams. Molar Mass has the same numerical value as *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 sulfuric 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 **sulfur** of mass 32.1	= 1 x 32.1	= 32.1 g mol^{-1}
4 atoms of **oxygen** of mass 16	= 4 x 16	= 64 g mol^{-1}
	Total mass	**= 98.1 g mol^{-1}**

Example 2

Calculate the Molar Mass of lead nitrate

$Pb(NO_3)_2$ Care! This molecule contains **TWO** nitrate groups.

1 atom of h**ead** of mass 207.2	= 1 x 207.2	= 207.2 g mol^{-1}
2 atoms of **nitrogen** of mass 14	= 2 x 14	= 2 8 g mol^{-1}
6 atoms of **oxygen** of mass 16	= 6 x 16	= 9 6 g mol^{-1}
	Total mass	**= 331.2 g mol^{-1}**

Example 3

Calculate the Molar Mass of $CuSO_4.5H_2O$

Care! This molecule has 5 molecules of water attached to each molecule of copper sulfate. 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	= 6 3 . 5 g mol^{-1}
	1 atom of **sulfur** of mass 32.1	= 1 x 32.1	= 3 2 . 1 g mol^{-1}
	4 atoms of **oxygen** each of mass 16	= 4 x 16	= 6 4 g mol^{-1}
In $5H_2O$	5 x 2 atoms of **hydrogen** each of	= 10 x 1	= 1 0 g mol^{-1}
	5 x 1 atoms of **oxygen** each of mass	= 5 x 16	= 8 0 g mol^{-1}
		Total mass	**= 249.6 g mol^{-1}**

Calculations of this type are generally written as follows

$CuSO_4.5H_2O$ = [63.5 + 32.1 + (4 x 16) + 5((2 x 1) + 16)] = 249.6 g mol^{-1}

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 (in *Section 12*). 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	H_2O
2	CO_2
3	NH_3
4	C_2H_5OH
5	C_2H_4
6	SO_2
7	SO_3
8	HBr
9	H_2SO_4
10	HNO_3
11	$CaSO_4$
12	$BaCl_2$
13	$AlCl_3$
14	$Al(NO_3)_3$
15	$Al_2(SO_4)_3$
16	$FeSO_4$
17	$FeCl_2$
18	$FeCl_3$
19	$Fe_2(SO_4)_3$

20 PbO

21 $PbCl_2$

22 $PbSO_4$

23 CuCl

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 6.02×10^{23} 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 Advanced 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}.*

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

$$\therefore \text{number of moles of oxygen} = \frac{64\text{ g of oxygen atoms}}{\text{RAM } O_2 \; 16\text{ g mol}^{-1}}$$

= 4 moles of oxygen atoms

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

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

The first stage of this calculation is to calculate the molar mass (RFM) of chlorine molecules. Molar mass of Cl_2 = 2 x 35.5 = 71 g mol^{-1}

$$\therefore \text{number of moles of } Cl_2 = \frac{142\text{ g of chlorine gas}}{\text{RFM chlorine } 71\text{ g mol}^{-1}}$$

= **2 moles of chlorine molecules**

c Calculate the number of moles of $CuSO_4.5H_2O$ in 100 g of the solid.

The Relative Molecular Mass of $CuSO_4.5H_2O$ =

[63.5 + 32.1 + (4 x 16) + 5{(2x1) + 16}] = 249.6 g mol^{-1}

$$\therefore \text{number of moles of } CuSO_4.5H_20 = \frac{100\text{ g of } CuSO_4.5H_2O}{\text{RFM of } CuSO_4.5H_20 = 249.5\text{ g mol}^{-1}}$$

= **0.4006 moles of $CuSO_4.5H_2O$ 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 sulfur dioxide SO_2.

1 mole of sulfur dioxide has a mass = 32.1 + (2 x 16) = 64.1 g mol^{-1}
∴ 3 moles of SO_2 = 3 x 64.1 **= 192.3 g**

b What is the mass of 0.05 moles of $Na_2S_2O_3.5H_2O$?

1 mole of $Na_2S_2O_3.5H_2O$ = [(23 x 2) + (32.1 x 2) + (16 x 3)] + 5[(2 x 1) +16] = 248.2 g mol^{-1}
∴ 0.05 moles of $Na_2S_2O_3.5H_2O$ = 0.05 x 248.2 **= 12.41**

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 dm^3 (24,000 cm^3) at room temperature and pressure**.

$$\therefore The\ volume\ of\ a\ given\ number\ of\ moles\ of\ gas = number\ of\ moles\ x\ 24000cm^3$$

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.

∴ 2 moles of carbon dioxide = 2 x 24 000 cm^3 = 48 000 cm^3 **= 48 dm^3**

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

Volume of 0.0056 moles of chlorine = 0.0056 x 24 000 cm^3 **= 134.4 cm^3**

Example 4

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

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

a Calculate the number of moles of hydrogen molecules in 240 cm^3 of the gas.

$$number\ of\ moles = \frac{240\ cm^3}{24\ 000\ cm^3} = 0.010\ moles$$

b How many moles of a gas are there in 1000 cm^3 of the gas?

number of moles of gas = 1000 cm^3/24000cm^3

= 0.0147 moles

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.

$$\therefore \text{number of moles of hydrogen } (H_2) = \frac{10\text{ g of hydrogen } H_2}{\text{RFM of } H_2 = 2\text{ mol}^{-1}}$$

10/2 =5 moles

∴ 5 moles of hydrogen = 5 x 24 000 cm^3 = 120 000 cm^3 **= 120 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 cm^3 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.

$$\therefore \text{number of moles of } CO_2 = \frac{1000\text{ cm}^3\text{ of } CO_2}{\text{volume of 1 mole of } CO_2 \text{ of } 24\,000\text{ cm}^3}$$

= 0.0147 moles

∴ 0.0147 moles of carbon dioxide = 0.0147 x 44 g **= 1.833 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, i.e. 24 000 cm^3. This is the molar mass of the gas.

For example, calculate the Relative Molecular Mass of a gas for which 100 cm^3 of the gas at room temperature and pressure have a mass of 0.0667 g.

100 cm^3 of the gas has a mass of 0.0667 g.

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

$$= 16 \text{ g}$$

$\therefore$ **The molar mass of the gas is 16 g mol^{-1}**

Exercise: 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 H_2O

2 88.0 g of CO_2

3 1.70 g of NH_3

4 230 g of C_2H_5OH

5 560 g of C_2H_4

6 0.641 g of SO_2

7 80.1 g of SO_3

8 18.20 g of HBr

9 0.0981 g of H_2SO_4

10 3.15 g of HNO_3

11 19.3 g of NaCl

Exercise: Calculation of the mass of material in a given number of moles of at material

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

1	2 moles of H_2O
2	3 moles of CO_2
3	2.8 moles of NH_3
4	0.50 moles of C_2H_5OH
5	1.2 moles of C_2H_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_2SO_4
10	0.15 moles of HNO_3
11	0.45 moles of NaCl
12	0.70 moles of $NaNO_3$

13 0.11 moles of Na_2CO_3

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, i.e. moles and equations.

Many reactions take place in solutions of known concentration.

Concentration in solution is generally measured as moles per 1000 cm^3 **of solution**. For example, sodium chloride may be labelled as 1M NaCl. This means that each 1000 cm^3 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 cm^3 of water as the volume of the mixture may no longer be 1000 cm^3.

The solution will have been made up by measuring out 58.5 g of the solid, dissolving it in about 500 cm^3 of water and then adding more water to make the total volume of the mixture up to 1000 cm^3. (1 dm^3)

Concentration in mol dm^{-3} is called **molarity**.

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

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

r

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

Method

To carry out these calculations, you need to calculate the actual amounts of materials in the volumes involved.

Example

25 cm^3 of 0.10 mol dm^{-3} NaOH react with 50 cm^3 of a solution of H_2SO_4. What is the molarity of the H_2SO_4?

$$2NaOH(aq) \quad + \quad H_2SO_4(aq) \quad \longrightarrow \quad Na_2SO_4(aq) \quad + \quad 2H_2O(l)$$

∴ 2 mol of NaOH react with 1 mol of H_2SO_4.

In this case you know the concentration of the sodium hydroxide so
∴ 1 mol of NaOH reacts with 0.5 mol of H_2SO_4.

always put the reactant you know as '1 mol'.

In this reaction you have used 25 cm^3 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 x 2.5 x 10^{-3} moles of H_2SO_4 = 1.25 x 10^{-3} moles of H_2SO_4
∴ 1.25 x 10^{-3} moles of H_2SO_4 will be found in 50 cm^3 of the solution.
∴ In 1000 cm^3 of the acid the same solution there will be

$$= \frac{1000 \times (1.25 \times 10^{-3})}{50} \text{ moles of } H_2SO_4$$

$$= \quad 0.0250 \text{ moles}$$

∴ The concentration of the sulfuric acid is **0.025 mol dm^{-3}**.

Exercise: Calculations based on concentrations in solution

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

---

1 25 cm^3 of 1.0 mol dm^{-3} <u>HCl</u>

---

2 50 cm^3 of 0.5 mol dm^{-3} <u>HCl</u>

---

3 250 cm^3 of 0.25 mol dm^{-3} <u>HCl</u>

---

4 500 cm^3 of 0.01 mol dm^{-3} <u>HCl</u>

---

5 25 cm^3 of 1.0 mol dm^{-3} <u>NaOH</u>

---

6 50 cm^3 of 0.5 mol dm^{-3} <u>KOH</u>

---

7 50 cm^3 of 0.25 mol dm^{-3} <u>HNO_3</u>

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

8 3.65 g of HCl in 1000 cm^3 of solution

9 3.65 g of HCl in 100 cm^3 of solution

10 6.624 g of $Pb(NO_3)_2$ in 250 cm^3 of solution

11 1.00 g of NaOH in 250 cm^3 of solution

12 1.962 g of H_2SO_4 in 250 cm^3 of solution

13 1.58 g of $KMnO_4$ in 250 cm^3 of solution

14 25.0 g of $Na_2S_2O_3.5H_2O$ in 250 cm^3 of solution

15 25.0 g of $CuSO_4.5H_2O$ in 250 cm^3 of solution

16 4.80 g of $(COOH)_2.2H_2O$ in 250 cm^3 of solution

17 10.0 g of $FeSO_4.(NH_4)_2SO_4.6H_2O$ in 250 cm^3 of solution

18 240 cm^3 of $NH_3(g)$ dissolved in 1000 cm^3 of solution

19 480 cm^3 of HCl(g) dissolved in 100 cm^3 of solution

Organic chemistry:

Watch The Whole of AQA - ORGANIC CHEMISTRY. GCSE Chemistry or Combined Science Revision Topic 7 for C2

Followed by The Functional Group Concept Explained | Organic Chemistry | Chemistry | FuseSchool

And make notes as you go along- especially if you did not finish this in year 11.

Can you list the functional groups present in molecules?

Can you recall or work out their general formula?

Can you state a test for an alkene?

Can you write an equation for combustion of an alkane and for an alkene and an alcohol?

How do you test the products of combustion?

What is cracking? Why is it useful?

Can you write an equation for the cracking of decane?

Naming of compounds

At Advanced GCE 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 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- organic will be covered in the course next year.

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, e.g.

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

Chlorides contain an element and chlorine, e.g.

$MgCl_2$ 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, e.g.

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, e.g.

NaOH	is	**Sodium Hydroxide**
$Ca(OH)_2$	Is	**Calcium Hydroxide Potassium**
KCN	is	**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 Lead**
$PbCl_4$	is	**(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, e.g.:

CO	is	**Carbon <u>Monoxide</u>** where mon- means one
CO_2	is	**Carbon <u>Dioxide</u>** where di- means two
SO_2	is	**Sulfur Dioxide.** This could be called **Sulfur (IV)**
SO_3	is	**Sulfur Trioxide**. This could be called **Sulfur (VI) Oxide**
PCl_3	is	**Phosphorus Trichloride**. This could be called **Phosphorus (III) Chloride**
PCl_5	is	**Phosphorus Pentachloride**. This could be called **Phosphorus(V) Chloride**
CCl_4	is	**Carbon Tetrachloride**
$SiCl_4$	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 **$-CO_3$**, **Sulfate** **$-SO_4$**, **Nitrate** **$-NO_3$**

A compound of sodium, carbon and oxygen would be Na_2CO_3 and would be called **Sodium Carbon<u>ate</u>**. For example:

$NaNO_3$	Is	**Sodium Nitrate**
$Mg(NO_3)_2$	is	**Magnesium Nitrate**
$Fe_2(SO_4)_3$	Is	**Iron (III) Sulfate**
$FeSO_4$	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:

KNO_2	is	**Potassium Nitrite** or **Potassium Nitrate (III)**
Na_2SO_3	is	**Sodium Sulfite** or **Sodium Sulfate (IV)**
K_2CrO_4	is	**Potassium Chromate (VI)**
$KMnO_4$	is	**Potassium Manganate (VII)**
$KClO_3$	is	**Potassium Chlorate(V).**

In

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, e.g.

CH_3COONa 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 $\mathbf{NH_4^+}$ and sits in the place generally taken by a metal in a formula.

NH_4Cl is **Ammonium Chloride**

$(NH_4)_2SO_4$ is **Ammonium Sulfate**

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

Sulfuric 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.

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_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.

Organic compounds with 2 carbons will either start with **Eth-** or have **-eth-** in their name, e.g.

C_2H_4	is	**Ethene**
C_2H_5OH	is	**Ethanol**
CH_3COOH	is	**Ethanoic Acid.**
C_2H_5Cl	is	**Chloroethane**

Exercise 3: Names from formulae

1	$CaSO_4$	
2	$BaCl_2$	
3	$AlCl_3$	
4	$Al(NO_3)_3$	
5	$Al_2(SO_4)_3$	
6	$FeSO_4$	
7	$FeCl_2$	
8	$FeCl_3$	
9	$Fe_2(SO_4)_3$	
10	PbO	
11	PbO_2	
12	$Pb(NO_3)_2$	
13	$PbCl_2$	
14	$PbSO_4$	
15	$Cu(NO_3)_2$	
16	$CuCl$	
17	$CuCl_2$	
18	$CuSO_4$	
19	$ZnCl_2$	
20	$AgNO_3$	
21	NH_4Cl	

22	$(NH_4)_2SO_4$
23	NH_4VO_3 (V is Vanadium)
24	$KClO_3$

The Periodic Table of the Elements

Key
atomic number
Symbol
name
relative atomic mass

(1) **1**	(2) **2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	(3) **13**	(4) **14**	(5) **15**	(6) **16**	(7) **17**	(0) **18**
1 **H** hydrogen 1.0																	2 **He** helium 4.0
3 **Li** lithium 6.9	4 **Be** beryllium 9.0											5 **B** boron 10.8	6 **C** carbon 12.0	7 **N** nitrogen 14.0	8 **O** oxygen 16.0	9 **F** fluorine 19.0	10 **Ne** neon 20.2
11 **Na** sodium 23.0	12 **Mg** magnesium 24.3											13 **A*l*** aluminium 27.0	14 **Si** silicon 28.1	15 **P** phosphorus 31.0	16 **S** sulfur 32.1	17 **C*l*** chlorine 35.5	18 **Ar** argon 39.9
19 **K** potassium 39.1	20 **Ca** calcium 40.1	21 **Sc** scandium 45.0	22 **Ti** titanium 47.9	23 **V** vanadium 50.9	24 **Cr** chromium 52.0	25 **Mn** manganese 54.9	26 **Fe** iron 55.8	27 **Co** cobalt 58.9	28 **Ni** nickel 58.7	29 **Cu** copper 63.5	30 **Zn** zinc 65.4	31 **Ga** gallium 69.7	32 **Ge** germanium 72.6	33 **As** arsenic 74.9	34 **Se** selenium 79.0	35 **Br** bromine 79.9	36 **Kr** krypton 83.8
37 **Rb** rubidium 85.5	38 **Sr** strontium 87.6	39 **Y** yttrium 88.9	40 **Zr** zirconium 91.2	41 **Nb** niobium 92.9	42 **Mo** molybdenum 95.9	43 **Tc** technetium	44 **Ru** ruthenium 101.1	45 **Rh** rhodium 102.9	46 **Pd** palladium 106.4	47 **Ag** silver 107.9	48 **Cd** cadmium 112.4	49 **In** indium 114.8	50 **Sn** tin 118.7	51 **Sb** antimony 121.8	52 **Te** tellurium 127.6	53 **I** iodine 126.9	54 **Xe** xenon 131.3
55 **Cs** caesium 132.9	56 **Ba** barium 137.3	57–71 lanthanoids	72 **Hf** hafnium 178.5	73 **Ta** tantalum 180.9	74 **W** tungsten 183.8	75 **Re** rhenium 186.2	76 **Os** osmium 190.2	77 **Ir** iridium 192.2	78 **Pt** platinum 195.1	79 **Au** gold 197.0	80 **Hg** mercury 200.6	81 **T*l*** thallium 204.4	82 **Pb** lead 207.2	83 **Bi** bismuth 209.0	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
87 **Fr** francium	88 **Ra** radium	89–103 actinoids	104 **Rf** rutherfordium	105 **Db** dubnium	106 **Sg** seaborgium	107 **Bh** bohrium	108 **Hs** hassium	109 **Mt** meitnerium	110 **Ds** darmstadtium	111 **Rg** roentgenium	112 **Cn** copernicium		114 **F*l*** flerovium		116 **Lv** livermorium		

57 **La** lanthanum 138.9	58 **Ce** cerium 140.1	59 **Pr** praseodymium 140.9	60 **Nd** neodymium 144.2	61 **Pm** promethium 144.9	62 **Sm** samarium 150.4	63 **Eu** europium 152.0	64 **Gd** gadolinium 157.2	65 **Tb** terbium 158.9	66 **Dy** dysprosium 162.5	67 **Ho** holmium 164.9	68 **Er** erbium 167.3	69 **Tm** thulium 168.9	70 **Yb** ytterbium 173.0	71 **Lu** lutetium 175.0
89 **Ac** actinium	90 **Th** thorium 232.0	91 **Pa** protactinium	92 **U** uranium 238.1	93 **Np** neptunium	94 **Pu** plutonium	95 **Am** americium	96 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fermium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium