This is your booklet for the term.

Look after it.

If you need to take it away to do prep, it is your responsibility to bring it to the next lesson.

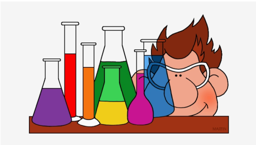
If you fail to bring your booklet to two consecutive lessons you'll be given another, your tutor contacted, an email sent home and a charge will be added to the bill.

If you manage to lose it and are unable to do your prep then it is your responsibility to download another copy of from **FROG**.

The Departmental Website.

You can log-on to the departmental website from anywhere by navigating to ellesmereonline.com and clicking

1. Pupils and Students
2. Science Link
3. Chemistry
4. AQA GCSE



10.1 - Tests for gases & ions

I have reviewed the syllabus statements for this topic	
I have completed the questions in this section	
I have read the relevant sections of the College Website	
I have made some revision material (mind-map, key-words & definitions etc)	
Prep Grade	
Test Grade	

Target	Pupil Signature

- The test for ***Hydrogen*** uses a burning splint held at the open end of a test tube of the gas. Hydrogen burns rapidly with a pop sound.
- The test for ***Oxygen*** relights a glowing splint inserted into a test tube of the gas.
- The test for CO_2 uses an aqueous solution of Calcium Hydroxide (***lime water***). When Carbon Dioxide is bubbled through Limewater the Limewater turns milky (cloudy).
- The test for ***Chlorine*** uses Litmus paper. When ***damp Litmus paper*** is put into Chlorine gas the Litmus paper is ***bleached*** and turns ***white***.

- ***Flame tests*** can be used to identify Lithium, Sodium, Potassium, Calcium and Copper compounds by producing distinctive colours in flame tests:

1. ***Lithium*** compounds result in a ***crimson*** flame
2. ***Sodium*** compounds result in a ***yellow*** flame
3. ***Potassium*** compounds result in a ***lilac*** flame
4. ***Calcium*** compounds result in an ***orange-red*** flame
5. ***Copper*** compounds result in a ***green*** flame.

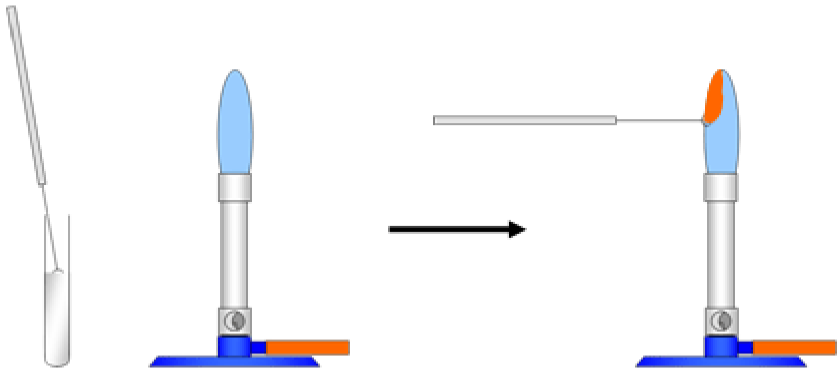
- If a sample containing a mixture of ions is used some flame colours can be masked.
- Elements and compounds can be detected and identified using ***instrumental methods***. Instrumental methods are ***accurate***, sensitive and rapid.
- Students should be able to state advantages of instrumental methods compared with the chemical tests in this specification.
- ***Flame emission spectroscopy*** is an instrumental method used to analyse metal ions in solutions. The sample is put into a flame & the light given out is passed through a spectroscope. The output is a ***line spectrum*** that can be analysed to identify the metal ions in the solution and measure their concentrations.
- Students should be able to interpret an instrumental result given appropriate data in chart or tabular form, when accompanied by a reference set in the same form, limited to flame emission spectroscopy.
- ***Sodium Hydroxide*** solution can be used to identify some metal ions (cations).
- Solutions of ***Aluminium***, ***Calcium*** and ***Magnesium*** ions form white precipitates when Sodium Hydroxide solution is added but only the Aluminium Hydroxide precipitate ***dissolves in excess Sodium Hydroxide*** solution.
- Solutions of ***Copper(II)***, ***Iron(II)*** and ***Iron(III)*** ions form ***coloured precipitates*** with ***Sodium Hydroxide***
- ***Copper***(***II***) forms a ***blue*** ppt, ***Iron(II)*** a ***green*** ppt & ***Iron(III)*** a brown ppt.
- Students ***should*** be able to write balanced equations for these reactions
- Equations are not expected for the production of Sodium Aluminate.
- ***Carbonates*** react with ***dilute acids*** to form ***Carbon Dioxide*** gas which can be identified with ***Limewater***.
- ***Halide*** ions in solution produce precipitates with Silver Nitrate solution ***in the presence of dilute Nitric acid***.
- ***Silver Chloride*** is ***white***, ***Silver Bromide*** is ***cream*** and ***Silver Iodide*** is ***yellow***.
- ***Sulphate*** ions in solution produce a ***white*** precipitate with ***Barium Chloride*** solution ***in the presence of dilute Hydrochloric acid***

Flame Tests.

The majority of solids are ***white*** powders and difficult to distinguish between just by looking.

Flame tests can give us a good indication of what ***metal ion*** it contains.

You should know the ***colours*** associated with four metal ions as well as the ***method*** used to carry out flame tests.



This lesson involves …

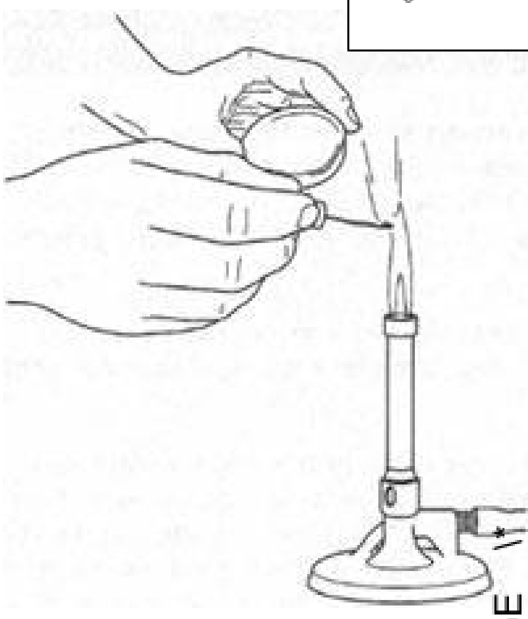
- *Collaborative
- *Concerned for Society
- *Confident

This lesson involves …

- *Generalisation
- *Connection-finding
- *Big Picture Thinking
- *Abstraction
- *Imagination

Method

Dip a Nichrome wire in a concentrated solution of Hydrochloric acid - this cleans the wire and makes it damp.



Nichrome is used because it is an alloy with a high melting point that will not affect the colour of the Bunsen flame.

The slightly damp wire is then dipped into the unidentified powder.

A blue Bunsen flame is needed and the colour change notice when the powder is placed in it should be noted.

Results

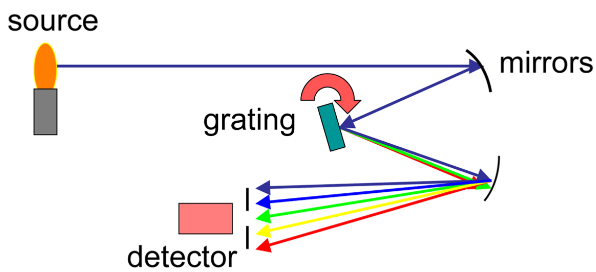
Metal ion	Formula of ion	Colour Seen
Lithium	Li^{+}	
Sodium	Na^{+}	
Potassium	K^{+}	
Calcium	Ca^{2+}	
Copper (II)	Cu^{2+}	

Problems with Flame tests.

It can be difficult to see the right colour even with a pure sample - with a sample that contains more than one ion we would not know what colours were mixed together.

And there are over 80 metals, ions of which often give very similar flames.

A ***Flame-emission Spectrometer*** splits up the colours from a flame and records the exact wavelength of the colours in the flame.



This lesson involves ...

*Generalisation
*Connection-finding
*Big Picture Thinking
*Abstraction
*Imagination

The wavelengths that each element emit are well established.

More than a dozen metal ions make blue flames that a person couldn't tell apart - but a machine can spot the difference with ease.

And if the sample is mixed, it can tell you exactly which metal ions are mixed together.

A solution containing more than one positive ion might give the following spectrum:

You can compare this spectrum against reference spectra:

Spectrum for ion X:

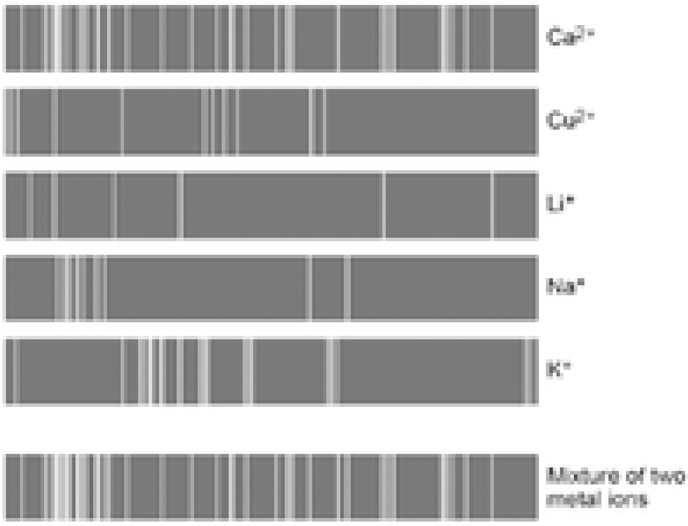
Spectrum for ion Y:

The combination of the spectra for ion X and for ion Y gives the spectra for the sample tested — this suggests the mixture contains ion X and ion Y.

Q1. (a) Flame emission spectroscopy can be used to analyse metal ions in solution.

Figure 3 gives the flame emission spectra of five metal ions, and of a mixture of two metal ions.

Figure 3



Use the spectra to identify the **two** metal ions in the mixture.

....

.... **(2)**

This lesson involves …

*Practice
*Perseverance
*Resilience

(b) Explain why a flame test could **not** be used to identify the two metal ions in the mixture.

....

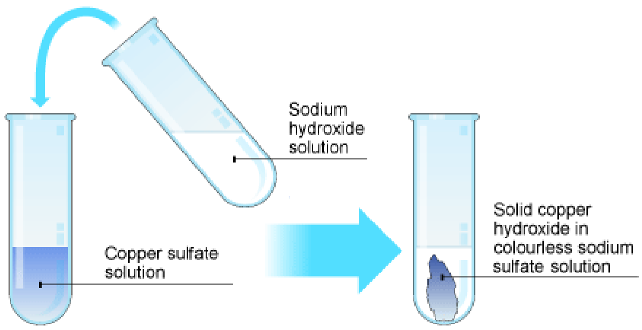
....

....

.... **(2)**

Other Cations

Precipitation Tests



- Pour 2cm depth of Cu^{2+}, Fe^{2+}, Fe^{3+} , Ca^{2+}, Mg^{2+} and Al^{3+} solutions into 6 separate test-tube.
- Carefully add Sodium Hydroxide solution to each until the test tube is around a third full
- 3Note the colour of the precipitate
- If the ppt is white, continue adding Sodium Hydroxide.

☐ SAFETY: EYE-PROTECTION

This lesson involves …

*Enquiring
*Creative/Enterprising
*Open-minded
*Risk-taking

Metal ion in solution	Formula of ion	…pt
Copper (II)	Cu^{2+}	
Iron (II)	Fe^{2+}	
Iron (III)	Fe^{3+}	
Calcium	Ca^{2+}	

Magnesium	Mg^{2+}	
Aluminium	Al^{3+}	

Q1. How do we know that the precipitates formed are insoluble?

Q2 Sometimes the copper precipitate is described as "gelatinous". What do you think this means?

Q3. Given time the colour of the Iron (II) precipitate changes. Why?

Q4. Three of the metal ions would be indistinguishable on first test. Which ones and why?

Q5. It would be possible to tell two of the three apart if we continue to add Sodium Hydroxide. How?

Q6. The colourful precipitates are all found in which region of the Periodic Table?

Anion Tests.

Halide ion tests - Ions from Group 7

- Pour about 2 cm of Chloride, Bromide and Iodide solutions into three separate test tubes.
- Add a few drops of Nitric acid.
- Then add a few drops of Silver Nitrate solution to each tube

This lesson involves ...

*Enquiring
*Creative/Enterprising
*Open-minded
*Risk-taking

☐ Safety: EYE-PROTECTION

Halide ion in solution	Formula of ion	Colour of ppt
Chloride	Cl^-	
Bromide	Br^-	
Iodide	I^-	

Q1. The positive silver ions bond to the negative halide ions as below:

$$\mathbf{Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \text{ --> } AgCl_{(s)}}$$ Silver Chloride ppt

a) Write a similar Ionic Equation for the Bromide Reaction

b) Write a similar Ionic Equation for the Iodide Reaction

Q2. Explain why it was necessary to add Nitric Acid at the start?

Q3. Hydrochloric Acid would have done the same job. Why would it have been a bad choice?

Sulphate ion Test (SO_4^{2-})

- Pour about 2 cm of Sulphate solution into a test tube.
- Add a few drops of Hydrochloric acid.
- Then add a few drops of Barium Chloride solution to each tube

☐ Safety: EYE-PROTECTION

Q1. What colour ppt do you observe?

Q2. Complete the Ionic Equation:

Ba^{2+} **(aq) +** SO_4^{2-} **(aq)-->** __________(s)   Barium Sulphate ppt

Q3. What evidence is there that Barium Sulphate is insoluble?

Q4. Explain why it was necessary to add Nitric Acid at the start?

Q5. Hydrochloric Acid would have done the same job. Why would it have been a bad choice?

This lesson involves ...

*Enquiring
*Creative/Enterprising
*Open-minded
*Risk-taking

Extended Response Question — Flame tests

Calcite is a mineral that contains calcium carbonate.

Arcanite is a mineral that contains potassium sulfate.

Samples of these two minerals have had their labels removed by mistake. These samples are currently labelled **A** and **B**.

Describe how samples of **A** and **B** could be chemically analysed in order to identify which sample is calcite and which sample is arcanite.

Your analysis should include valid results.

[6 marks]

Mark scheme

Answers	Mark
Level 3: The analysis would lead to the production of a valid outcome. The key steps are identified and logically sequenced.	5–6
Level 2: The analysis would not necessarily lead to a valid outcome. Most steps are identified, but the method is not fully logically sequenced.	3–4
Level 1: The analysis plan would not lead to a valid outcome. Some relevant steps are identified, but links are not made clear.	1–2
No relevant content	0

Indicative content

Chemical analysis

Testing for positive metal ion

- Add dilute hydrochloric acid to each sample.
- Use a nichrome wire loop **or** spray metal solution **or** dip a wooden splint in solution.
- Place the sample in a blue Bunsen flame.

Testing for negative ion

- Crush samples of **A** and **B** and place in separate test tubes.
- Add dilute hydrochloric acid to each sample.
- Add barium chloride (solution) to the acidified sample of **A** and of **B**.

Results

Testing for metal ion

- Calcium ions give an orange-red flame colour.
- Potassium ions give a lilac flame colour.

Testing for negative ion

- Fizzing after addition of dilute hydrochloric acid indicates the presence of carbonate ions.
- The gas given off from the addition of hydrochloric acid could be bubbled through limewater; the formation of a cloudy solution indicates carbon dioxide gas given off.
- The formation of a white precipitate after the addition of barium chloride indicates the presence of sulfate ions.

Identification

- Calcite will give an orange-red flame colour **or** will fizz with acid.
- Arcanite will give a lilac flame colour **or** will give a white precipitate/solid with barium chloride.
- Only testing of the metal ion or of the negative ion is needed to identify the mineral.

Tests for gases

Test for Hydrogen

Hydrogen is less dense than air and explosive when mixed with Oxygen.

- Place a few centimetres of hydrochloric acid in a test-tube in a rack.
- Drop in a strip of magnesium.
- Hold another inverted tube above the first to collect the Hydrogen made
- Light a splint and swiftly turn the test tube upwards a little towards the flame to allow the Hydrogen to mix with air.

Q1. What do you hear?

This lesson involves ...

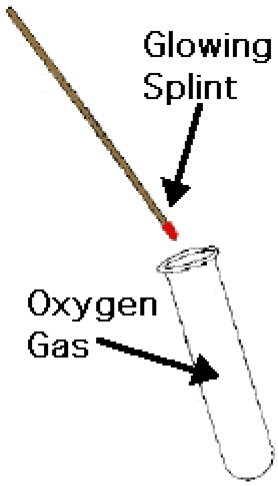
*Enquiring
*Creative/Enterprising
*Open-minded
*Risk-taking

Q2. Complete the equations below:

Hydrogen + Oxygen --> __________

$2H_{2(g)}$ + $O_{2(g)}$ --> __________

Test for Oxygen



- Make some Oxygen by pouring a centimetre of Hydrogen Peroxide solution into a test-tube.
- Put the tube in a rack.
- Add a small spatula of Manganese Oxide Catalyst.
- Place a glowing splint in the neck of the tube.

Q1. What do you observe?

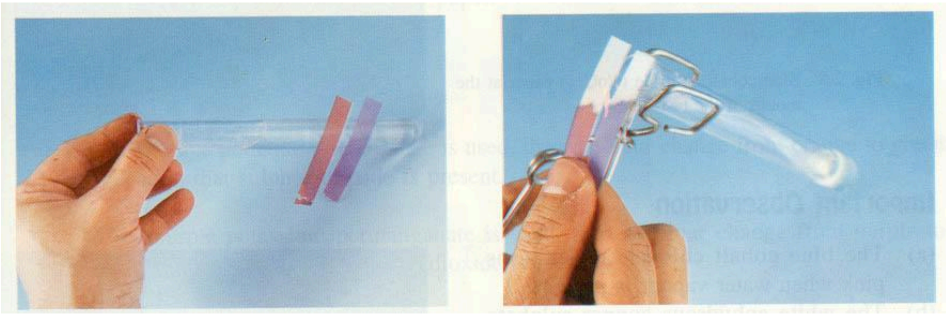
Q2. Why does this happen?

Test for Chlorine

Chlorine is a soluble, ***green*** gas that dissolves to form an ***acidic*** solution.

But it can look yellow in small amounts and it may be very difficult to see at all.

Your teacher will demonstrate holding damp blue Litmus paper in the Chlorine



Q1. What **two** colour changes do you see?

Q2. What causes the two colours?

Carbon Dioxide ion test

Carbonates react with dilute acids to produce Carbon Dioxide gas

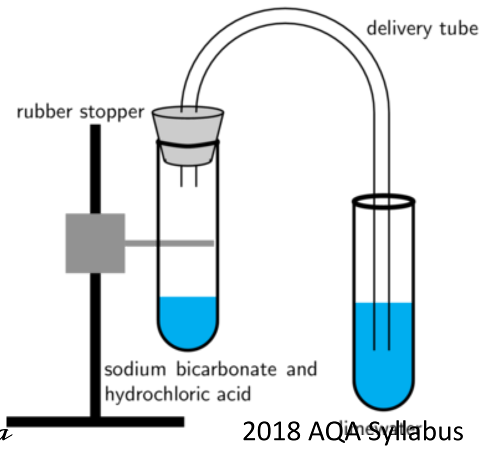
Carbonate + Acid --> Salt + Water + Carbon Dioxide

So adding acid to a substance and showing it makes CO_2 is a test to show the substance is a Carbonate.

We really should know the test for CO_2 from Year 7

- ☐ Pour some carbonate solution into one test-tube and limewater into another.
- ☐ Add a little hydrochloric acid to the carbonate and push in the bung

❖ SAFETY – EYE-PROTECTION



Q1. What should you have seen?

Q2. What happens if we continue to bubble Carbon Dioxide through the Limewater?

Extended Response Question

The proportions of different gases in the atmosphere

The gases present in the atmosphere can be tested and identified.

A sample of the air near a swimming pool was tested because there were complaints over the possibility of a chlorine gas leak.

.1 Describe how chlorine gas could be tested for, including the positive result. **[2 marks]**

.2 Describe a test for a gas, including the positive result, that you would expect to find in the sample of gas. **[2 marks]**

There was no chlorine gas found to be present in the sample. However, there was some sulfur dioxide gas found.

.3 Describe the problems caused by the presence of the acidic gas, sulfur dioxide, in the atmosphere. **[4 marks]**

Question	Answers	Mark
.1	**Test for chlorine** ● Use of damp litmus paper. ● Litmus paper is turned white/decolourised.	2
.2	**Gas expected in atmosphere** ● Oxygen ● Glowing splint relights OR ● Carbon dioxide ● Turns limewater cloudy OR ● Allow nitrogen gas (no simple test)	2
.3	**Level 2**: Scientifically relevant facts, events or processes are identified and given in detail to form an accurate account.	3–4
	Level 1: Facts, events or processes are identified and simply stated but their relevance is not clear.	1–2
	No relevant content	0
	Indicative content **Sulfur dioxide problems** ● Sulfur dioxide gas causes respiratory problems / affects asthma suffers. ● Sulfur dioxide dissolves to form acid rain or forms sulfuric acid. **Acid rain problems** ● Acid dissolves **or** corrodes limestone buildings. ● Acid reacts with **or** corrodes metals. ● Damages (waxy layer on) plants – allow: kills plants. ● Makes it difficult for plants to absorb nutrients. ● Makes lakes/ponds too acidic for fish/aquatic life.	

11 A student carried out a series of tests on a solid, **M**, in order to identify the ions that could be present.

The table shows her results.

Test	Method	Result
Test 1	Carry out a flame test on solid **M**	Lilac flame
Test 2	Dissolve solid **M** in water, and divide the solution into three portions, A, B and C	

ous

(b) Describe how the student should carry out a flame test on solid **M**.

(3)

....

....

....

....

....

....

....

....

(c) (i) Why was dilute nitric acid added to the solution of solid **M** before using silver nitrate solution?

(1)

....

(ii) Why should dilute hydrochloric acid **not** be used in place of dilute nitric acid in this test?

(2)

....

....

....

(d) The tests for negative ions that the student carried out involved precipitation.

Suggest **one** negative ion that cannot be identified by a precipitation reaction.

(1)

....

(Total for Question 11 = 10 marks)

8 A student was asked by his teacher to perform a flame test on a solid.

He used this method.

- dip the tip of a clean platinum wire into hydrochloric acid and then into the solid
- adjust the air hole of the Bunsen burner to obtain a non-roaring, non-luminous Bunsen flame
- place the tip of the platinum wire into the edge of the flame
- observe the colour in the flame

(a) (i) Why is it important that the platinum wire is clean?

(1)

....

....

(ii) Why is it important to use a non-luminous flame?

(1)

....

....

(iii) What colour would be observed in the flame if the solid contained sodium ions?

(1)

....

(b) Another student was given a pale violet solid. He was told that it contained two cations (positive ions) from this list

Li^{+} Ca^{2+} Fe^{2+} NH_4^{+} Fe^{3+}

He performed a flame test on the solid.

He then dissolved a small sample of the solid in water. A yellow solution was formed.

He added sodium hydroxide solution and then warmed the mixture.

The table shows his observations.

Test	Observation
flame test	no positive result
add sodium hydroxide solution and warm	brown precipitate a pungent-smelling gas was evolved the gas turned damp red litmus paper blue

(i) The flame test gave no positive result.

State the two cations from the list that are **not** present in the solid.

(1)

.... and ....

(ii) Identify the pungent-smelling gas given off and explain why the red litmus paper must be damp before it is used.

(2)

....

....

....

....

(iii) Identify the two cations present in the pale violet solid.

(2)

.... and ....

(d) Sodium chloride can be made by many different reactions.

A student prepared a sample of sodium chloride using the following method.

Step 1 She added an excess of a solid sodium compound, X, to dilute hydrochloric acid. The mixture fizzed as the solid reacted.

Step 2 She filtered the mixture produced to remove the excess solid X. The filtrate was a colourless liquid.

Step 3 She evaporated the colourless liquid. A white solid remained.

(i) Describe a chemical test that the student could do to show that the colourless liquid in Step 2 contained chloride ions, Cl^-.

(3)

Test ....

....

....

Result ....

(ii) The student concluded that solid X was sodium hydroxide.

State one reason why this conclusion was **not** correct.

Suggest a possible identity of solid X.

(2)

Reason ....

....

Solid X could be ....

3 A student added some pieces of iron to a boiling tube containing dilute hydrochloric acid. She observed fizzing and the formation of a solution, X.

(a) Identify the gas that causes the fizzing and describe a test for it.

(2)

Gas ....

Test ....

....

(b) Solution X contains chloride ions.

(i) The student confirmed this by adding some silver nitrate solution. She observed a white precipitate of silver chloride.

Give the formula of the white precipitate, and name the other solution she should have added before the silver nitrate solution.

(2)

Formula of white precipitate ....

Other solution ....

(ii) Complete the word equation for the reaction in this test.

(1)

iron chloride + silver nitrate → silver chloride + ...

(c) Solution X also contains ions of iron. The student thought that these ions had the formula Fe^{2+} or Fe^{3+}.

What reagent should she add to decide whether solution X contains Fe^{2+} or Fe^{3+} ions? State the result of the test in each case.

(3)

Reagent ....

Result with Fe^{2+} ions ....

....

Result with Fe^{3+} ions ....

....

(Total for Question 3 = 8 marks)

(iii) The presence of chloride ions in HCl(aq) can be shown by mixing it with silver nitrate solution and dilute nitric acid.

State the result of this test and complete the chemical equation for the reaction by adding the state symbols.

(3)

result ....

....

$AgNO_3$(..............) + HCl(aq) → AgCl(..............) + HNO_3(..............)

(b) Grape vines can be attacked by a fungus that ruins the grapes. The fungus can be killed using Bordeaux mixture, a solid containing copper(II) sulfate and calcium hydroxide.

(i) State a test to show that Bordeaux mixture contains calcium ions.

(2)

test for calcium ions ....

observation ....

(ii) A sample of Bordeaux mixture is dissolved in water.

Describe separate tests to show that this solution contains copper(II) ions and sulfate ions.

(5)

test for copper(II) ions ....

....

observation ....

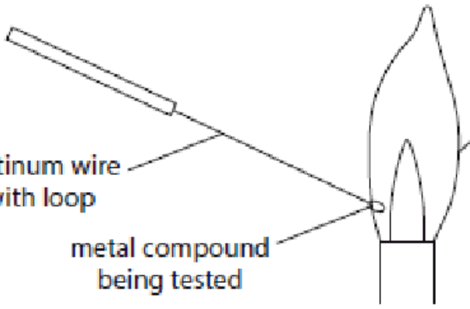
test for sulfate ions ....

....

observation ....

8 A flame test is carried out on three metal compounds, X, Y and Z.

The diagram shows the apparatus used.



(a) (i) Suggest two reasons why platinum is a suitable metal to use as the wire in this test.

(2)

1 ....

....

2 ....

....

(ii) Why should the platinum wire be cleaned between each test?

(1)

....

....

(iii) Why is a luminous Bunsen flame not suitable for carrying out a flame test?

(1)

....

....

(b) The three metal compounds are also tested separately with three reagents.

The reagents used are

- aqueous acidified silver nitrate
- aqueous acidified barium chloride
- aqueous sodium hydroxide

The table shows the results of all the tests.

Metal compound	Flame test	Aqueous acidified silver nitrate	Aqueous acidified barium chloride	Aqueous sodium hydroxide
X	yellow	white precipitate	no precipitate	no precipitate
Y	red	no precipitate	white precipitate	no precipitate
Z	no colour	no precipitate	no precipitate	green precipitate

(i) Give the name of compound X and of compound Y.

(4)

compound X ....

compound Y ....

(ii) Identify the cation present in compound Z.

(1)

....

(c) Aqueous sodium hydroxide can be used to distinguish between solutions containing iron(II) ions (Fe^{2+}) and iron(III) ions (Fe^{3+}).

State the observation made when aqueous sodium hydroxide is added separately to each solution.

(2)

Fe^{2+}(aq)...

Fe^{3+}(aq)...

8 Three aqueous solutions are sodium chloride, sodium iodide and silver nitrate. They are in containers labelled X, Y and Z. It is not known which solution is in each container.

The solutions are mixed together as shown in the table, and the observations recorded.

Experiment	Observation
solution X added to solution Y	yellow precipitate formed
solution X added to solution Z	no change

(d) A student uses a flame test to distinguish between separate samples of calcium chloride and potassium chloride.

This is the student's method.

There is one mistake in step 1 and one mistake in step 3.

step 1 dip a platinum wire into some concentrated sodium hydroxide solution

step 2 dip the platinum wire into the sample

step 3 place the wire and sample into a luminous Bunsen flame

step 4 record the colour of the flame

Describe a correct method for step 1 and step 3.

(2)

step 1 ....

....

step 3 ....

....

(e) What colour is the flame when the test on potassium chloride is carried out correctly?

(1)

☐ **A** green

☐ **B** lilac

☐ **C** orange

☐ **D** red

4 Chemical tests can be used to detect ions in solids and in aqueous solutions.

(b) When dilute nitric acid is added to an aqueous solution, followed by silver nitrate solution, a yellow precipitate forms.

Which of these halide ions is present in the aqueous solution?

(1)

☒ **A** Br^-

☒ **B** Cl^-

☒ **C** F^-

☒ **D** I^-

(c) When dilute hydrochloric acid is added to a solid, a gas forms.

Which of these ions is present in the solid?

(1)

☒ **A** carbonate

☒ **B** hydroxide

☒ **C** nitrate

☒ **D** sulfate

(d) Sodium hydroxide solution is added separately to three solutions.

One solution contains Cu^{2+} ions, another contains Fe^{2+} ions and the third solution contains Fe^{3+} ions.

Which row shows the correct colours of the precipitates that form?

(1)

	Cu^{2+}	Fe^{2+}	Fe^{3+}
☒ **A**	green	blue	brown
☒ **B**	brown	green	blue
☒ **C**	blue	green	brown
☒ **D**	blue	brown	green

(e) When barium chloride solution is added to an aqueous solution of a compound, a white precipitate forms. When dilute hydrochloric acid is added to the mixture, the precipitate disappears and a colourless solution forms.

Which of these ions is present in the aqueous solution?

(1)

☒ **A** carbonate

☒ **B** chloride

☒ **C** nitrate

☒ **D** sulfate

(d) These tests are carried out on two separate samples of iron(III) sulfate solution.

test 1 add sodium hydroxide solution

test 2 add dilute hydrochloric acid, then add barium chloride solution

(i) Which observation is correct for test 1?

(1)

☐ **A** brown precipitate

☐ **B** brown solution

☐ **C** green precipitate

☐ **D** green solution

(ii) Give the names of the two products formed in test 1.

(2)

.... and ....

(iii) In test 2, there is no visible change after adding dilute hydrochloric acid.

State why the acid is added.

(1)

....

....

(iv) In test 2, barium sulfate is formed after adding barium chloride solution.

State the observation that is made.

(1)

....

....

(b) Copper and iron both form compounds containing ions with a 2+ charge.

Sodium hydroxide is used to distinguish between solutions containing Cu^{2+} and Fe^{2+} ions.

What would be observed when aqueous sodium hydroxide is added separately to each solution?

(2)

Cu^{2+} ....

Fe^{2+} ....

6 Some iron(II) sulfate ($FeSO_4$) is dissolved in water to make a solution.

(a) A reaction takes place when sodium hydroxide solution is added to a solution of iron(II) sulfate.

(i) Complete the word equation to show this reaction.

(2)

sodium hydroxide + iron(II) sulfate ⟶ .... + ...

(ii) State what you would observe in this reaction.

(1)

....

(b) Barium chloride is used to test for sulfate ions.

(i) Barium chloride solution is added to another solution of iron(II) sulfate. A white precipitate forms.

Identify the white precipitate.

(1)

....

(ii) In this test, another substance should be added to react with any carbonate ions that might be present.

Identify this other substance and state one observation that would be made if carbonate ions were present.

(2)

Other substance

....

Observation

....

10.2 - Energetics

I have reviewed the syllabus statements for this topic	
I have completed the questions in this section	

I have read the relevant sections of the College Website	
I have made some revision material (mind-map, key-words & definitions etc)	
Prep Grade	
Test Grade	

Target	Pupil Signature

10.2 - Energetics

- ***Energy is conserved*** in chemical reactions. The amount of energy in the universe at the end of a chemical reaction is the same as before the reaction takes place.
- If a reaction transfers energy to the surroundings the product molecules must have less energy than the reactants, by the amount transferred.
- ***Exothermic*** reactions transfer energy to the surroundings and ***raise*** its ***temperature***.
- ***Exothermic*** reactions include ***combustion***, many ***oxidation*** reactions and ***neutralisation***.
- Everyday uses of exothermic reactions include ***self-heating cans*** and ***hand warmers***.
- E***ndothermic*** reactions take in energy from the surroundings & ***lower*** its ***temperature*** eg ***thermal decomposition,*** & Citric Acid with Sodium Hydrogencarbonate.
- Some ***sports injury packs*** are based on endothermic reactions.
- Students should be able to distinguish between exothermic and endothermic reactions on the basis of the temperature change of the surroundings and evaluate uses of exothermic and endothermic reactions given appropriate information.
- Limited to measurement of temperature change. ***Calculation of (ΔH) is not required***.

- Chemical reactions occur only when reacting particles ***collide with sufficient energy.***
- The ***minimum*** amount of energy that particles must have to react is called the ***Activation Energy.***
- Reaction profiles can be used to show the ***relative energies of reactants and products***, the ***Activation Energy*** and the ***overall energy change*** of a reaction.
- Students should be able to 1 ***draw*** simple reaction profiles (energy level diagrams) for exothermic and endothermic reactions showing the relative energies of reactants and products, the activation energy and the overall energy change, ***with a curved line to show the energy as the reaction proceeds. 2 use*** reaction profiles to ***identify*** reactions as exothermic or endothermic 3 ***explain*** that the activation energy is the energy needed for a reaction to occur.
- During a chemical reaction: 1energy must be ***supplied to break bonds in the reactants*** 2 energy is ***released*** when ***bonds in the products are formed***.
- Energies needed to break bonds & released when formed can be calculated from ***bond energies***.
- The ***difference*** between the sum of energies needed to break bonds in reactants and sum of the energies released when bonds in the products form is the overall energy change of the reaction.
- In an ***exothermic reaction***, the energy released from forming new bonds is ***greater*** than the energy needed to break existing bonds.
- In an ***endothermic reaction***, the energy needed to break existing bonds is ***greater*** than the energy released from forming new bonds.
- Students should be able to calculate the energy transferred in reactions using bond energies.

Cells & Fuel Cells

- ***Cells*** contain chemicals which react to produce electricity.
- The ***voltage*** produced by a cell is dependent upon a number of factors including the type of ***electrode*** and ***electrolyte***
- A simple cell can be made by connecting ***two different metals in contact with an electrolyte.***
- ***Batteries*** consist of two or more cells connected together in series to provide a greater voltage.
- ***Alkaline*** batteries are non-rechargeable. In non-rechargeable cells and batteries the chemical reactions stop when one of the reactants has been used up.
- ***Rechargeable*** cells and batteries can be recharged because the ***chemical reactions are reversed*** when an external electrical current is supplied.
- Students should be able to interpret data for reactivity of metals & evaluate the use of cells.
- Students do not need to know details of cells and batteries other than those specified.
- ***Fuel cells*** are supplied by an external source of fuel (eg Hydrogen) and Oxygen or air.

- The fuel is ***oxidised*** electrochemically within the fuel cell to produce a ***potential difference***.
- The overall reaction in a Hydrogen fuel cell involves the oxidation of Hydrogen to produce water.
- Hydrogen fuel cells offer a potential alternative to rechargeable cells and batteries.
- Students should be able to:

1. evaluate the use of Hydrogen fuel cells in comparison with rechargeable cells and batteries
2. (HT only) write the half equations for the electrode reactions in the Hydrogen fuel cell.

Exo- and Endothermic Reactions

Some reactions ***give out heat*** – making the ***temperature rise*** and their container feel ***warmer***.

They are ***exothermic*** – ***Thermic*** is to do with heat, ***Exo*** means outwards (you go out through an ***ex***it)

Other reactions ***take in heat*** – making the ***temperature fall*** their container feel ***colder***.

They are ***endothermic*** –, ***Endo*** means inwards (you go in through an ***en***trance).

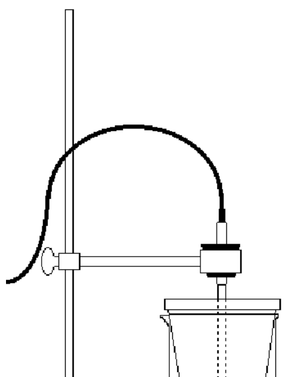
It's difficult to measure heat energy – but easy to measure temperature changes.

This lesson involves…

*Enquiring
*Creative/Enterprising
*Open-minded
*Risk-taking

Experiment 1.

- Measure 25 cm^3 of dilute Hydrochloric acid.
- Pour the acid into a plastic cup.



- Place the cup in a beaker and cover with a lid.
- Place a thermometer into the acid, through the lid and wait for the temperature to stop changing.
- Record the Start Temperature
- Add 3cm of Magnesium strip
- Record the highest temperature reached
- Repeat the experiment with fresh acid

Results

	Start Temperature (°C)	**Highest Temperature (°C)**	**Temperature Change (°C)**
Experiment 1			
Experiment 2			

Average Temperature Change = ______________

Q1. Were your two results similar? Was your temperature change similar to your neighbours'? If not suggest why not?

Q2. Hand-warmers are an example of an __________-thermic reaction

Cool-packs are an example of an __________-thermic reaction

Q3. It is impossible to do this experiment without some heat losses. Where does the heat go?

This lesson involves ...

*Enquiring
*Creative/Enterprising
*Open-minded
*Risk-taking

Experiment 2.

- Measure 25 cm^3 of dilute Citric acid.
- Pour the acid into a plastic cup
- Place the cup in a beaker and cover with a lid.
- Place a thermometer into the acid, through the lid and wait for the temperature to stop changing
- Add 3 spatulas of Sodium Hydrogen Carbonate and record the lowest temperature reached
- Repeat the experiment with fresh acid

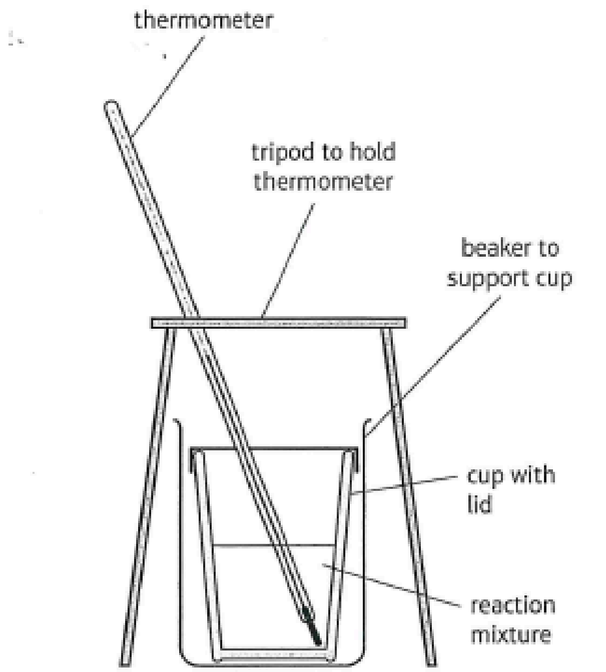
Results

	Start Temperature (°C)	Lowest Temperature (°C)	Temperature Change (°C)
Expt 1			
Expt 2			

Q1. Is the reaction Exo or Endothermic?

Chemistry

Practical 4: Temperature changes



Safety

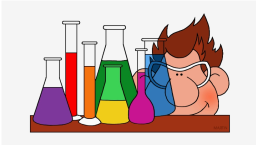
- Wear safety goggles.
- Sodium hydroxide and hydrochloric acid at this concentration are corrosive and very damaging to eyes.

Method

A Use a measuring cylinder to put 30 cm^3 of 2.0 mol/dm^3 hydrochloric acid into a polystyrene cup.

B Clean out the measuring cylinder with water.

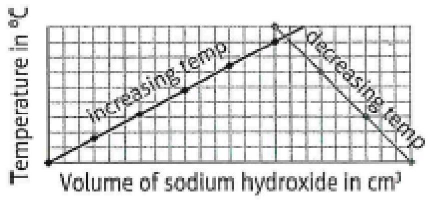
C Place the polystyrene cup into a glass beaker to make it more stable. Measure the



Volume of sodium hydroxide solution added in cm³	Maximum temperature recorded at each stage in °C				
	Experiment 1	Experiment 2			Mean
0					
5					
10					
15					
20					
25					
30					
35					
40					

2 On the next page, draw a graph of your results with the volume of sodium hydroxide solution added in cm³ (the independent variable) on the *x*-axis and the mean maximum temperature in °C (the dependent variable) on the *y*-axis.

3 Draw two straight lines of best fit. One will be through all the points which are increasing in temperature and the other will be through all the points which are decreasing in temperature. You need to make sure the lines cross, so you might need to extend them. The diagram on the right should give you an idea of what this should look like.



4 What volumes of hydrochloric acid and sodium hydroxide would produce the largest temperature rise?

____

____

5 Why does the temperature start to fall towards the end of the experiment?

____

____

Evaluation

6 How could you make the data you collect from this investigation more accurate?

____

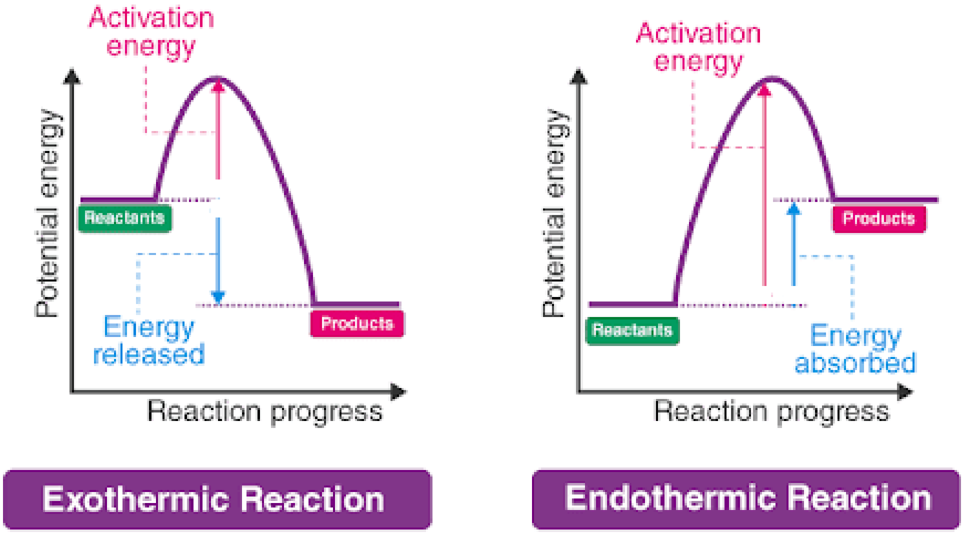
____

7 It is difficult to find the exact volume of sodium hydroxide solution that would give the maximum temperature rise. What further work would you need to carry out to find this exact volume of sodium hydroxide?

____

____

Energy Profiles



Whether a reaction is Exo or Endothermic some energy must be put in before the reaction can start.

We call this the **ACTIVATION ENERGY Ea.**

The **energy released/absorbed** by in the reaction (Δ**H**) is the difference in energy between the reactants and the products.

Q1 How do we provide the activation energy to allow matches to burn?

Q2 Lighting a charcoal barbeque usually requires a paraffin-based firelighter to “get it going”. Does the Fire-lighter or the charcoal have the highest Activation energy?

Q3 What does the Δ symbol in the energy change ΔH indicate in Maths?

Q4 In an exothermic reaction the ΔH points **up/down,** this means that the value for ΔH will be **positive/negative** - delete as appropriate.

Q5 In an enoothermic reaction the ΔH points **up/down,** this means that the value for ΔH will be **positive/negative** - delete as appropriate.

18 Draw and compare the reaction profiles of exothermic and endothermic reactions.

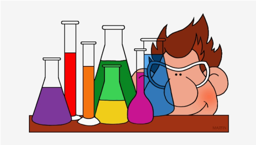
[6 marks]

Question	Answers	Mark
18	**Level 2:** Scientifically relevant features are identified; the way(s) in which they are similar/different is made clear and (where appropriate) the magnitude of the similarity/difference is noted.	4–6
	Level 1: Relevant features are identified and differences noted.	1–3
	No relevant content	0
	Indicative content ● Axes drawn and labelled energy (y-axis) and progress of reaction (x-axis). ● Horizontal lines drawn and labelled reactants and products. ● Products line to the right of reactants. ● Arrows drawn from reactants to products. ● Labelled ΔH. ● Curves drawn. ● Activation energies labelled. ● Exothermic reaction has products line at a lower energy than reactants line ● But endothermic reaction has products line at a higher energy than reactants line. ● ΔH arrow is pointing downwards for the exothermic reaction ● Because energy is transferred to the surroundings / given out ● But pointing upwards for the endothermic reaction ● Because energy is taken in. ● Both have activation energies from the reactants line to the top of the curve. ● The activation energy for the endothermic reaction is greater than that for the exothermic reaction. Credit can be given for all points on labelled diagrams.	

Bonds and Energy.

It takes energy to ***break*** bonds

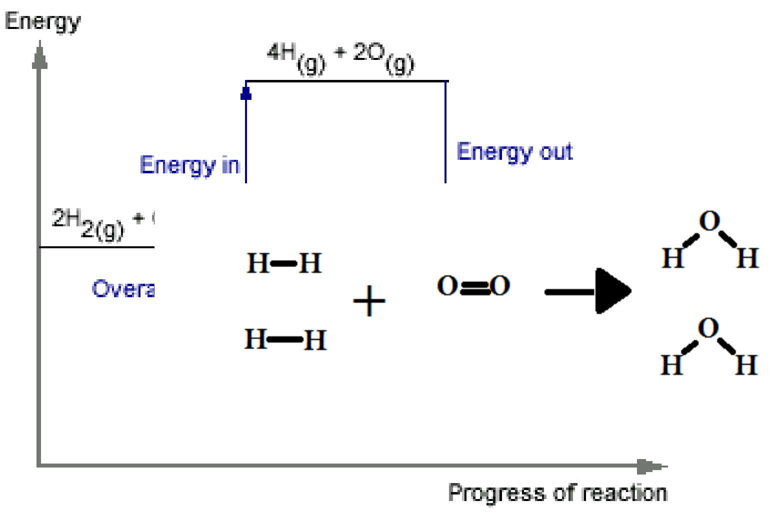
But ***making*** new bonds releases energy.



In a reaction bonds in molecules of reactants must break ***before*** new bonds can form in the products.

So all reactions start off with an input of energy

Hydrogen reacts with Oxygen to make water



If we show the bonds in this reaction:

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

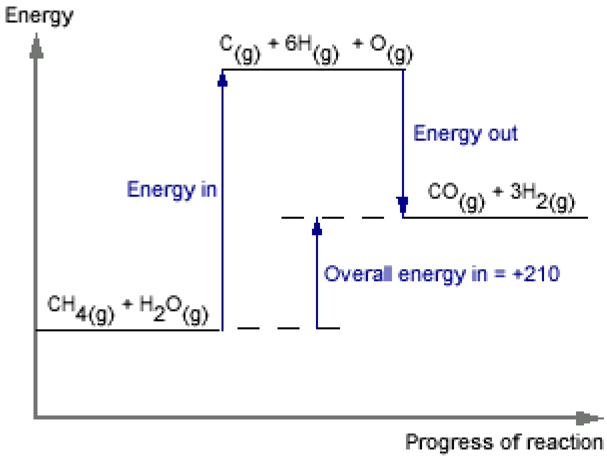
It’s clear we’d have to put in enough energy t break two H-H bonds, and one O=O bond before the reaction can happen.

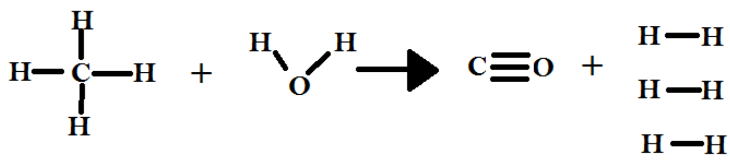
But energy is released when four O-H bonds are made.

The energy released is ***more than*** the energy needed at the start – so this is an ***Exothermic*** reaction.

Question.

The reaction below converts Methane (CH_4) and Water (H_2O) to Carbon Monoxide (CO) and Hydrogen (H_2)





Q1 Is this an Exo or an Endo-thermic reaction? How can you tell?

Q2. Which bonds need to be broken before this reaction can happen?

Q3. Which bonds are formed during this reaction?

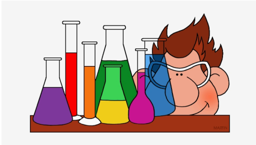
This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

Q4. Why is less energy released than was needed at the start?

Q5. What would happen to the temperature during this reaction?

Bond Energies



Chemists can calculate the energy needed to ***break*** particular bonds.

Not every bond takes as much energy to break.

Generally, ***shorter*** bonds are the ***strongest***.

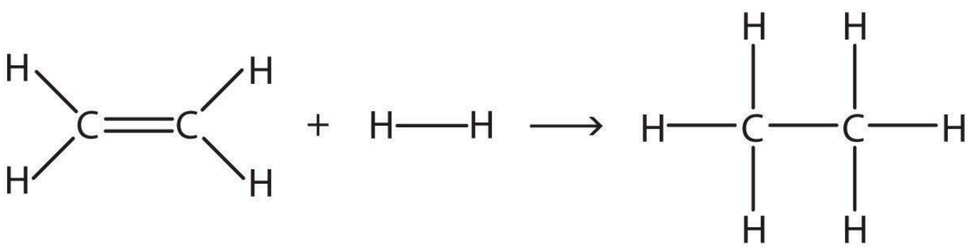
Notice the units are ***kJ per mole –*** the energy needed to break a mole of bonds - because breaking a single bond would require so little energy it would be difficult to measure.

We can use these Bond Enthalpies to work out how much energy a reaction will release.

Bond Lengths and Bond Energies

	Bond Length (nm)	Bond Energy (kJ/mol)
H–H	0.074	435
H–Cl	0.127	431
Cl–Cl	0.198	243
H–C	0.109	414
C–Cl	0.177	328
C–C	0.154	331
C=C	0.134	590
C≡C	0.120	812
C–O	0.143	326
C=O	0.120	803
C≡O	0.113	1075
N–N	0.145	159
N=N	0.125	473
N≡N	0.110	941

Example.



In this reaction, the ***energy input*** (energy needed to break bonds) would be:

1(C=C) + 4(C-H) + 1(H-H) = 590 + (4 x 414) + 435 = 2681 kJ/mol

The ***energy output*** (energy released by new bonds forming) would be:

1(C-C) + 6(C-H) = 331 + (6 x 414) = 2815 kJ/mol

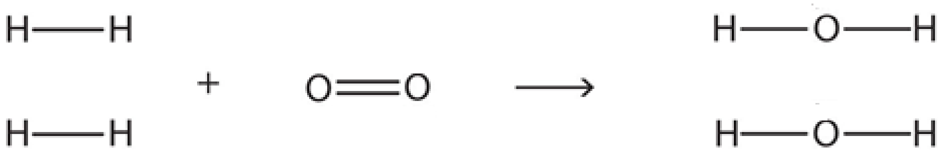
So the **Enthalpy change (ΔH)** for this reaction would be: **Output – Input** = 2681 – 2815 = -134

This is an exothermic reaction. A negative **ΔH** means less energy was put in than was released.

Questions

Use the bond energies on the previous page to answer the following questions. You will also need the following bond energies : (O-H) = 463, (O=O) = 496

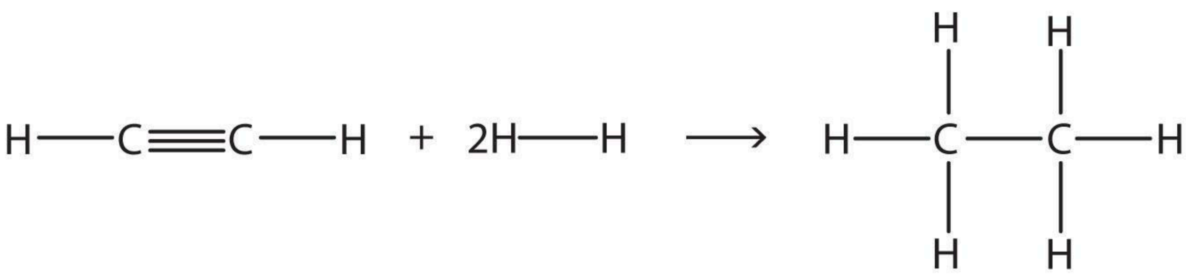
Q1. Calculate the enthalpy change for this reaction: $2H_2 + O_2 \rightarrow 2H_2O$



This lesson involves …

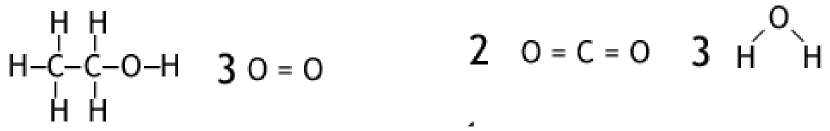
*Logical Thinking
*Precision
*Multi-step problem-solving

Q2. Calculate the enthalpy change for this reaction: $C_2H_2 + 2H_2 \rightarrow C_2H_4$

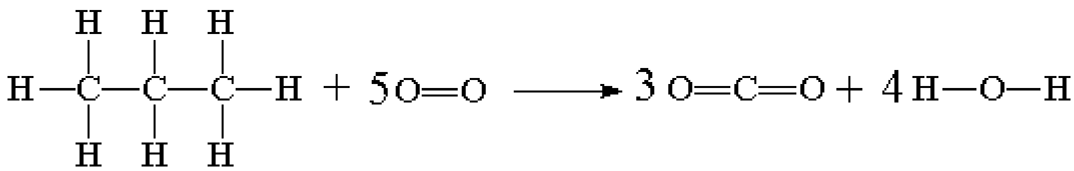


Q3. Calculate the enthalpy change in the following reaction

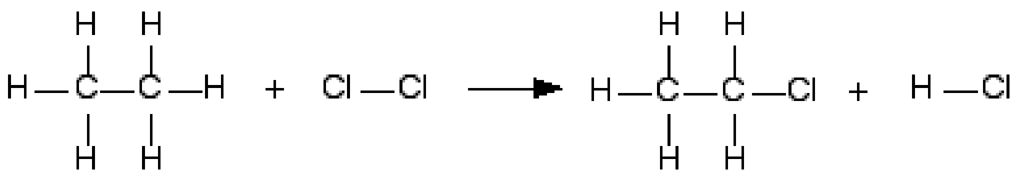
$$C_2H_5OH\ (g) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O\ (g)$$



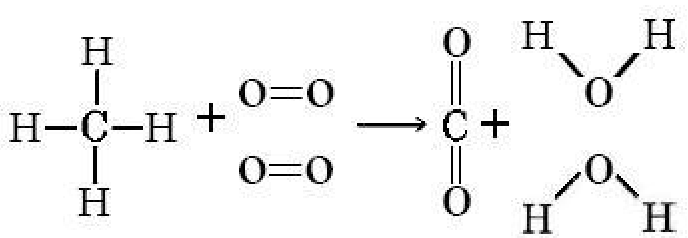
Q4. Calculate the enthalpy change in the following reaction: **$C_3H_{11} + 5O_2 \rightarrow 3\ CO_2 + 4H_2O$**



Q5. Calculate the enthalpy change in the following reaction: **$C_2H_6 + Cl_2$ □ C_2H_4Cl+ HCl**



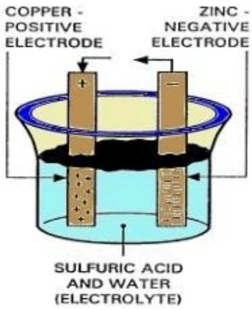
Q6. Calculate the enthalpy change in the following reaction: **$CH_4 + 2O_2$ □ $CO_2 + 2H_2O$**



What is a cell?

When chemicals reacts there is always an energy change - usually energy is released – this is **an E___thermic reaction**

But the energy doesn't have to be released as heat, it could be electrical energy.



In the cell above there are two electrodes of different metals separated by an electrolyte - a conducting solution, often an acid.

- WHY MUST THE SOLUTION BE IONIC, NOT COVALENT?

- All metals try to react by ____________electrons

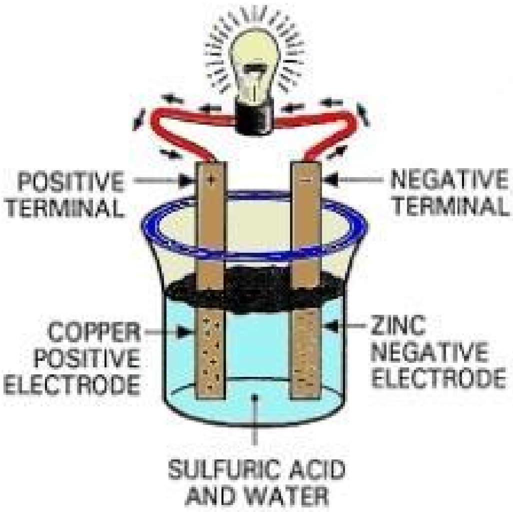
- If there are two metals than the **most/least** reactive metal will lose electrons and the **most/least** reactive will gain them.
- ***Zinc is more reactive than Copper.*** So it loses electrons and they flow **to/from** the Copper.

- _______ move through the electrolyte to complete the circuit.

This is an electrical cell.

EXPERIMENT

- Pour 50 cm^3 of Sulphuric acid in a beaker.
- Connect a zinc and a copper electrode via voltmeter.
- Measure the voltage ________v.
- How can we tell which is the positive and which is the negative terminal?



What affects the voltage?

A cell like this will generate a small voltage, which can be changed by:

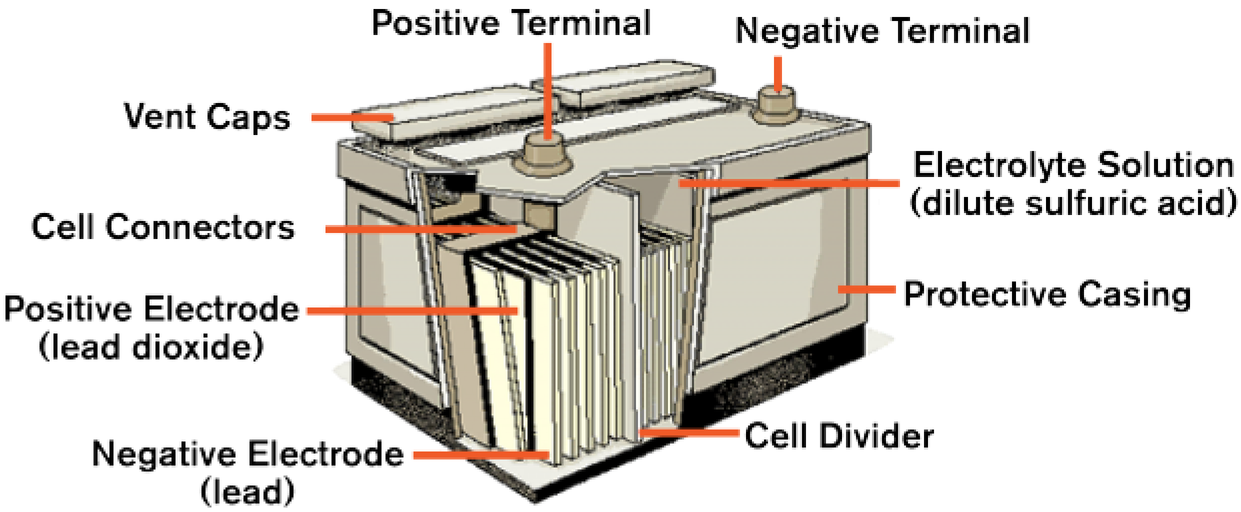
1. Using a different electrolyte
2. Using different electrodes

We ***don't*** have to predict anything about the electrolyte but we should know that ***the greater the difference in Reactivity of the metal electrodes the higher the voltage.***

- **Swap the Zinc electrode for a small strip of Magnesium. What is the voltage now?**

What is a battery?

So often ***multiple cells*** are connected in a ***battery*** - such as the ones used in cars.



Car batteries are rechargeable – What does this mean?

____

____

But the familiar ***Alkaline batteries*** are not rechargeable - they will eventually stop working.

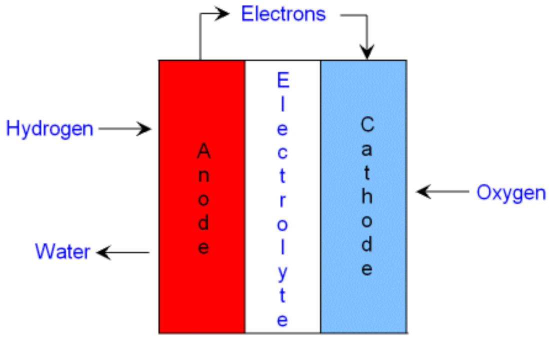
Why?

____

____

Fuel Cells

A ***fuel cell*** simply continually re-fuels the cell so that it carries on generating a voltage without ever running out of chemicals (providing we don't run out of fuel)



❖ Why not just burn the Hydrogen and generate electricity from the heat?

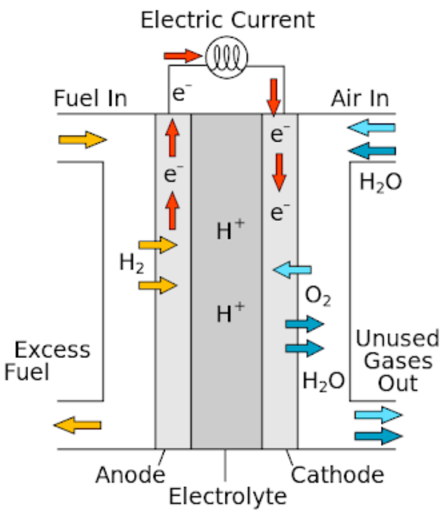
____

____

In the fuel cell the Hydrogen and Oxygen ***never meet*** - they are separated by the ***electrolyte***.

This lesson involves …

*Generalisation
*Connection-finding
*Big Picture Thinking
*Abstraction
*Imagination



❖ How do they react if they never meet?

Half Equations

Balance the half equations using e^- to represent electrons

Half-eqn 1: $H_{2(g)}$ --> ___$H^+_{(aq)}$ + ___ e^-

The electrons move away along a wire (they are the electric current); the Hydrogen ions pass through the electrolyte to the other side of the cell where they can react with the Oxygen.

The simplest half-equation for this is

Half-eqn 2: ___$H^+_{(aq)}$ + $O_{2(g)}$ + ___ e^- --> ___ $H_2O_{(g)}$

So, only water vapour is produced.

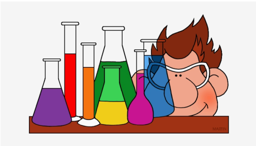
You may notice that the number of electrons in the two half-equations doesn't match.

In reality all the electrons made & used must match so we multiple **Half-eqn 1** by 2

2 x Half-eqn 1: ___$H_{2(g)}$ --> ___$H^+_{(aq)}$ + ___ e^-

Add the two half-equations together and cancelling what appears in both.

Overall : ________ + ______ --> ________



Advantages and Disadvantages of Fuels Cells

Advantages

This lesson involves …

*Collaborative
*Concerned for Society
*Confident

❖ What are the advantages of fuel cells?

❖ What are the disadvantages of fuel cells?

Extended Response Question 2018

Hydrogen fuel cells and rechargeable lithium-ion batteries can be used to power electric cars.

0 3 . 4 Complete the balanced equation for the overall reaction in a hydrogen fuel cell.

[2 marks]

$$\_\_\_\_\_\_\_\_ H_2 + \_\_\_\_\_\_\_\_\_\_ \rightarrow \_\_\_\_\_\_\_\_\_ H_2O$$

0 3 . 5 **Table 1** shows data about different ways to power electric cars.

Table 1

	Hydrogen fuel cell	**Rechargeable lithium-ion battery**
Time taken to refuel or recharge in minutes	5	30
Distance travelled before refuelling or recharging in miles	Up to 415	Up to 240
Distance travelled per unit of energy in km	22	66
Cost of refuelling or recharging in £	50	3
Minimum cost of car in £	60 000	18 000

Evaluate the use of hydrogen fuel cells compared with rechargeable lithium-ion batteries to power electric cars.

Use **Table 1** and your own knowledge.

[6 marks]

____

____

____

____

____

Mark Scheme

03.5	**Level 3:** A judgement, strongly linked and logically supported by a sufficient range of correct reasons, is given.	5–6
	Level 2: Some logically linked reasons are given. There may also be a simple judgement.	3–4
	Level 1: Relevant points are made. This is not logically linked.	1–2
	No relevant content	0

Indicative content

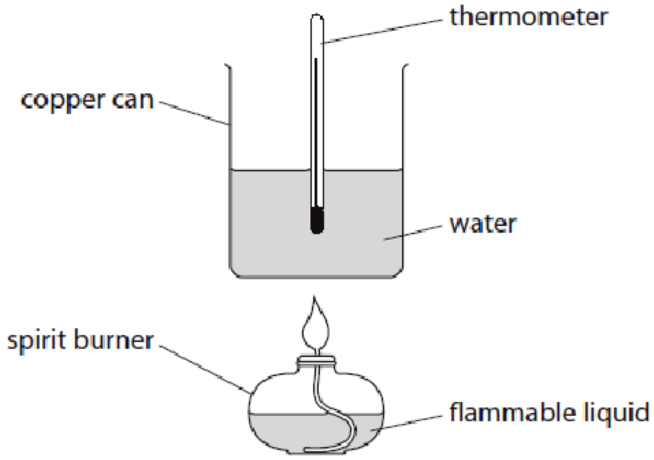
reasons why fuel cells could be judged as better

from the table	from other knowledge
• time for refuelling a fuel cell is faster than recharging or a fuel cell does not need to be recharged • a fuel cell has a greater range	• hydrogen can be renewable if made by electrolysis using renewable energy • lithium-ion batteries can catch fire • produces only water or no pollutants produced • lithium-ion batteries may release toxic chemicals on disposal • lithium-ion batteries (eventually cannot be recharged so) have a finite life

reasons why the lithium-ion battery could be judged as better

from the table	from other knowledge
• lithium-ion uses energy more efficiently • cost of lithium-ion car much less • cost of recharging much less than refuelling with hydrogen	• hydrogen is often made from fossil fuels so is not renewable • charging points are more widely available than hydrogen filling stations • hydrogen takes up a lot of space or is difficult to store • hydrogen can be highly flammable / explosive • no emissions produced • (catalyst in the hydrogen fuel-cell eventually becomes poisoned so) have a finite life

7 A student investigates the temperature rise of water in a copper can placed above a spirit burner containing a flammable liquid. The diagram shows the apparatus he uses.



This is the student's method.

- place 200 g of water in the copper can and record the temperature of the water
- weigh the spirit burner containing the flammable liquid
- place the spirit burner underneath the copper can and light the burner
- after two minutes extinguish the flame and record the maximum temperature of the water
- reweigh the spirit burner containing the remaining flammable liquid

(a) State whether each of the changes listed in the table would increase, decrease or have no effect on the value of the maximum temperature of the water.

(3)

Change	Effect on the value of the maximum temperature of the water
increasing the distance between the spirit burner and the copper can	
using a thermometer with divisions at 0.2 °C instead of 0.5 °C	
adding insulation to the side of the copper can	

9 A student investigates the reaction of aqueous sodium hydroxide with two different aqueous solutions of hydrochloric acid, solution X and solution Y.

She carries out two experiments.

Experiment 1

- Using a measuring cylinder, she pours 20 cm^3 of aqueous sodium hydroxide into a conical flask and records its temperature.
- Using a burette, she adds 5 cm^3 of solution X to the conical flask.
- She stirs the mixture with the thermometer and records the temperature.
- She adds further 5 cm^3 volumes of solution X and stirs with the thermometer.
- She records the temperature after each addition of solution X.
- She stops when a total of 40 cm^3 of solution X has been added.

Experiment 2

- She empties the burette and rinses it first with water and then with solution Y. She then fills the burette with solution Y.
- She repeats the experiment using solution Y.

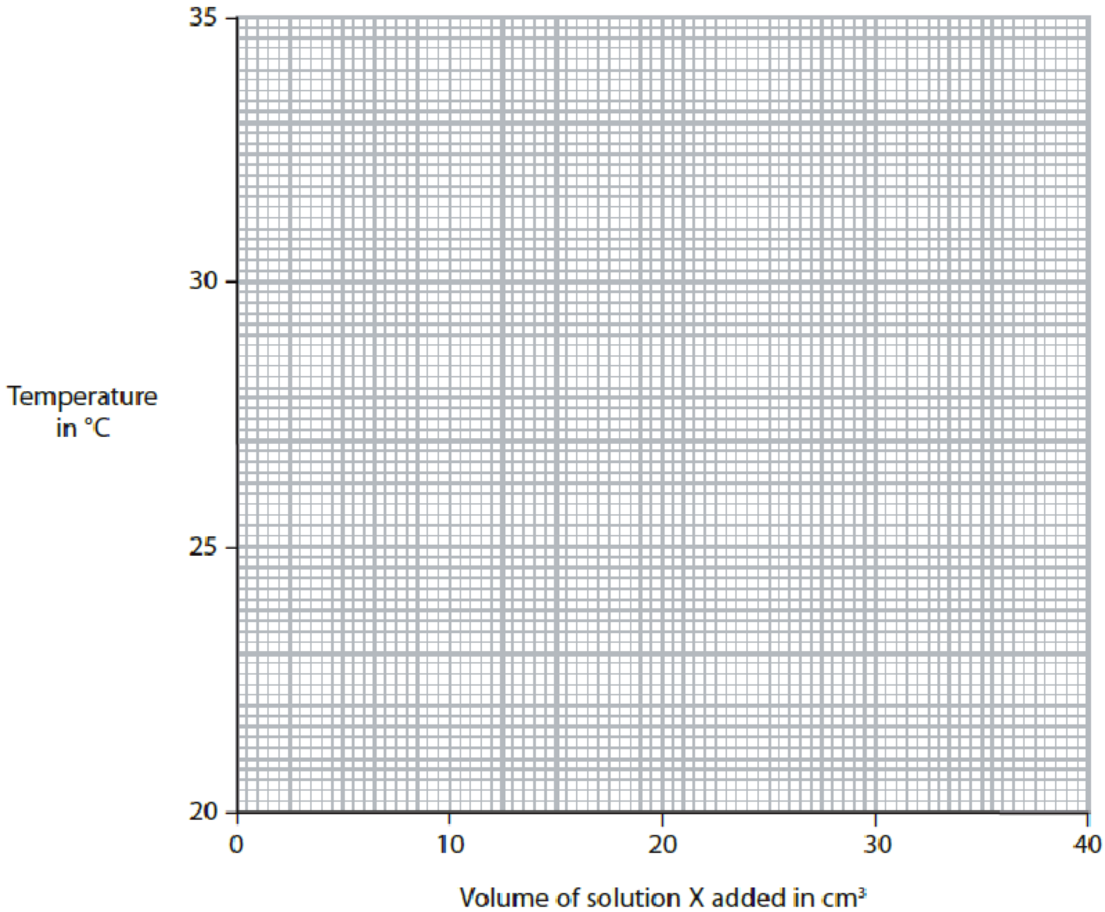
The table shows the results she obtains in Experiment 1.

Experiment 1 – Solution X	
Volume in cm^3 of solution X added	**Temperature in °C**
0	23.0
5	27.0
10	31.0

(a) Plot the results for Experiment 1 on the grid.
Draw a straight line of best fit through the first three points and a second straight line of best fit through the last six points.

Make sure that the two straight lines cross.

(4)



(b) (i) Use the graph to determine the volume of solution X that will produce the maximum temperature rise when added to 20 cm³ of the aqueous sodium hydroxide.

(1)

volume of solution X = ... cm³

(ii) Use the graph to determine the maximum temperature rise.

(1)

maximum temperature rise = ... °C

(c) Why did the student rinse the burette first with water, and then with solution Y, before performing Experiment 2?

(2)

water ....

....

solution Y ....

....

(d) The maximum temperature rise in Experiment 2 was less than that in Experiment 1. Suggest a reason why.

(1)

....

....

....

(b) The student prepares six samples of potassium nitrate, each with a mass of 2.0 g.

She pours 50 cm^3 of water into a 100 cm^3 beaker and records the temperature of the water.

She then uses this method to find the change in temperature as she adds each sample of potassium nitrate.

- add the first sample of potassium nitrate to the beaker and stir until the sample dissolves
- record the temperature of the solution

(ii) From the graph, find the mass of potassium nitrate that would be needed to produce a temperature change of 10.0 °C.

(1)

....

....

(iii) Explain how the student's results show the type of heat change that occurs when potassium nitrate dissolves in water.

(2)

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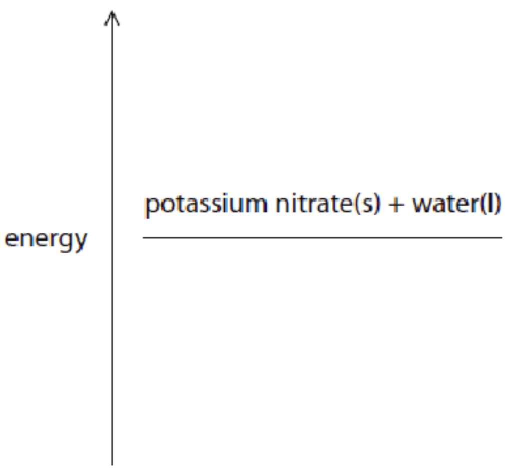
....

....

(iv) Complete the energy level diagram for this experiment.

(1)

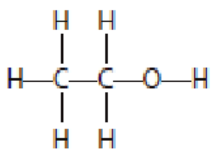


(b) Ethanol can be used as a fuel.

This is the equation for the complete combustion of ethanol.

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$

These are the displayed formulae for ethanol, oxygen, carbon dioxide and water.

	O═O	O═C═O	H—O—H
ethanol	oxygen	carbon dioxide	water

The table gives some average (mean) bond energies.

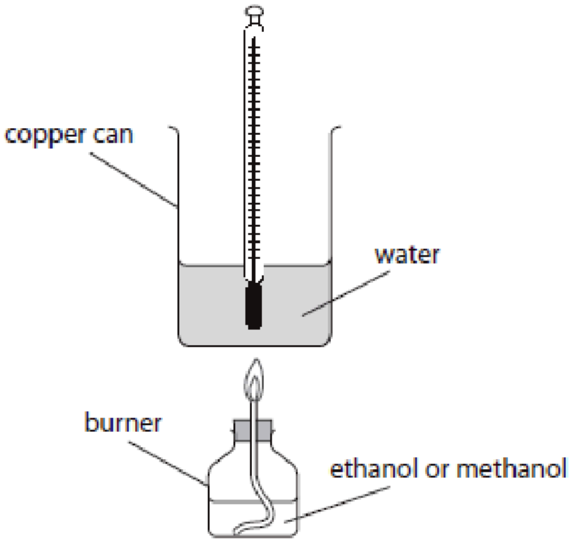
Bond	Average bond energy in kJ/mol
C—C	348
C—H	412
C—O	360
H—O	463
O═O	496
C═O	743

Use this information to calculate the enthalpy change (Δ*H*) when one mole of ethanol is completely burned.

(4)

(c) Ethanol and methanol can both be used as fuels.

A student uses this apparatus to find out how much energy is produced when one mole of ethanol and one mole of methanol are burned.



(ii) The student uses the same burner and copper can in each experiment.

State two other factors that the student should keep the same in each experiment.

(2)

1 ....

....

2 ....

....

(iii) A data book states that the energy given out when 1 mol of ethanol is burned is 1371 kJ.

Suggest two reasons why the student's value is much less than this.

(2)

1 ....

....

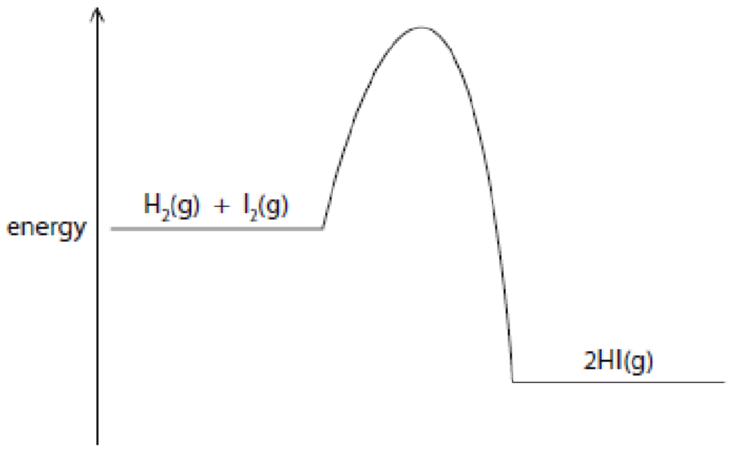
2 ....

8 Hydrogen iodide can be manufactured from its elements using this reaction.

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad \Delta H = -9 \text{ kJ/mol}$$

A temperature of 500 °C, a pressure of 4 atm and a platinum catalyst are used in this manufacturing process.

(a) The diagram shows the reaction profile if a catalyst is not used.



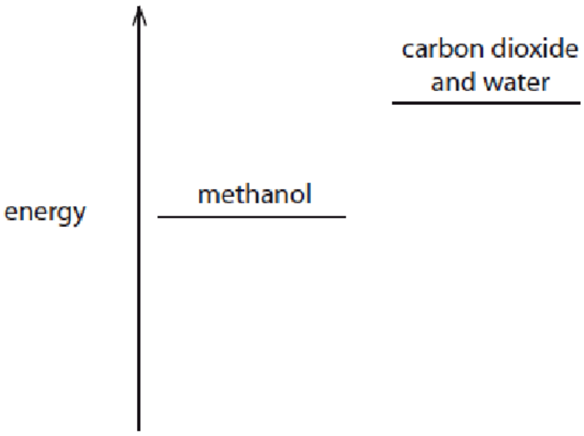
(i) On the diagram, draw the reaction profile when a platinum catalyst is used.

(1)

(ii) Label the diagram to show the enthalpy change (ΔH) and the activation energy (E_{cat}) for the reaction with the catalyst.

(2)

(iii) The student draws an energy level diagram for the complete combustion of methanol.



Identify the two mistakes in his diagram.

(2)

1 ....

....

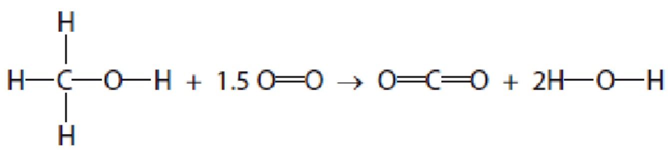
2 ....

....

(c) The student is given this table of average (mean) bond energies.

Bond	C—H	C—O	O—H	O═O	C═O
Average bond energy in kJ/mol	412	360	463	496	743

The equation for the complete combustion of methanol is



Use this equation and the information in the table to calculate another value for the molar enthalpy change, ΔH, for the combustion of methanol.

(4)

11 Hydrogen is used as a fuel. The equation for its combustion is

$$2H_2 + O_2 \rightarrow 2H_2O$$

This lesson involves …

*Automaticity
*Speed & Accuracy

…erms of the energy changes involved in breaking and making bonds, why … gives out heat.

(3)

....

....

....

....

....

....

3 A group of students planned an experiment to find the temperature rise in a neutralisation reaction. This is their method.

- Use a measuring cylinder to add 25 cm^3 of an alkali to a 100 cm^3 beaker
- Record the temperature of the alkali
- Use a burette to add an acid to the alkali in 5.0 cm^3 portions
- Record the temperature of the mixture after adding each portion of acid
- Stop the experiment when the neutralisation is complete

(a) The teacher asked the students about their method.

Suggest an answer to each of her questions.

(i) Why would it be better to use a pipette instead of a measuring cylinder?

(1)

....

....

(ii) It would be better if a polystyrene cup were used instead of a beaker.

What property of polystyrene makes this an improvement?

(1)

....

....

(iii) What extra step should there be between adding each portion of acid and measuring the temperature?

(1)

....

....

(iv) How would you know when the neutralisation was complete?

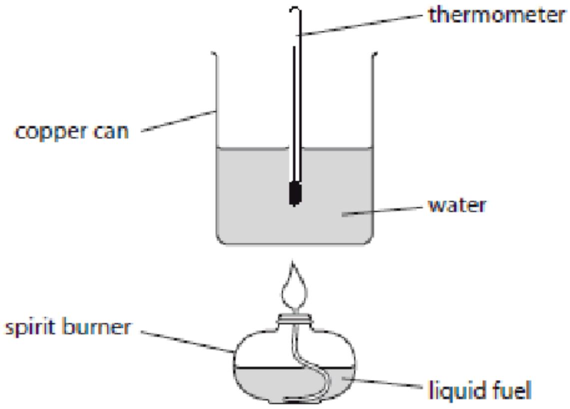
(1)

....

....

11 A student burned four liquid fuels in order to compare the amount of energy they released, in the form of heat.

She used this apparatus.



This lesson involves …

*Automaticity
*Speed & Accuracy

The energy released when each fuel was burned was used to raise the temperature of 100 g of water. For each fuel, the student recorded the mass of fuel burned and the increase in temperature of the water.

Her results are shown in the table.

Fuel	Average relative formula mass	Mass of fuel burned in g	Amount of fuel burned in mol	Increase in temperature in °C
diesel	170	4	0.024	15
ethanol	46	3	0.065	10
methanol	32	2	0.063	5
petrol	114	1	0.009	4

The best fuel is the one that releases the most energy.

(a) The student suggested that petrol was the best fuel.

Explain why, using the information in the table.

(1)

....

....

(b) Another student suggested that diesel was the best fuel.

Explain why, using the information in the table.

(1)

....

....

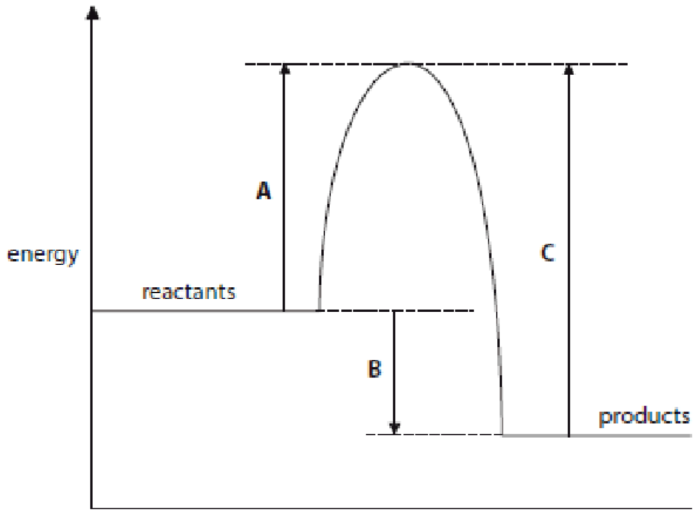
(c) In another experiment, a student burned propanol and then used his results to calculate the energy released when one mole of propanol was burned.

He then compared his result with a value from a data book.

The values are shown in the table.

	Energy released per mole of propanol burned in kJ
Student's result	1020
Data book value	2010

(d) The diagram shows the energy profile for burning a fuel.



Which of the energy changes A, B or C represents

- the activation energy for the reaction
- the amount of energy given out during the reaction?

(2)

Activation energy = ...

Energy released = ...

(e) Explain, in terms of bond breaking and bond making, why this reaction gives out energy.

(3)

....

....

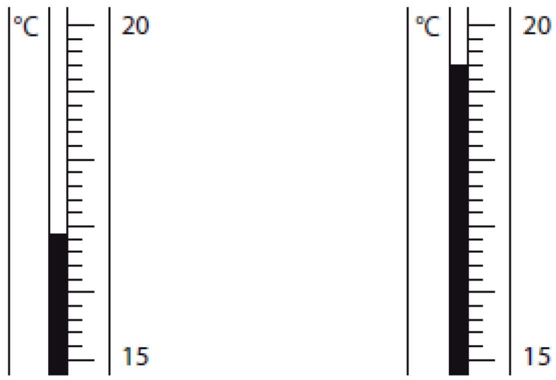
....

....

....

(Total for Question 11 = 9 marks)

(b) The diagram shows the thermometer readings in one experiment.



before adding alkali after adding alkali

This lesson involves …

*Automaticity
*Speed & Accuracy

Write down the thermometer readings and calculate the temperature change.

(3)

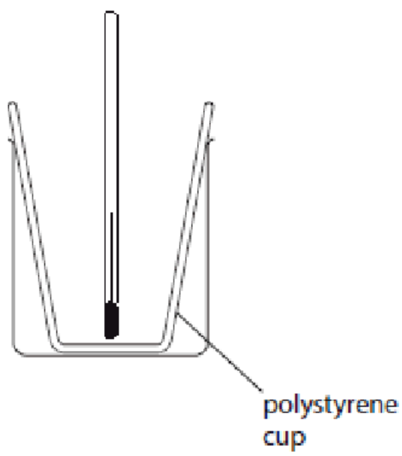
temperature after adding alkali .... °C

temperature before adding alkali .... °C

temperature change .... °C

4 A student investigated the neutralisation of acids by measuring the temperature changes when alkalis were added to acids of known concentrations.

He used this apparatus to add different volumes of sodium hydroxide solution to a fixed volume of dilute nitric acid.



He used this method.

- measure the temperature of 25.0 cm^3 of the acid in the polystyrene cup
- add the sodium hydroxide solution in 5.0 cm^3 portions until a total of 30.0 cm^3 has been added

(a) State two properties of the sodium hydroxide solution that should be kept constant for each 5.0 cm^3 portion.

(2)

1 ....

(c) The student carried out the experiment three times.

The table shows his results.

Volume of alkali added in cm^3	Temperature in °C		
	experiment 1	experiment 2	experiment 3
0.0	17.4	16.6	15.9
5.0	18.5	21.0	18.0
10.0	19.6	24.5	20.0
15.0	20.5	23.6	22.2
20.0	21.4	22.7	23.6
25.0	22.5	21.4	22.8
30.0	23.4	20.5	22.0

The teacher said that only the results for experiment 3 showed the expected increase and decrease in temperature.

(i) Why was there no temperature decrease in experiment 1?

(1)

☐ **A** The alkali was added too quickly

☐ **B** The starting temperature of the acid was too high

☐ **C** The acid concentration was half what it should have been

☐ **D** The volume of acid used was 50.0 cm^3 instead of 25.0 cm^3

(ii) Why were the temperature increases in experiment 2 much greater than expected?

(1)

☐ **A** The starting temperature of the acid was too high

☐ **B** The acid concentration was double what it should have been

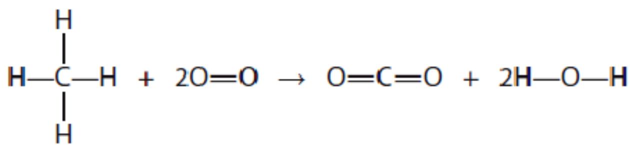
☐ **C** The volume of acid used was 50.0 cm^3 instead of 25.0 cm^3

☐ **D** The alkali was added in 10.0 cm^3 portions but were recorded as 5.0 cm^3 portions

(c) The student uses a table of average bond energies to calculate another value for the molar enthalpy of combustion of methane.

Bond	C—H	O═O	C═O	H—O
Average bond energy in kJ/mol	412	496	743	463

The equation for the combustion can be shown using displayed formulae.



(i) Use values from the table to calculate the energy taken in when the bonds in the reactants are broken.

(2)

energy taken in = ... kJ

(ii) Use values from the table to calculate the energy given out when the bonds in the products are formed.

(2)

energy given out = ... kJ

(iii) Use your answers to (i) and (ii) to calculate the molar enthalpy change for the combustion of methane.

(1)

molar enthalpy change = ... kJ/mol

(c) Balance the equation for the reaction used in this preparation of oxygen.

(1)

$$\ldots\ldots\ H_2O_2 \rightarrow \ldots\ldots\ H_2O + \ldots\ldots\ O_2$$

(d) The manganese(IV) oxide acts as a catalyst.

What is meant by the term **catalyst**?

(2)

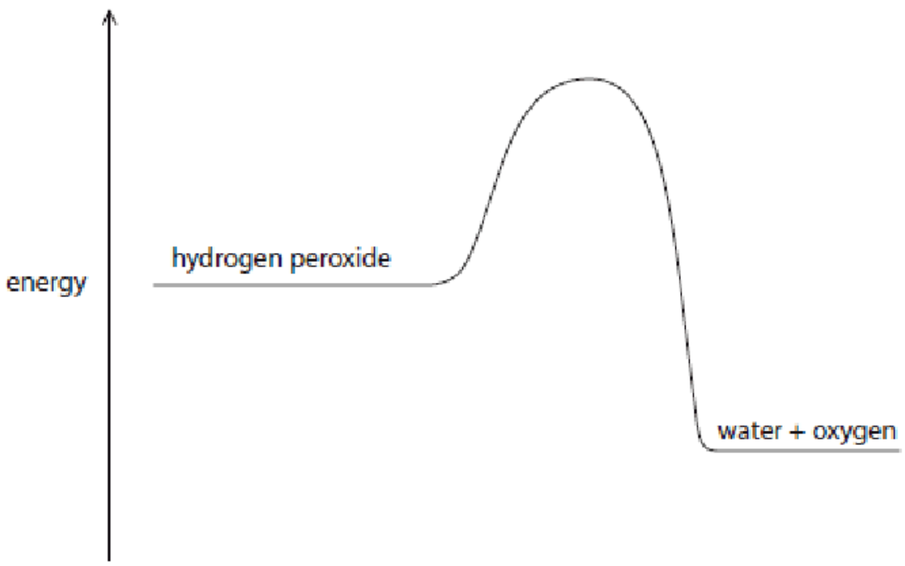
....

....

....

....

(e) The diagram shows the reaction profile for the decomposition of hydrogen peroxide without a catalyst.



(i) Label the diagram to show the activation energy (E_a) for this reaction.

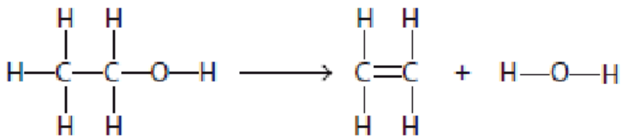
(1)

(ii) On the diagram, draw a curve to represent the reaction profile for the same reaction when a catalyst is used.

(1)

(Total for Question 4 = 9 marks)

(d) The equation for the conversion of ethanol into ethene can be written using displayed formulae.



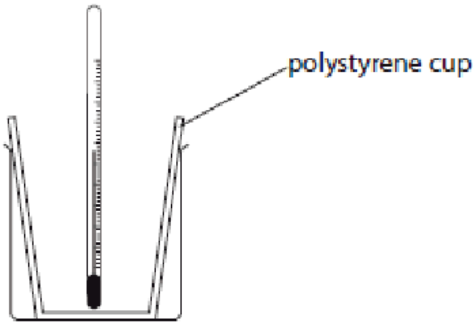
The table gives some average bond energies.

Bond	Average bond energy in kJ/mol
C—C	348
C=C	612
C—H	412
C—O	360
O—H	463

Use information from the table to calculate the enthalpy change, in kJ/mol, for the conversion of ethanol into ethene.

(4)

5 A student uses this apparatus to investigate the temperature change that occurs when potassium hydroxide is dissolved in water.



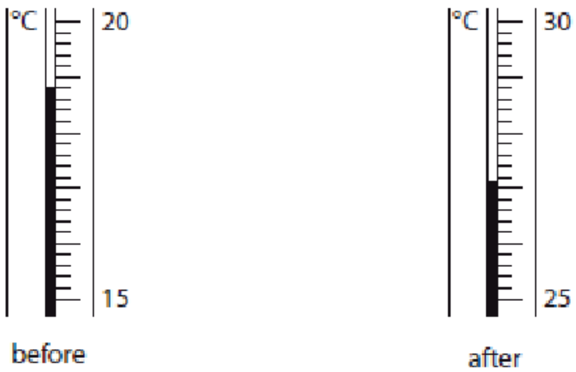
This lesson involves …

*Automaticity
*Speed & Accuracy

She uses this method.

- pour 50 cm^3 of water into the polystyrene cup and measure the temperature of the water
- add 3 g of potassium hydroxide and stir
- record the highest temperature of the solution

(a) These diagrams show the thermometer readings before and after the student added the potassium hydroxide.



Use the readings to complete the table.

(3)

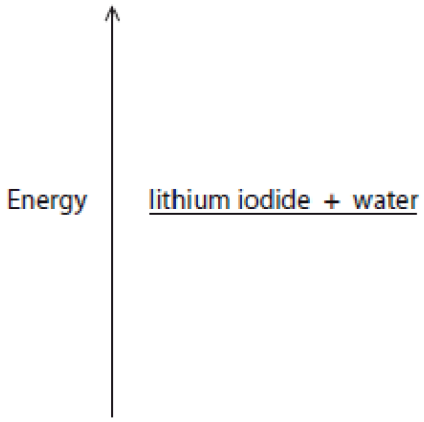
temperature in °C after adding potassium hydroxide	
temperature in °C before adding potassium hydroxide	
temperature change in °C	

(ii) The temperature change in this experiment shows that dissolving lithium iodide in water to form lithium iodide solution is an exothermic process.

Complete the energy level diagram to show the position of the lithium iodide solution.

Label the diagram to show Δ*H*, the molar enthalpy change.

(2)



5 A teacher investigates the temperature changes that occur when sodium hydroxide solution is added to dilute hydrochloric acid.

This is the method she uses.

- place some of the acid in a glass beaker and measure its temperature
- add a known volume of sodium hydroxide solution
- stir the mixture and record the highest temperature reached
- repeat the experiment with different volumes of sodium hydroxide solution

(a) State two factors that the teacher must keep constant to make this a valid investigation (a fair test).

(2)

1 ....

2 ....

(b) Explain how the use of a polystyrene cup, in place of a glass beaker, will affect the accuracy of the results.

(2)

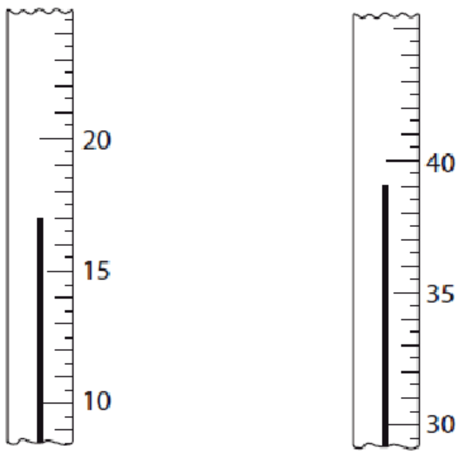
....

....

....

....

(c) (i) The diagram shows the thermometer readings for one of the experiments.



Record the temperatures and calculate the temperature change.

(3)

final temperature of mixture	....°C
initial temperature of acid	....°C
temperature change	....°C

(ii) State how the temperature change shows whether the reaction between sodium hydroxide and hydrochloric acid is exothermic or endothermic.

(1)

....

....

10.3 - Moles

I have reviewed the syllabus statements for this topic	

I have completed the questions in this section	
I have read the relevant sections of the College Website	
I have made some revision material (mind-map, key-words & definitions etc)	
Prep Grade	
Test Grade	

Target	Pupil Signature

10.3 - Moles

- The ***law of conservation of mass*** states that no atoms are lost or made during a chemical reaction so the mass of the products equals the mass of the reactants.
- This means that chemical reactions can be represented by symbol equations which are ***balanced*** in terms of the numbers of atoms on both sides of the equation.
- Students should understand the use of the multipliers in equations in normal script before a formula and in subscript within a formula.
- ***Some reactions may appear to involve a change in mass*** but this can usually be explained because a reactant or product is a ***gas*** & its mass has not been taken into account. Eg: when a metals with Oxygen the mass of the Oxide produced is greater than the mass of the metal or in thermal decompositions of Carbonates where Carbon Dioxide escapes into the atmosphere

- Students should be able to ***explain*** any observed changes in mass in non-enclosed systems during a chemical reaction ***given the balanced symbol equation*** for the reaction and explain these changes in terms of the particle model.
- Whenever measurements are made there's always some ***uncertainty*** about results
- Students should be able to:

1. represent the distribution of results and make ***estimations of uncertainty***
2. use the ***range*** of measurements about the mean as a ***measure of uncertainty***.

- The ***relative formula mass*** (M_r) of a compound is the ***sum of the relative atomic masses*** of the atoms in the numbers shown in the formula.
- In a ***balanced equation***, the sum of the relative formula masses of the reactants in the quantities shown ***equals*** the sum of the relative formula masses of the products in the quantities shown.

- Chemical amounts are measured in ***moles***.
- The symbol for the unit mole is ***mol***.
- The ***mass of 1 mole*** of a substance is ***equal to its relative formula mass***.
- One mole of a substance contains the ***same number*** of the stated particles, atoms, molecules or ions as one mole of any other substance.
- The number of atoms, molecules or ions in a mole of a given substance is the ***Avogadro constant.*** The value of the Avogadro constant is ***6.02×10^{23} per mole***.
- Students should understand that the measurement of amounts in moles can apply to atoms, molecules, ions, electrons, formulae and equations, for example that in one mole of Carbon (C) the number of atoms is the same as the number of molecules in one mole of Carbon Dioxide (CO_2).
- Students should be able to use the ***relative formula mass*** of a substance to calculate the ***number of moles*** in a given mass of that substance ***and vice versa***.
- The masses of reactants & products can be calculated from balanced equations.
- Chemical equations can be interpreted in terms of moles. For example:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

shows that one mole of magnesium reacts with two moles of Hydrochloric acid to produce one mole of Magnesium Chloride and one mole of Hydrogen gas.

- Students should be able to:

1. calculate the masses of substances shown in a balanced symbol equation
2. calculate the masses of reactants and products from the balanced symbol equation and the mass of a given reactant or product.

- In a chemical reaction involving two reactants, it is common to use an ***excess*** of one of the reactants to ensure that all of the other reactant is used.
- The reactant that is ***completely used*** up is called the ***limiting reactant*** because it limits the amount of products.

- Students should be able to explain the effect of a limiting quantity of a reactant on the amount of products it is possible to obtain in terms of amounts in moles or mass

- Many chemical reactions take place in solutions.
- The ***concentration*** of a solution can be measured in ***grams per dm^3*** (***g/dm^3***).
- Students should be able to ***calculate*** the mass of solute in a given volume of solution of known concentration in terms of mass per given volume of solution
- (HT only) explain how the mass of a solute and the volume of a solution is related to the concentration of the solution.

- The ***concentration*** of a solution can be measured in ***mol/dm^3*** .
- The amount in moles of solute or the mass in grams of solute in a given volume of solution can be calculated from its concentration in mol/dm^3 .
- If the volumes of two solutions that react completely are known and the concentration of one solution is known, the concentration of the other solution can be calculated.
- Students should be able to ***explain*** how the concentration of a solution in mol/dm^3 is related to the mass of the solute and the volume of the solution.

- The volumes of acid and alkali solutions that react with each other can be measured by ***titration*** using a suitable indicator.
- Students should be able to

1. describe how to ***carry out titrations*** using strong acids and strong alkalis only (Sulphuric, Hydrochloric and Nitric acids only) to find the reacting volumes ***accurately***
2. (HT Only) calculate quantities in titrations involving concentrations in mol/dm^3 and in g/dm^3 .

Required practical 2: determination of the reacting volumes of solutions of a strong acid and a strong alkali by titration.

(HT only) determination of the concentration of one of the solutions in mol/dm^3 and g/dm^3 from the reacting volumes and the known concentration of the other solution.

- Equal amounts in moles of gases occupy the ***same volume*** under the same conditions of temperature and pressure.
- The volume of one mole of any gas at room temperature and pressure (20°C and 1 atmosphere pressure) is ***24 dm^3*** .
- The volumes of gaseous reactants and products can be calculated from the ***balanced equation*** for the reaction.

- Students should be able to:

1. ***calculate the volume of a gas*** at room temperature and pressure from its mass and relative formula mass
2. calculate volumes of gaseous reactants and products from a balanced equation and a given volume of a gaseous reactant or product.

- The balancing numbers in a symbol equation can be calculated from the masses of reactants and products by converting the masses in grams to amounts in moles and converting the numbers of moles to simple whole number ratios.
- Students should be able to balance an equation given the masses of reactants and products.
- Students should be able to change the subject of a mathematical equation.

Conservation of mass

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

The above is an equation for burning Methane (CH_4) in Oxygen (O_2), which produces Carbon Dioxide (CO_2) and Water Vapour (H_2O)

This lesson involves:

*Logical Thinking
*Precision
*Multi-step problem-solving

But when you count the atoms on either side...

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

C=1		C=1
H=4	$\neq$	H=2
O=2		O=3

.., we can see it's unbalanced, its not possible for Hydrogen atoms to disappear or for Oxygen atoms to appear from nowhere.

Hopefully you'd know to balance the Hydrogen atoms first...

$$CH_4 + O_2 \rightarrow CO_2 + 2H_2O$$

C=1		C=1
H=4	$\neq$	H=4
O=2		O=4

But that we'd still need to balance the Oxygen atoms...

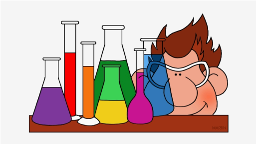
$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

C=1		C=1
H=4	=	H=4
O=4		O=4

This equation is now balanced because it has the same number of each type of element on both sides of the equation.

Which means that the total mass on the left equals the total mass on the right.

This is an example of the ***Law of Conservation of Mass -*** mass can't be created or destroyed, just redistributed.



You should be able to do the same with the following…

QUIZ #1- Balancing chemical equations

Balance the following formula equations.

1- $\_\_\_C_3H_8 + \_\_\_O_2 \rightarrow \_\_\_CO_2 + \_\_\_H_2O$

2- $\_\_\_Al_2(SO_4)_3 + \_\_\_Fe_3(PO_4)_2 \rightarrow \_\_\_AlPO_4 + \_\_\_FeSO_4$

3- $\_\_\_Na + \_\_\_Cl_2 \rightarrow \_\_\_NaCl$

4- $\_\_\_N_2 + \_\_\_O_2 \rightarrow \_\_\_N_3O_7$

5- $\_\_\_NaNO_3 + \_\_\_Al_2(CO_3)_3 \rightarrow \_\_\_Na_2CO_3 + \_\_\_Al(NO_3)_3$

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

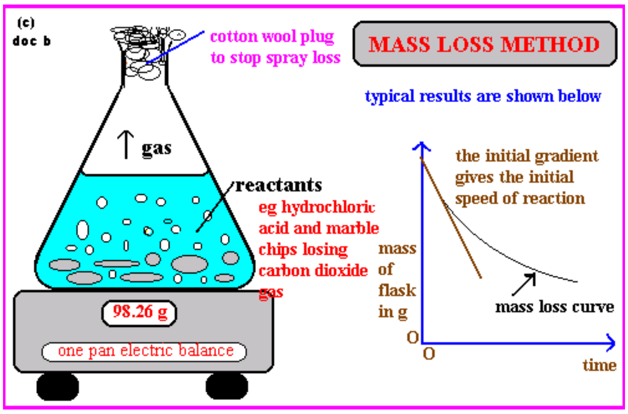
This lesson involves …

*Automaticity
*Speed & Accuracy

Why do some reactions seem to disobey the law of conservation of mass?

Losing Gases

- Put a few marble chips in acid and quickly place the flask on a balance.
- Record the mass every 30 secs until the reaction stops.



A

This lesson involves ...

*Generalisation
*Connection-finding
*Big Picture Thinking
*Abstraction
*Imagination

- The mass decreases – where does it go?

____

____

- How does the equation below show that this reaction obeys the law of conservation of mass despite the apparent mass loss?

$$CaCO_{3(s)} + 2\ HCl_{(aq)}\ \square\ CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

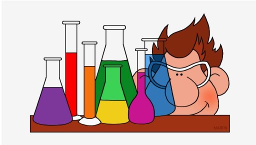
____

____

____

Gaining Gas

Rusting is an example of a reaction where a solid (Iron) reacts with a gas (Oxygen).



❖ Why does the nail get heavier?

____

❖ An equation for rusting is shown below. How does it show that the law of conservation of mass isn't broken during rusting?

$$4Fe(s) + 3O_2(aq) + 3H_2O(l) \longrightarrow 2Fe_2O_3 \cdot 3H_2O(s)$$

____

This lesson involves …

*Generalisation
*Connection-finding
*Big Picture Thinking
*Abstraction
*Imagination

Uncertainty.

If we did the rusting experiment we might find that the results didn't quite match what you were expecting.

That could be because rust flakes off, removing mass.

But, even if no flakes were lost, there might be some problems due to ***uncertainty***.

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

If this digital balance measured the mass of something as 14.7g ….

a) What is the maximum mass it might read if we changed it to measure 2 dp?

____

b) What is the minimum mass it might read if we changed it to measure 2 dp?

____

c) What is the maximum mass it might read if we changed it to measure 9 dp?

____

d) What is the minimum mass it might read if we changed it to measure 9 dp?

____

e) If we use the balance on a 1dp setting then it is really reading

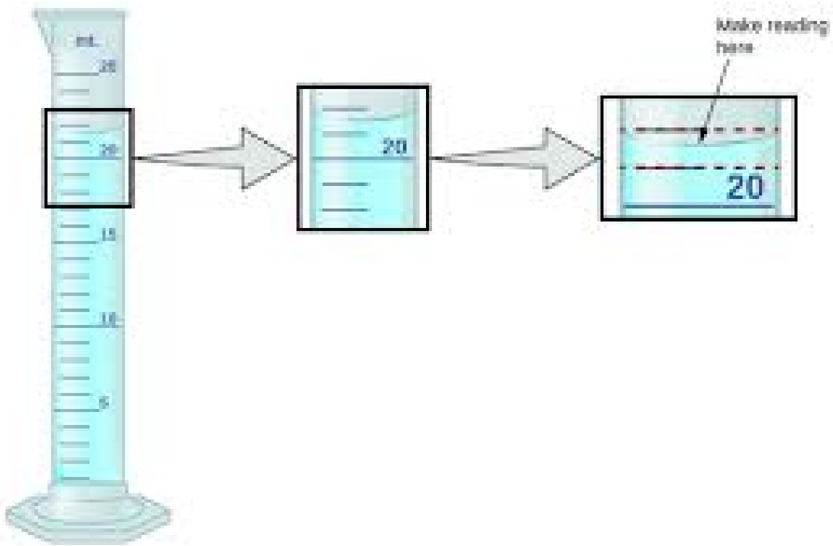
So, the reading is only **14.7 +/-** _____ **g**

❖ If we read this measuring cylinder to the nearest cm3 we'd have to decide whether it read 20 or 21.

❖ But the uncertainty would be +/- _______

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving



❖ Which of these measuring cylinders would have the lowest uncertainty? How do we know?

The law of conservation of mass states that no atoms are lost or made during a chemical reaction. This means that the mass of the products equals the mass of the reactants. A student carried out two experiments to investigate this law.

Experiment 1

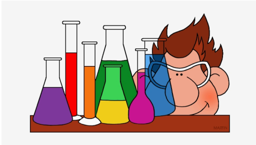
The student found the mass of an ice cube in a beaker. They left the ice cube to melt and found the mass of the beaker and water. There was no change in mass.

Experiment 2

The student found the mass of a strip of magnesium ribbon in a crucible. They placed a lid on the crucible and heated it while lifting the lid a little from time to time. After cooling the crucible, the student found that the mass of crucible and contents had increased.

Explain why there was no change in mass during Experiment 1 but the mass increased during Experiment 2.

[6 marks]



Mark scheme

Answers	Mark
Level 3: Relevant points (reasons/causes) are identified, given in detail and logically linked to form a clear account.	5–6
Level 2: Relevant points (reasons/causes) are identified, and there are attempts at logically linking. The resulting account is not fully clear.	3–4
Level 1: Points are identified and stated simply, but their relevance is not clear and there is no attempt at logical linking.	1–2
No relevant content	0
Indicative content Experiment 1 • This is a physical change / solid changes to liquid. • There is no change in the number of hydrogen and oxygen atoms. • There is no change in the number of (H_2O) molecules in the ice and the water • As no new molecules are added and none are removed. • Therefore, the masses are the same and this confirms the law of conservation of mass. Experiment 2 • This is a chemical change. • Magnesium reacts with oxygen (from the air) • To form magnesium oxide. • $2Mg + O_2 \rightarrow 2MgO$ • There are the same number of atoms of magnesium and oxygen on each side of the equation. • The student only weighs the magnesium and magnesium oxide. • Magnesium oxide is heavier than magnesium as it also contains oxygen. • The student could not weigh the oxygen from the air that reacted with the magnesium: • *mass of magnesium + mass of oxygen = mass of magnesium oxide* • So this also confirms the law of conservation of mass. • The student could determine the mass of oxygen: • *mass of oxygen used = mass of magnesium oxide – mass of magnesium*	

Relative Formula Masses

Questions

Find M_r of the following using the table (right)

Element	Symbol	Relative atomic mass
hydrogen	H	1
carbon	C	12
nitrogen	N	14
oxygen	O	16
sulfur	S	32
chlorine	Cl	35.5

a. H_2O b. CO_2

c. SO_3 d. H_2SO_4

e. H_2O_2 f. $HClO_3$

g. NH_4Cl h. NH_4NO_3

i. C_2H_6 j. C_2H_5OH

M_r and brackets

Sometimes a formula has brackets.

This shouldn't make calculating the M_r much more difficult.

You already know how to remove brackets in Maths. Eg 2(x+3) = 2x + 6

You can do the same in Chemistry – it's just that the number comes after the brackets.

Eg. $Sr(HCO_3)_2$ = $SrH_2C_2O_6$

Q1. You will need to use your Periodic Table to look up Masses of the elements and use them to find the Mr of the following compounds.

a. $Mg(OH)_2$ b. $Fe(OH)_3$

c. $Ca(HSO_4)_2$ d. $Ga(HCO_3)_2$

e. $Fe(NO_3)_2$ f. $Fe(NO_3)_3$

g. $(NH_4)_2SO_4$ h. $(NH_4)_3PO_4$

M_r and Water of Crystallisation

Some salts absorb water as they crystallise – the water of crystallisation.

To calculate the M_r we just find the M_r of the salt on its own and add the mass of water to it, Eg. $CuSO_4 \bullet 5H_2O = CuSO_4$ & $5\ H_2O$

M_r ($CuSO_4$) = $_{(Cu)}$ 63.5 + $_{(S)}$ 32 + $_{(O)}$16 x 4 = 65.5 + 32 + 64 = 159.5

M_r (H_2O) = $_{(H)}$1 x 2 + $_{(O)}$16 = 18

M_r ($CuSO_4.5H_2O$) = 159.5 + 5x18 = 249.5

Q2. **Find the M_r of**

a. $CoCl_2 \bullet 6H_2O$

b. $CaCl_2 \cdot 2H_2O$

c. $SnCl_2 \bullet 2H_2O$

Conservation of Mass Again.

potassium	+	water	→	potassium hydroxide	+	hydrogen
2K	+	$2H_2O$	→	2KOH	+	H_2

The sum of M_rs on the left = __________________ g/mol

The sum of M_rs on the right = __________________ g/mol

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$$

The sum of M_rs on the left = __________________ g/mol

The sum of M_rs on the right = __________________ g/mol

So, if the sum on the left doesn't equal the sum on the right – you equation isn't balanced.

Moles.

A mole is an amount of stuff like a dozen eggs (12) or a trio of chocolate desserts (3).

You could equally have a trio of eggs and a dozen chocolate desserts.

You can't really have a mole of eggs because you would need 600,000,000,000,000,000,000,000 of them, which is millions of times more eggs than there are chickens.

Why use moles?

Look at this balanced equation: $2\ H_2O_{2(aq)} \rightarrow 2\ H_2O_{(l)} + O_{2(g)}$

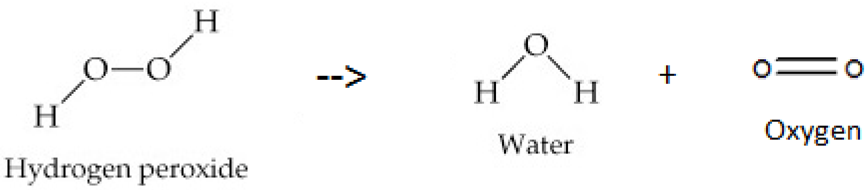
We know that unwritten numbers are 1's. So, $2\ H_2O_{2(aq)} \rightarrow 2\ H_2O_{(l)} + 1\ O_{2(g)}$

If this equation worked in mass it would mean that 2g of H_2O_2 made 2g of water ***and*** 1g of Oxygen.

That's clearly not possible.

It actually means that 2 molecules of peroxide makes 2 molecules of water and 1 of oxygen.

This is fine because there are the same atoms on both sides of the equation



But we can't really deal with 1 molecule at a time so we use moles (or fractions of moles, or multiples of moles)

$2\ H_2O_{2(aq)} \rightarrow 2\ H_2O_{(l)} + 1\ O_{2(g)}$

Means that 2 moles of peroxide makes 2 moles of water and 1 mole of oxygen.

So, 20 moles of peroxide makes 20 moles of water and 10mole of oxygen.

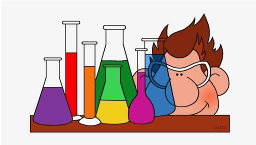
Or, 0.2 moles of peroxide makes 0.2 moles of water and 0.1mole of oxygen etc

This is because the reacting ratio is

$H_2O_2 : H_2O : O_2$

2 : 2 : 1

Reacting ratios



Example

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Ratio CH_4 : O_2 : CO_2 : H_2O

1 : 2 : 1 : 2

1. How many moles of CO_2 would be made from 1.2 moles of CH_4?

CH_4 : CO_2

1 : 1 So, 1.2 moles of CH_4 makes 1.2 moles of CO_2

2. How many moles of H_2O would be made from 0.25 moles of CH_4?

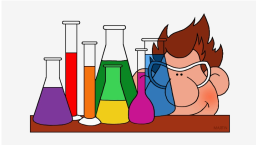
CH_4 : H_2O

1 : 2 So, 0.25 moles of CH_4 makes 0.5 moles of H_2O

3. How many moles of CO_2 would be made from 2.4 moles of O_2?

O_2 : CO_2

2: 1 So, 2.4 moles of O_2 makes 1.2 moles of CO_2



Using ratios

1. $$2H_2 + O_2 \rightarrow 2H_2O$$

a. Write the reacting ratio…

b. How much water is made from i) 20 moles of Hydrogen
ii) 20 moles of Oxygen
iii) 0.6 moles of Hydrogen
iv) 0.6 moles of Oxygen

c. How many moles of Oxygen are required to make 0.44 moles of water?

d. How many moles of Oxygen are required to react exactly with 1.28 moles of Hydrogen?

2. $$\underset{\text{methane}}{CH_4} + \underset{\text{oxygen}}{2O_2} \longrightarrow \underset{\text{carbon dioxide}}{CO_2} + \underset{\text{water}}{2H_2O}$$

a. Write the reacting ratio…

b. How much water is made from i) 2 moles of Methane
ii) 10 moles of Oxygen
iii) 1.6 moles of Methane
iv) 1.6 moles of Oxygen

c. How much CO_2 is made from i) 0.4 moles of Methane
ii) 5 moles of Oxygen
iii) 0.8 moles of Methane
iv) 2.6 moles of Oxygen

d. How many moles of Oxygen are required to make 0.44 moles of water?

e. How many moles of Oxygen are required to react exactly with 1.08 moles of Methane?

3

$$C_3H_8 + 5O_2 \rightarrow 4H_2O + 3CO_2$$

a. Write the reacting ratio…
b. How much water is made from i) 2 moles of Propane (C_3H_8)
ii) 10 moles of Oxygen
iii) 1.6 moles of Propane
iv) 1.6 moles of Oxygen

c. How much CO_2 is made from i) 0.4 moles of Propane (C_3H_8)
ii) 5 moles of Oxygen
iii) 0.8 moles of Propane
iv) 2.6 moles of Oxygen

d. How many moles of Oxygen are required to make 0.44 moles of water?
e. How many moles of Oxygen are required to react exactly with 1.08 moles of Propane?

4.

$$4Al + 3O_2 \rightarrow 2Al_2O_3$$

a. Write the reacting ratio…
b. How much Aluminium Oxide is made from i) 8 moles of Aluminium (Al)
ii) 6 moles of Oxygen
iii) 1.6 moles of Aluminium
iv) 1.8 moles of Oxygen

c. How many moles of Oxygen are required to make 0.8 moles of Aluminium Oxide?
d. How many moles of Oxygen are required to react exactly with 1.2 moles of Aluminium?

1.

$$2Al + 3H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3H_2$$

a. Write the reacting ratio…

b. How much Aluminium Sulphate is made from
i) 8 moles of Aluminium (Al)
ii) 6 moles of H_2SO_4
iii) 1.6 moles of Aluminium
iv) 1.8 moles of H_2SO_4

c. How many moles of Aluminium are required to make 0.8 moles of Aluminium Sulphate?

d. How many moles of H_2SO_4 are required to react exactly with 1.4 moles of Aluminium?

6.

$$2C_4H_{10} + 13O_2 \longrightarrow 8CO_2 + 10H_2O$$

a. Write the reacting ratio…

b. How much CO_2 is made from
i) 7 moles of Butane (C_4H_{10})
ii) 13 moles of Oxygen
iii) 1.5 moles of Butane
iv) 2.7 moles of Oxygen

c. How many moles of Butane are required to make 1.9 moles of Water?

d. How many moles of Butane are required to react exactly with 1.4 moles of Oxygen?

How many is a mole?

We already saw that a mole is the number of ^{12}C atoms in 12g of Carbon-12.

Or the number of ^{20}Ne atoms in 20g of Neon-10, or the number of ^{39}K atoms in 39g of Potassium-39

That number is 6.02×10^{23} or 602,000,000,000,000,000,000,000 – which we call Avogadro's Number.

- So, how many ^{23}Na atoms in 46g of ^{23}Na?

1 mole of ^{23}Na would weigh 23g.

This must be 2 moles.

2 x Avogadro's Number = $12.04 \times 10^{23} = 1.204 \times 10^{24}$

- How many ^{7}Li atoms in 0.7g of ^{7}Li?

1 mole of ^{7}Li would weigh 7g.

This must be 1/10 moles.

1/10 x Avogadro's Number = $0.602 \times 10^{23} = 6.02 \times 10^{22}$

Q1. How many ^{40}Ca atoms in 400g of ^{40}Ca?

1 mole of ^{40}Ca would weigh _______g.

This must be _____ moles.

_____ x Avogadro's Number = _________

Q2. How many ^{24}Mg atoms in 12g of ^{24}Mg?

1 mole of ^{24}Mg would weigh _______g.

This must be _____ moles.

_____ x Avogadro's Number = _________

An exam question might ask you:

"How many atoms of Oxygen in 10 moles of Oxygen gas (O_2)"

Answer:

10 moles = 10 x Avogadro's No = 6.02×10^{24} Oxygen molecules

But each molecule contains 2 Oxygen atoms so a total of 12.04×10^{24} Oxygen atoms

"How many atoms of Oxygen in 20 moles of Carbon Dioxide gas (CO_2)"

Answer:

____

"How many atoms of Oxygen in 0.1 moles of Sodium Oxide (Na_2O)"

Answer:

____

"How many atoms of Oxygen in 100 moles of Calcium Carbonate ($CaCO_3$)"

Answer:

____

"How many atoms of Oxygen in 0.2 moles of Nitric Acid (HNO_3)"

Answer:

____

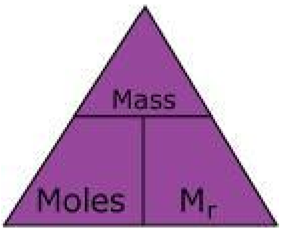
In reality though, we're much more likely to be asked questions that involve turning mass to moles and vice versa.

Changing Mass to Moles. #1

You need to learn the formula Moles = Mass ÷ M_r

And to be able to re-arrange it to Mass = Moles x Mr

And Mr = Mass ÷ Moles



Q1. Use the equation to find the number of moles in:

a) 4.4g of CO_2 M_r (CO_2) = 44 Moles = Mass ÷ M_r = 4.4 ÷ 44 = 0.1 moles

b) 3.6g of H_2O

c) 4.9 g of H_2SO_4

d) 2.4g of SO_3

e) 1.1g of C_3H_8

f) 0.85g of $NaNO_3$

Changing Mass to Moles. #1

g) 2.96g of Ca_3N_2

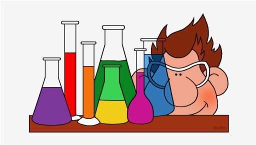
h) 0.81g of $Ca(HCO_3)_2$

i) 0.5g of $Mg(OH)_2$

j) 0.5g of $Fe(OH)_3$

k) 1.6g of $(NH_4)_2SO_4$

l) 2.6g of $CuSO_4$



m) 2.6g of $CuSO_4.5H_2O$

n) 0.01g of $Ga(HCO_3)_2$

o) 0.5g of $C_{20}H_{42}$

Changing Mass to Moles. #2

Q2. Which has more moles?

a) 0.4g of CO_2 or 0.8g of $Fe(OH)_3$

b) 0.5g of H_3PO_4 or 1.5g of $(NH_4)_3PO_4$

Which has more moles?

c) 4g of $Mg(OH)_2$ or 12g of $C_{20}H_{42}$

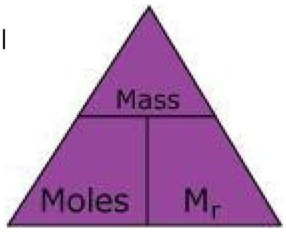
d) 3g of $CuSO_4$ or 8g of $CuSO_4.5H_2O$

e) 2.6 g of SO_3 or 2g of C_3H_7OH

Find the M_r of an unknown substance

If you know the formula of a substance you can add up its M_r easily.

But if you don't know the formula you could still find the Mr as long as you know tl substance and how many moles it contains.



Q1. Find the M_r of a substance in which:

a) 2g contains 0.4 moles.

Moles = Mass ÷ M_r □ M_r = Mass ÷ Moles □ M_r = 2 ÷ 0.4 = 5 g/mol

b) 24g contains 1.2 moles

c) 420g contains 20 moles.

Find the M_r of a substance in which:

d) 6g contains 1.5 moles.

e) 0.3g contains 0.05 moles.

f) 25g contains 1.25 moles.

g) 0.3g contains 1.4×10^{-3} moles.

h) 1.44 ***kg*** contains 3 moles

Find the M_r of an unknown substance #2

2. Which substance has the bigger M_r?

a) 2 moles of Substance A weighing 54g or 3 moles of Substance B weighing 66g

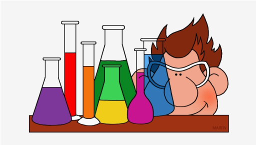
b) 0.1 moles of Substance A weighing 26g or 2 moles of Substance B weighing 512g

Which substance has the bigger M_r?

c) 3 moles of Substance A weighing 180g or 0.2 moles of Substance B weighing 16g

d) 0.2 moles of Substance A weighing 4g or 0.05 moles of Substance B weighing 1.2g

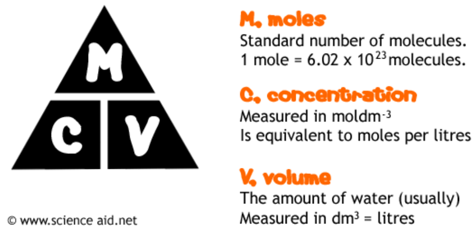
e) 12 moles of Substance A weighing 3.6 ***kg*** or 0.2 moles of Substance B weighing 120g



Concentration and Volume

If we know how much of a solution we have (its volume) and its concentration then we can find the number of moles by:

Moles = Conc. X Volume



However, IGCSE questions usually give volumes in **cm^3**.

And concentration in **mol/dm^3**

Since the units don't match we must first convert the **cm^3** to **mol/dm^3**.

1cm^3 = 1 millilitre 1dm^3 = 1 litre

So, to convert cm^3 to dm^3 we just divide by 1000.

eg. 30 cm^3 = 30 ÷ 100 dm^3 = 0.03 dm^3.

Q1. Convert the following to dm^3.

a) 100 cm^3 b) 7 cm^3 c) 100,000 cm^3

d) 78 cm^3 e) 0.7 cm^3 f) 1.7 x 10^3 cm^3

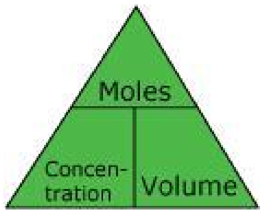
Q2. Convert the following to cm^3.

a) 4 dm^3 b) 0.7 dm^3 c) 100 dm^3

d) 1.45×10^{-3} dm^3 e) 0.01dm^3 f) 2000 dm^3

Finding the concentration.

To find the concentration we use the formula: **Conc. = Moles ÷ Volume**



But IGCSE exams want concentrations in mol/dm^3 and give you the volume in cm^3.

So you must remember to convert first.

Q1) Find the concentration of:

a) A solution of volume 250 cm^3 containing 2.4 moles.

Volume = 250 cm^3 = 250 ÷ 1000 = 0.25 dm^3

Conc = Mol ÷ Vol (dm^3)
2.4 ÷ 0.25 = 9.6 mol/dm^3

Find the concentration of:

b) A solution of volume 500 cm^3 containing 1.2 moles.

c) A solution of volume 200 cm^3 containing 0.2 moles.

d) A solution of volume 750 cm^3 containing 0.02 moles.

e) A solution of volume 750 cm^3 containing 1.5 moles.

f) A solution of volume 2000 cm^3 containing 6 moles.

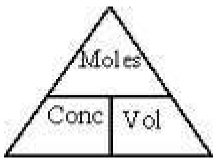
g) A solution of volume 10,000 cm^3 containing 6 moles.

h) A solution of volume 200 cm^3 containing 0.6 moles.

Finding Moles from Concentration

You will be often asked to calculate the number of moles of substance using the formula

Moles = Concentration x Volume (dm^3)



Just remember to change cm^3 to dm^3. But only when it is necessary!

Q1) How many moles in:

a) 2.5 dm^3 of 2 mol/dm^3 $HCl_{(aq)}$

b) 0.5 dm^3 of 1.4 mol/dm^3 $H_2SO_{4(aq)}$

c) 2.5 dm^3 of 0.1 mol/dm^3 $HNO_{3(aq)}$

d) 0.5 dm^3 of 0.4 mol/dm^3 $NaOH_{(aq)}$

e) 0.05 dm^3 of 0.2 mol/dm^3 $NaNO_{3(aq)}$

Q2) How many moles in:

f) 250 dm^3 of 2.4 mol/dm^3 $HCl_{(aq)}$

g) 50 cm^3 of 2.0 mol/dm^3 $H_2SO_{4(aq)}$

h) 200 cm^3 of 0.5 mol/dm^3 $HNO_{3(aq)}$

i) 5 cm^3 of 4.8 mol/dm^3 $NaOH_{(aq)}$

j) 0.15 dm^3 of 0.8 mol/dm^3 $NaNO_{3(aq)}$

Putting it together.

GCSE examiners like to include questions with multiple parts.

Eg 1. A student reacts 6.9g of Sodium with 250 cm^3 of water. Find the concentration of the Sodium Hydroxide formed. $Na + H_2O \rightarrow NaOH + ½ H_2$

i) How many moles of Sodium reacted?

Moles = Mass ÷ M_r = 6.9/23 = 0.3

ii) How many moles of NaOH were made?

From the equation: Na:NaOH 1:1 $\rightarrow$ Moles (NaOH) = 0.3

iii) What is the concentration of NaOH?

Concentration = moles ÷ volume (dm^3) = 0.3 ÷ (250 ÷ 1000) = 1.2 mol/dm^3

Eg 2. A student reacts 5.9g of Potassium Hydroxide (KOH) with 100 cm^3 of Sulphuric Acid. Find the concentration of the Potassium Sulphate (K_2SO_4) formed.

$2 KOH + H_2SO_4 \rightarrow K_2SO_4 + 2H_2O$

a) How many moles of KOH reacted?

Moles = Mass ÷ M_r = 5.9/59 = 0.1

ii) How many moles of K_2SO_4 were made?

From the equation: KOH: K_2SO_4 2:1 $\rightarrow$ Moles (K_2SO_4) = 0.3/2 = 0.15

iii) What is the concentration of K_2SO_4 ?

Concentration = moles ÷ volume (dm^3) = 0.15 ÷ (100 ÷ 1000) = 1.5 mol/dm^3

Q1. If 8g of Calcium reacts with 150cm^3 of Sulphuric Acid, what is the concentration of Calcium Sulphate ($CaSO_4$) solution produced?

$Ca + H_2SO_4 \rightarrow CaSO_4 + H_2$

a) How many moles of Calcium reacted?

b) How many moles of $CaSO_4$ were made?

c) What was the concentration of the $CaSO_4$ solution made?

Q2. If 2.4g of Magnesium reacts with 250cm^3 of Sulphuric Acid, what is the concentration of Magnesium Sulphate (mgSO_4) solution produced?

$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$

a) How many moles of Magnesium reacted?

b) How many moles of $MgSO_4$ were made?

c) What was the concentration of the $MgSO_4$ solution made?

Q3. If 1.4g of Lithium reacts with 50cm^3 of Nitric Acid, what is the concentration of Lithium Nitrate ($LiNO_3$) solution produced?

$$Li + HNO_3 \rightarrow Li\ NO_3 + \frac{1}{2} H_2$$

a) How many moles of Lithium reacted?

b) How many moles of Li NO_3 were made?

c) What was the concentration of the Li NO_3 solution made?

Q4. If 2.1g of Lithium reacts with 200cm^3 of Sulphuric Acid, what is the concentration of Lithium Sulphate (Li_2SO_4) solution produced?

$$2Li + H_2SO_4 \rightarrow Li_2SO_4 + H_2$$

a) How many moles of Lithium reacted?

b) How many moles of Li_2SO_4 were made?

c) What was the concentration of the Li_2SO_4 solution made?

Q5. If 25cm^3 of Sodium Hydroxide (NaOH) with a concentration of 2 mol/dm^3 reacts with 40cm^3 of Sulphuric Acid, what is the concentration of the Sodium Sulphate (Na_2SO_4) solution produced?

$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$

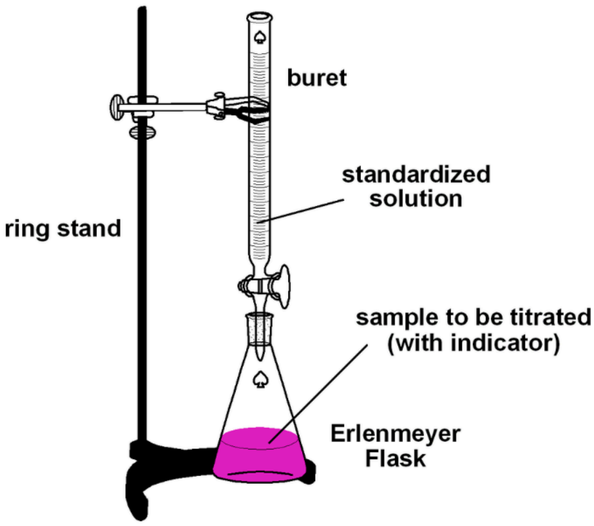
a) How many moles of NaOH reacted?

b) How many moles of Na_2SO_4 were made?

c) What was the concentration of the Na_2SO_4 solution made? (Note, the total volume is 65cm^3)

Titration – finding an unknown concentration from a known concentration

- Using a volumetric pipette, add 25cm^3 NaOH to a conical flask.
- Add some phenolphthalein indicator.
- Fill a burette to almost 0.00cm^3 with 2 mol/dm^3 Hydrochloric acid. Record the volume in the burette.
- Slowly neutralise the NaOH, swirling constantly – your teacher will demonstrate.
- Record the final volume in the burette.
- Calculate the volume of acid used (Titre).
- Repeat using fresh NaOH and clean equipment.
- Repeat a third time if your results are not similar



$$NaOH + HCl \rightarrow NaCl + H_2O$$

	Trial Run	Accurate run 1	Accurate run 2
Final Volume in burette (cm^3)			
Initial Volume in burette (cm^3)			
Titre (cm^3)			

Calculate a ***Mean Titre*** from two concordant results (results with 0.2 cm^3)

Mean Titre = ______

Finding the Concentration of the NaOH

Moles of HCl used = (vol/1000) x Concentration = __________ mol

Reacting Ratio is 1:1

Moles of NaOH neutralised = _________ mol

Concentration of NaOH = ___________ mol/dm^3

Titration Calculation 1

A student neutralises 25cm^3 of NaOH using 2mol/dm^3 HCl

$NaOH + HCl \rightarrow NaCl + H_2O$

Here is her results:

	Trial Run	Accurate run 1	Accurate run 2
Final Volume in burette (cm^3)	**0.00**	**12.40**	**24.10**
Initial Volume in burette (cm^3)	**12.40**	**24.10**	**35.90**
Titre (cm^3)			

Calculate a ***Mean Titre*** from two concordant results (results with 0.2 cm^3)

Mean Titre = ______

Finding the Concentration of the NaOH

Moles of HCl used = (vol/1000) x Concentration = __________ mol

Reacting Ratio is 1:1

Moles of NaOH neutralised = _________ mol

Concentration of NaOH = ___________ mol/dm^3

Titration Calculation 2

A student neutralises 50 cm^3 of NaOH using 2mol/dm^3 HCl

$NaOH + HCl \rightarrow NaCl + H_2O$

Here is her results:

	Trial Run	Accurate run 1	Accurate run 2
Final Volume in burette (cm^3)	**0.00**	**0.00**	**23.10**
Initial Volume in burette (cm^3)	**25.40**	**23.10**	**46.30**
Titre (cm^3)			

Calculate a ***Mean Titre*** from two concordant results (results with 0.2 cm^3)

Mean Titre = ______

<u>Finding the Concentration of the NaOH</u>

Moles of HCl used = (vol/1000) x Concentration = ___________ mol

Reacting Ratio is 1:1

Moles of NaOH neutralised = __________ mol

Concentration of NaOH = ____________ mol/dm^3

Titration Calculation 3

A student neutralises 25 cm^3 of NaOH using 2mol/dm^3 H_2SO_4

$$2\ NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2\ H_2O$$

Here is her results:

	Trial Run	Accurate run 1	Accurate run 2
Final Volume in burette (cm^3)	0.00	15.40	0.00
Initial Volume in burette (cm^3)	15.40	30.10	14.60
Titre (cm^3)			

Calculate a ***Mean Titre*** from two concordant results (results with 0.2 cm^3)

Mean Titre = ______

<u>Finding the Concentration of the NaOH</u>

Moles of H_2SO_4 used = (vol/1000) x Concentration = ___________ mol

Reacting Ratio is 2:1

Moles of NaOH neutralised = __________ mol

Concentration of NaOH = ____________ mol/dm^3

The other Concentration.

School chemistry is mostly about **mol/dm³.**

But industry often uses **g/dm³**.

Converting is easy enough.

Eg. What is the concentration in **mol/dm³** of a solution of NaOH that is 80 **g/dm³**.

1 dm^3 of NaOH contains 80g.

M_r (NaOH) = 40 g/mol

Moles of NaOH = Mass/Mr = 80/40 = 2 mols

So there are 2 mols of NaOH in each dm^3 ------ Concentration = 2 mol/dm^3

Eg. 2 What is the concentration in **g/dm³** of a solution of NaOH that is 0.4 **mol/dm³**.

1 dm^3 of NaOH contains 0.4 mol

M_r (NaOH) = 40 g/mol

Mass of NaOH = Moles x Mr = 0.4 x 40 = 16 g

So there are 16 g of NaOH in each dm^3 ------ Concentration = 16 g/dm^3

Q1. Convert the following to mol/dm^3

a. 21 g/dm^3 HCl

b. 14 g/dm^3 H_2SO_4

c. 32 g/dm^3 $CaCl_2$

d. 4 g/dm^3 KBr

Q2. Convert the following to g/dm^3

a. 2.1 mol/dm^3 HCl

b. 1.4 mol/dm^3 H_2SO_4

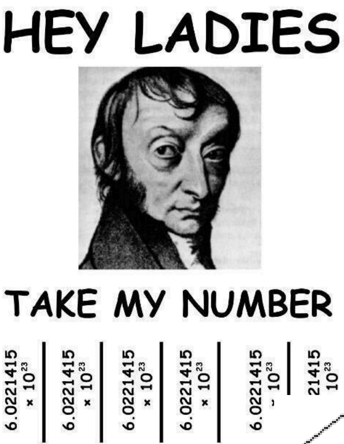
c. 3.2 mol/dm^3 $CaCl_2$

d. 0.4 mol/dm^3 KBr

Molar Gas Volume.

Amedeo Avagadro is not usually considered to have been a good looking man.

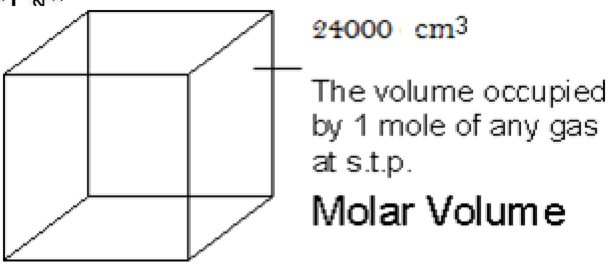
But he's more famous 150 years after his death than he was when he was alive.



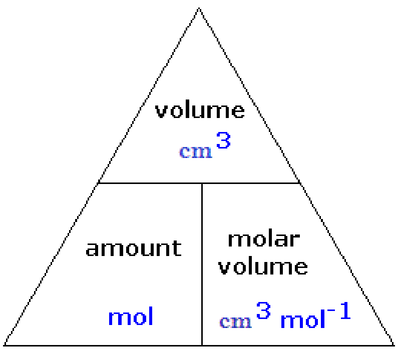
Why?

Because he came up with the idea that equal volumes of any gas contain the same number of particles (as long as they are at the same temperature and pressure)

This is ***Avogadro's Law*** – 1 mole of any gas occupies 24,000 cm^3 (at 25° and 1 atmosphere pressure)



So, it's not difficult to convert gas volumes to moles, or moles to volumes, if you remember:



In other words, **Moles = Volume/24000** or **Volume = Moles x24000**

Molar Gas Volumes

Q1 How many moles of gas are contained in:

a) 24000cm^3 of Oxygen

b) 24000cm^3 of Nitrogen

c) 48000cm^3 of Carbon Dioxide

d) 480cm^3 of Argon

e) 12 cm^3 of Sulphur Dioxide.

Q2. How much space will the following amounts of gas occupy at standard temperature and pressure?

a) 1 mole of Ether

b) 1 mole of Hydrogen

c) 6 moles of Helium

d) 0.6 moles of Chlorine

e) 1.3 moles of Hydrogen Bromide.

Q3. Convert the following masses to moles and then calculate their volume at STP.

a) 22g of CO_2

b) 180g of water vapour (H_2O)

c) 0.2g of Hydrogen (H_2)

Q4. Convert the following volumes of gas to moles, then work out how much they weigh. (Remember Mass = Moles x M_r)

a) 24 cm^3 of CO_2

b) 240 cm^3 of NH_3

c) 12000 cm^3 of SO_2

d) 120 cm^3 of N_2O

e) 120 cm^3 of NO_2

Q5. How much space will the following gases occupy at 25° and 1 atmosphere?

a) 44 g of CO_2

b) 170 g of NH_3

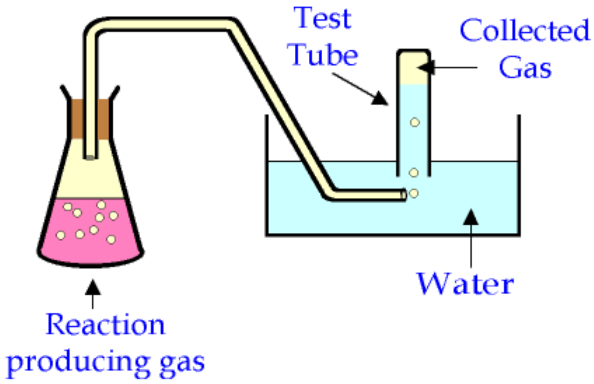
c) 640 g of SO_2

d) 4.2g of N_2O

e) 460g of NO_2

Molar Gas Vol. Experiment

- Set up your equipment as shown in the diagram.
- Place 40 cm^3 of Sulphuric acid in a conical flask.
- Accurately weigh a 3cm strip of Magnesium.
- Drop the Magnesium in the acid and collect the Hydrogen gas in a measuring cylinder.
- If the acid runs out before all the Magnesium is used, add another $10cm^3$



$Mg_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4\,(aq)} + H_{2(g)}$

Results.

Mass of Magnesium used = _________g

Moles of Magnesium = Mass/ 24 = _______

From the equation above, 1 mole of Magnesium should produce 1 mole of Hydrogen gas. So, **Moles of Magnesium used = Moles of Hydrogen made**

Volume of gas expected = Moles of Magnesium used x 24000 = _________cm^3

Actual Volume of gas collected = __________ cm^3

Q1. If the actual volume of gas is different to the volume expected, what could have gone wrong?

Q2. How could we prove that the gas produced was Hydrogen?

15 Hydrobromic acid can be neutralised by adding sodium hydroxide solution.

The equation for the reaction is

$$HBr + NaOH \rightarrow NaBr + H_2O$$

A solution of hydrobromic acid has a concentration of 0.200 mol/dm^3.

(a) Calculate the amount, in moles, of HBr in 20.0 cm^3 of the hydrobromic acid solution.

(2)

amount of HBr = ... mol

(b) Calculate the volume of 0.100 mol/dm^3 sodium hydroxide solution needed to exactly neutralise the hydrobromic acid.

Give the unit.

(2)

volume of sodium hydroxide = ...

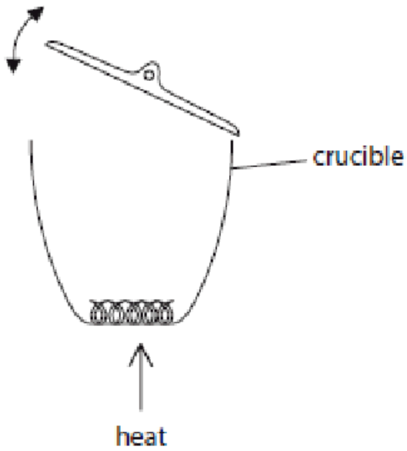
(c) In another neutralisation reaction, a student uses 30.0 cm^3 of 0.200 mol/dm^3 aqueous sodium hydroxide solution.

Calculate the mass of sodium hydroxide contained in this solution.

(2)

mass of sodium hydroxide = ... g

10 A student uses this apparatus to find the mass of magnesium oxide that forms when a strip of magnesium ribbon is burned in air.



This is the student's method.

- weigh an empty crucible and lid
- place a coiled strip of freshly cleaned magnesium ribbon in the crucible
- weigh the crucible, lid and ribbon
- heat the crucible strongly for several minutes, lifting the lid and quickly replacing it several times
- leave the crucible, lid and contents to cool and then reweigh

(a) (i) Describe the appearance of the freshly cleaned magnesium ribbon and the appearance of the magnesium oxide that forms.

(2)

magnesium ribbon ....

....

magnesium oxide ....

....

(ii) Explain why the student lifts the lid and quickly replaces it several times during the experiment.

(2)

lifts lid ....

....

replaces lid ....

....

(b) Calculate the mass of oxygen required to completely react with 0.6 g of magnesium to form magnesium oxide.

The equation for the reaction is

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

(2)

mass of oxygen = ... g

(c) The mass of magnesium oxide that forms is less than expected.

This could be because some of the magnesium reacts with nitrogen to form magnesium nitride, Mg_3N_2

Write a chemical equation for the reaction between magnesium and nitrogen to form magnesium nitride.

(1)

....

(f) Sodium nitrate decomposes when heated, as shown by the equation

$$2NaNO_3 \rightarrow 2NaNO_2 + O_2$$

A 1.70 g sample of sodium nitrate (M_r = 85) was completely decomposed to sodium nitrite ($NaNO_2$) and oxygen.

Calculate the mass of sodium nitrite formed.

(3)

7 Sodium azide (NaN_3) is a stable compound at room temperature but decomposes when heated to 300 °C. The equation for the decomposition is:

$$2NaN_3(s) \rightarrow 2Na(l) + 3N_2(g)$$

Sodium azide is used to produce nitrogen gas to inflate car airbags.

(ii) Use your answer to (a)(i) to calculate the mass, in grams, of sodium azide required to produce 108 dm^3 of nitrogen.

(3)

(iii) Using your answer to (d)(ii), calculate the concentration, in mol/dm^3, of the diluted sulfuric acid.

(2)

Concentration of the diluted sulfuric acid = ... mol/dm^3

(iv) Using your answer to (d)(iii), calculate the concentration, in mol/dm^3, of the **original, concentrated** sulfuric acid.

(1)

Concentration of the original, concentrated acid = ... mol/dm^3

(Total for Question 8 = 11 marks)

12 Lead can be extracted from lead(II) sulfide, PbS, in two stages.

Stage 1: Lead(II) sulfide is heated in air. It reacts with oxygen to produce lead(II) oxide and sulfur dioxide.

Stage 2: The lead(II) oxide is then heated in a blast furnace with coke.

(a) Write a chemical equation for the reaction in **Stage 1**.

(2)

....

(b) The equation for the reaction that occurs when lead(II) oxide is heated with coke in a blast furnace is:

$$2PbO + C \rightarrow 2Pb + CO_2$$

(i) State, with a reason, whether PbO is oxidised or reduced in this reaction.

(1)

....

....

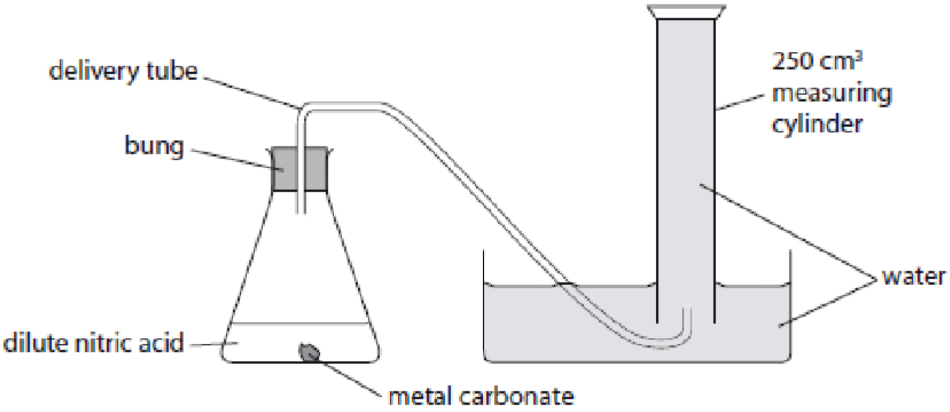
....

(ii) Calculate the minimum mass, in tonnes, of coke needed to react with 44.6 tonnes of lead(II) oxide.
[1 tonne = 10^6 g]

(3)

Mass of coke needed = ... tonnes

7 A student set up this apparatus to measure the volume of carbon dioxide given off when a sample of a carbonate of a Group 2 metal was reacted with dilute nitric acid.



She weighed out some of the carbonate and put it in a conical flask. She then added an excess of dilute nitric acid.

After adding the acid she placed the bung and delivery tube into the conical flask.

She measured the total volume of gas collected at room temperature and pressure (rtp) in the measuring cylinder.

Her results are shown in the table.

Mass of Group 2 carbonate	0.888 g
Volume of gas collected	144 cm^3

The equation for the reaction is

$$XCO_3(s) + 2HNO_3(aq) \rightarrow X(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$

where X is the symbol for the Group 2 metal.

(a) (i) Calculate the amount, in moles, of carbon dioxide gas collected.
(Assume that one mole of gas has a volume of 24 000 cm^3 at rtp)

(2)

Amount of carbon dioxide gas collected = ... mol

(ii) Deduce the amount, in moles, of the carbonate that reacted.

(1)

Amount of carbonate reacted = ... mol

(iii) Using the mass of the carbonate and your answer to (a)(ii), calculate the relative formula mass (M_r) of this carbonate.

Give your answer to the nearest whole number.

(2)

Relative formula mass = ...

(iv) Calculate a value for the relative atomic mass of the Group 2 metal, X, and use the Periodic Table on page 2 to suggest its identity.

(3)

Relative atomic mass of X = ...

Identity of X = ...

(d) Some of the ammonia from the Haber process is made into fertilisers such as ammonium nitrate.

The equation for the formation of ammonium nitrate is

$$NH_3 + HNO_3 \rightarrow NH_4NO_3$$

A manufacturer makes a batch of 40 kg of ammonium nitrate.

(i) Calculate the amount, in moles, of NH_4NO_3 in 40 kg of ammonium nitrate.

(2)

Amount of NH_4NO_3 ....

(ii) Deduce the amount, in moles, of ammonia needed to make this amount of ammonium nitrate.

(1)

Amount of ammonia .... mol

(iii) Calculate the mass, in kg, of ammonia needed.

(2)

Mass of ammonia .... kg

(e) The manufacturer also produces fertilisers that are labelled NPK.

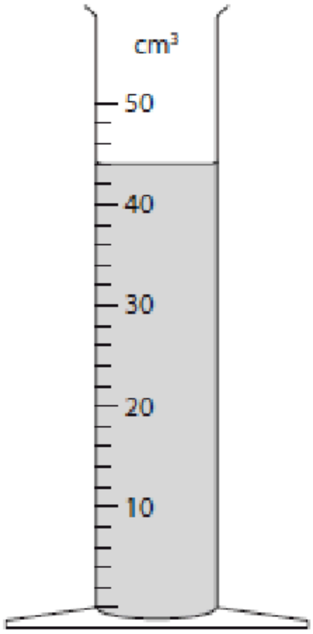
Suggest the names of two elements, other than nitrogen, that are likely to be present in NPK fertilisers.

(1)

.... and ....

(Total for Question 10 = 13 marks)

9 The piece of apparatus shown contains 0.010 mol/dm³ hydrochloric acid.



(a) (i) Give the name of this piece of apparatus.

(1)

....

(ii) What volume of hydrochloric acid is in the apparatus?

(2)

....

(iii) Use your answer in (a)(ii) to calculate the amount, in moles, of hydrochloric acid in the apparatus.

(2)

Amount = ... mol

(b) A student poured a solution containing 0.010 mol of hydrochloric acid into a beaker.

He then added 0.0075 mol of zinc powder and collected the hydrogen given off in a gas syringe.

The equation for the reaction is

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Is the zinc or the hydrochloric acid in excess? Explain your answer.

(2)

....

....

....

(c) The student repeated the experiment with 0.0075 mol of magnesium powder with the same total surface area as the zinc.

The equation for the reaction is

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

(i) What effect would this change have on the rate at which the hydrogen is given off?

(1)

....

....

....

(ii) What effect would this change have on the volume of hydrogen produced?

(1)

....

....

....

(Total for Question 9 = 9 marks)

(e) The equation for the formation of ammonia is

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

(i) Calculate the amount, in moles, of ammonia, that could be formed in the Haber process from 112 kilograms of nitrogen, assuming all the nitrogen is converted into ammonia.

(3)

Amount of ammonia = ... mol

(d) Ethanol can be manufactured by the hydration of ethene. The equation for this reaction is

$$C_2H_4(g) + H_2O(g) \rightarrow C_2H_5OH(g)$$

(i) Identify the catalyst and state the temperature used in this process.

(2)

Catalyst ....

Temperature ....

(ii) A 20 mol sample of ethanol was produced using this reaction.

Deduce the amount, in moles, of ethene needed and the volume, in dm^3, that this amount of ethene would occupy at room temperature and pressure.

Assume that all of the ethene is converted into ethanol and that the molar volume of ethene is 24 dm^3 at rtp.

(3)

Amount of ethene .... mol

Volume of ethene

Volume = ... dm^3

12 Magnesium reacts with dilute hydrochloric acid. The equation for the reaction is

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

(a) 0.0960 g of magnesium was added to 25.0 cm³ of 0.400 mol/dm³ hydrochloric acid.

(i) Calculate the amount, in moles, of magnesium used.

(2)

amount of magnesium = ... mol

(ii) Calculate the amount, in moles, of HCl in the 25.0 cm³ of hydrochloric acid.

(2)

amount of HCl = ... mol

(b) Use your answers from (a) to determine which of the reactants is in excess.

Show your reasoning.

(2)

The reactant in excess is ....

(Total for Question 12 = 6 marks)

(d) The equation for the formation of ammonium nitrate from ammonia is

$$NH_3(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq)$$

25.0 cm^3 of a solution of ammonia of concentration 0.300 mol/dm^3 were reacted with a solution of HNO_3

15.0 cm^3 of HNO_3 were required to exactly neutralise the ammonia solution.

Calculate the concentration, in mol/dm^3, of the HNO_3 solution.

(3)

concentration of HNO_3 = ... mol/dm^3

(Total for Question 13 = 9 marks)

(c) The equation for the fermentation of glucose is

$$C_6H_{12}O_6 \rightarrow 2CH_3CH_2OH + 2CO_2$$

A mass of 3600 kg of glucose was completely fermented.

(i) Calculate the amount, in moles, of glucose that was fermented.
(M_r of glucose = 180)

(2)

amount of glucose = ... mol

(ii) Deduce the amount, in moles, of ethanol produced in this reaction.

(1)

amount of ethanol = ... mol

(iii) Calculate the volume, in dm^3 at rtp, of carbon dioxide produced in this reaction.
(1 mol of carbon dioxide occupies 24 dm^3 at rtp)

(2)

volume of carbon dioxide = ... dm^3

(Total for Question 7 = 11 marks)

15 (a) A student made a solution of sodium hydroxide by dissolving 10.0 g of solid sodium hydroxide in distilled water to make 250 cm^3 of solution.

(i) Calculate the amount, in moles, of NaOH in 10.0 g of sodium hydroxide.

(3)

amount = ... mol

(ii) Calculate the concentration, in mol/dm^3, of this solution of sodium hydroxide.

(2)

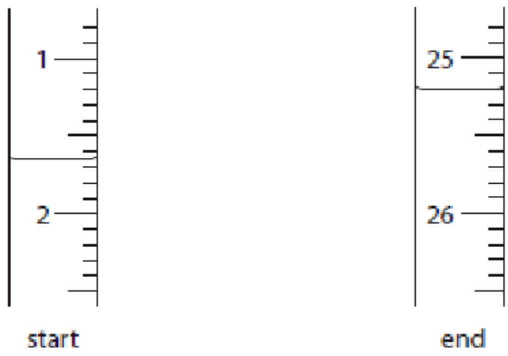
concentration = ... mol/dm^3

(b) (i) The student uses the sodium hydroxide solution to find the concentration of a solution of hydrochloric acid.

He uses this method

- use a pipette to put 25.0 cm^3 of the sodium hydroxide solution into a conical flask
- add a few drops of methyl orange indicator to the solution
- gradually add the hydrochloric acid from a burette until the solution in the flask just changes colour

The diagram shows his burette readings.



Complete the table, giving all values to the nearest 0.05 cm^3.

(3)

burette reading at end in cm^3	
burette reading at start in cm^3	
volume of acid added in cm^3	

(ii) State the colour of the methyl orange at the start and at the end of the experiment.

(2)

colour at start ....

colour at end ....

(iii) Why is a burette used instead of a pipette for adding the acid?

(1)

....

....

(c) Sodium hydroxide reacts with carbon dioxide.

The equation for this reaction is

$$2NaOH + CO_2 \rightarrow Na_2CO_3 + H_2O$$

A solution of sodium hydroxide of concentration 2.00 mol/dm^3 is used.

(i) Calculate the amount, in moles, of sodium hydroxide in 200 cm^3 of this solution.

(2)

amount of sodium hydroxide = ... mol

(ii) Deduce the maximum mass, in grams, of carbon dioxide that can react with this solution of sodium hydroxide.

(2)

mass of carbon dioxide = ... g

(c) In an experiment using a different metal oxide, a mass of 2.8 g of metal is obtained from 3.6 g of the metal oxide.

(i) Calculate the mass of oxygen in the sample of the metal oxide.

(1)

mass of oxygen = ... g

(ii) Calculate the amount, in moles, of oxygen atoms in the sample of the metal oxide.

(2)

amount of oxygen = ... mol

(iii) The formula of the metal oxide is MO, where M is the symbol of the metal.

Deduce the amount, in moles, of M in the sample of the metal oxide.

(1)

amount of M = ... mol

(iv) What is the relative atomic mass of M?

(2)

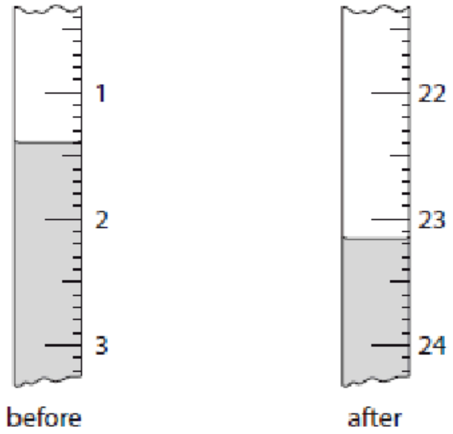
11 A student uses the neutralisation method to make a sample of the soluble salt, sodium sulfate.

The equation for the reaction he uses is

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

(a) He does a titration using these steps to find the ratio of the volumes of reactants needed.

(b) The diagram shows the burette readings in one experiment before and after adding the acid.



Use the readings to calculate the volume of acid added, entering all values to the nearest 0.05 cm^3.

(3)

burette reading after adding acid ....cm^3

burette reading before adding acid ....cm^3

volume of acid added ....cm^3

(c) The student repeats the experiment and records these results.

burette reading in cm^3 after adding acid	25.20	25.05	23.65	23.50
burette reading in cm^3 before adding acid	2.90	3.10	2.55	2.30
volume of acid added in cm^3	22.30	21.95	21.10	21.20
titration results to be used (✓)				

The average (mean) volume of acid added should be calculated using only concordant results (those that differ from each other by 0.20 cm^3 or less).

(i) Identify the concordant results by placing ticks (✓) in the table where appropriate.

(1)

(ii) Use your ticked results to calculate the average volume of acid added.

(2)

(d) The student uses 200 cm^3 of sodium hydroxide solution of concentration 0.300 mol/dm^3 to prepare a sample of sodium sulfate solution.

(i) Calculate the amount, in moles, of NaOH in the sodium hydroxide solution.

(2)

amount of NaOH = ... mol

(ii) Calculate the amount, in moles, of H_2SO_4 needed to neutralise this amount of NaOH.

(1)

amount of H_2SO_4 = ... mol

(iii) Calculate the mass, in grams, of this amount of H_2SO_4

(2)

mass of H_2SO_4 = ... g

(f) Another student used sulfuric acid instead of nitric acid in her experiments. She started with 25.0 cm^3 of sulfuric acid of concentration 0.650 mol/dm^3.

She added 0.500 mol/dm^3 sodium hydroxide solution until the acid was completely neutralised.

The equation for this reaction is

$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$

(i) Calculate the amount, in moles, of sulfuric acid used.

(2)

amount = ... mol

(ii) Calculate the amount, in moles, of sodium hydroxide needed to neutralise this amount of sulfuric acid.

(1)

amount = ... mol

(iii) Calculate the volume, in cm^3, of sodium hydroxide solution needed to neutralise this amount of sulfuric acid.

(2)

volume = ... cm^3

7 Sulfuric acid can be manufactured from sulfur in a four-stage process.

stage 1 sulfur is burned in air to form sulfur dioxide

$$S + O_2 \rightarrow SO_2$$

stage 2 the sulfur dioxide is reacted with more oxygen to form sulfur trioxide

$$SO_2 + \frac{1}{2}O_2 \rightarrow SO_3$$

stage 3 the sulfur trioxide is absorbed in concentrated sulfuric acid to make oleum

$$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$$

stage 4 the oleum is carefully diluted with water to form sulfuric acid

(a) Write a chemical equation for the formation of sulfuric acid from oleum.

(1)

....

(b) A mass of 80 tonnes of sulfur is reacted with oxygen in stage 1.

Calculate the maximum mass, in tonnes, of sulfur trioxide that can be produced in stage 2.

[1 tonne = 1.0×10^6 g]

(3)

(c) Calculate the minimum volume at rtp, in cubic decimetres (dm^3), of oxygen required to completely react with 64 tonnes of sulfur dioxide.

[1 mol of oxygen at rtp has a volume of 24 dm^3]

(2)

volume of oxygen = ... dm^3

8 A student carries out a titration to find the concentration of some dilute sulfuric acid.

She is given

- a supply of the dilute sulfuric acid
- sodium hydroxide solution of concentration 0.150 mol/dm^3
- apparatus suitable for carrying out a titration
- phenolphthalein indicator

She uses this method to do the titration.

step 1 add 25.0 cm^3 of the sodium hydroxide solution to a conical flask

step 2 add 3 drops of phenolphthalein indicator to the conical flask

step 3 fill a burette with the sulfuric acid

step 4 add the sulfuric acid to the conical flask until the phenolphthalein indicator just changes colour

(a) Name the piece of apparatus that the student should use to add the sodium hydroxide solution in step 1.

(1)

....

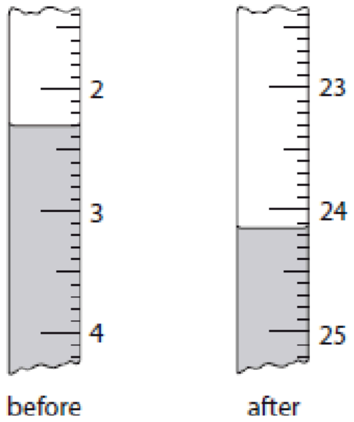
(c) Why is it better to use phenolphthalein indicator rather than universal indicator in this titration?

(1)

....

....

(d) The diagram shows the burette readings in one titration.



Use the readings to complete the table, entering all values to the nearest 0.05 cm^3.

(3)

burette reading in cm^3 after adding acid	
burette reading in cm^3 before adding acid	
volume of acid added in cm^3	

(e) The student repeats the experiment using the same sodium hydroxide solution but another solution of sulfuric acid of a different concentration.

The table shows her results.

burette reading in cm^3 after adding acid	27.65	27.80	27.75	27.40
burette reading in cm^3 before adding acid	0.50	1.50	1.00	1.00
volume of acid added in cm^3	27.15	26.30	26.75	26.40
titration results to be used (✓)				

The average (mean) volume of acid should be calculated using only concordant results.

Concordant results are those volumes that differ from each other by 0.20 cm^3 or less.

(i) Identify the concordant results by placing ticks (✓) in the table where appropriate.
(1)

(ii) Use your ticked results to calculate the average volume of acid added.
(2)

average volume of acid = ... cm^3

(f) The student uses a similar method to find the concentration of a solution of phosphoric acid (H_3PO_4).

The equation for the reaction is

$$3NaOH + H_3PO_4 \rightarrow Na_3PO_4 + 3H_2O$$

The table shows her results.

volume of sodium hydroxide solution added to conical flask	25.0 cm^3
concentration of sodium hydroxide solution	0.180 mol/dm^3
average volume of phosphoric acid solution added from burette	28.30 cm^3

(i) Calculate the amount, in moles, of NaOH in 25.0 cm^3 of the sodium hydroxide solution.

(2)

amount of NaOH = ...mol

(ii) Calculate the amount, in moles, of H_3PO_4 in the phosphoric acid solution.

(1)

amount of H_3PO_4 = ...mol

(iii) Calculate the concentration, in mol/dm^3, of the phosphoric acid.

(2)

concentration of phosphoric acid = ... mol/dm^3

(e) The manufacturer makes a batch of ethanoic acid from methanol and carbon monoxide using this reaction.

$$CH_3OH + CO \rightarrow CH_3COOH$$

He starts with 64 kg of methanol.

Calculate the maximum mass of ethanoic acid he could obtain.

(3)

maximum mass of ethanoic acid = ... kg

15 The formula for hydrated iron(II) sulfate is $FeSO_4.xH_2O$

The value of x is a whole number between 1 and 10. It can be determined by carrying out a titration with 0.0200 mol/dm³ potassium manganate(VII) ($KMnO_4$) solution as follows:

- dissolve a sample of $FeSO_4.xH_2O$ in water to make 250 cm³ of solution
- measure out 25.0 cm³ of this solution into a conical flask
- add the $KMnO_4$ solution using a burette until the end point is reached
- record the volume of solution added
- repeat the titration three more times

The table shows the results.

titration number	1	2	3	4
volume in cm³ of $KMnO_4$ solution added	22.80	22.10	22.50	22.20
concordant titration results (✓)				

(a) Concordant results are those within 0.20 cm³ of each other.

Place ticks (✓) in the table to show the concordant results.

(1)

(b) Using the concordant results, calculate the average (mean) volume of $KMnO_4$ solution added. Give your answer to 2 decimal places.

(2)

average volume added = ... cm³

(c) Which is the most suitable piece of apparatus to measure out 25.0 cm³ of $FeSO_4$ solution?

(1)

- ☐ **A** beaker
- ☐ **B** gas syringe
- ☐ **C** measuring cylinder
- ☐ **D** pipette

(d) These results were obtained in another titration.

mass of $FeSO_4.xH_2O$ in 250 cm^3 of the $FeSO_4$ solution	5.56 g
average volume of $KMnO_4$ solution added to 25.0 cm^3 of solution	20.00 cm^3
concentration of the $KMnO_4$ solution	0.0200 mol/dm^3

(i) Calculate the amount, in moles, of $KMnO_4$ in 20.00 cm^3 of solution.

(2)

amount of $KMnO_4$ = ... mol

(ii) In this reaction one mole of $KMnO_4$ reacts with five moles of $FeSO_4$

Calculate the amount, in moles, of $FeSO_4$ in 25.0 cm^3 of the $FeSO_4$ solution.

(1)

amount of $FeSO_4$ in 25.0 cm^3 = ... mol

(iii) Calculate the amount, in moles, of $FeSO_4$ in 250 cm^3 of this $FeSO_4$ solution.

(1)

amount of $FeSO_4$ in 250 cm^3 = ... mol

(iv) Using your answer from (d)(iii), calculate the mass, in grams, of $FeSO_4$ in the 5.56 g of $FeSO_4.xH_2O$.

[M_r of $FeSO_4$ = 152]

(1)

mass of $FeSO_4$ = ...g

(e) In another experiment it is found that 24.2 g of $FeSO_4.xH_2O$ contains 15.2 g of iron(II) sulfate ($FeSO_4$).

(i) Calculate the mass of water in 24.2 g of $FeSO_4.xH_2O$

(1)

mass of water = ... g

(ii) Calculate the amount, in moles, of H_2O in this mass of water.

(1)

amount of H_2O = ... mol

(iii) Calculate the amount, in moles, of $FeSO_4$ in 15.2 g of iron(II) sulfate.
[M_r of $FeSO_4$ = 152]

(1)

amount of $FeSO_4$ = ... mol

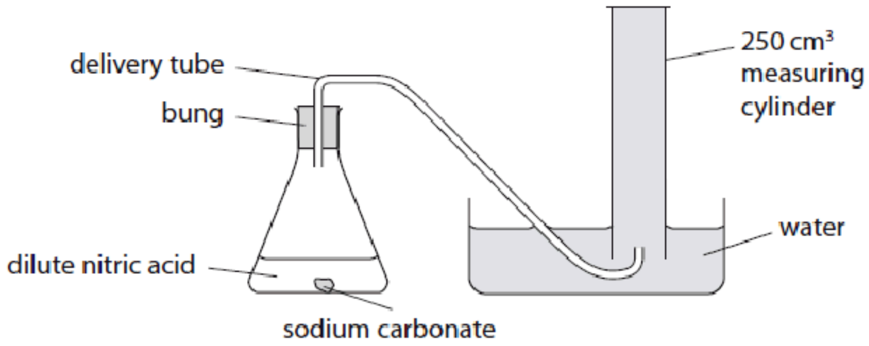
(iv) Using your answers to parts (ii) and (iii), calculate the value of x in $FeSO_4.xH_2O$.

(1)

value of x = ...

(Total for Question 15 = 13 marks)

12 A student uses this apparatus to determine the volume of one mole of carbon dioxide gas.



This is the student's method.

- a solid lump of sodium carbonate of mass 0.53 g is placed into the conical flask
- an excess of dilute nitric acid is added and the bung is put in place
- when all of the sodium carbonate has reacted, the volume of carbon dioxide collected is measured

The equation for the reaction is

$$Na_2CO_3 + 2HNO_3 \rightarrow 2NaNO_3 + H_2O + CO_2$$

(a) (i) Calculate the amount, in moles, of sodium carbonate that reacts.
[M_r: Na_2CO_3 = 106]

(2)

amount of sodium carbonate = ... mol

(ii) The volume of carbon dioxide collected is 110 cm^3.

Use this information and your answer to (a)(i) to calculate the volume, in cm^3, of one mole of carbon dioxide.

(2)

volume of one mole of carbon dioxide = ... cm^3

(b) The equation for the reaction between lithium nitride and water is

$$Li_3N(s) + 3H_2O(l) \rightarrow 3LiOH(aq) + NH_3(g)$$

(ii) A sample of 1.40 g of lithium nitride is added to an excess of water.

Calculate the amount, in moles, of Li_3N in the sample of lithium nitride.

(2)

amount of Li_3N = ... mol

(iii) Calculate the amount, in moles, of LiOH in the lithium hydroxide formed.

(1)

amount of LiOH = ... mol

(iv) Calculate the volume of 0.500 mol/dm^3 sulfuric acid required to neutralise exactly the amount of lithium hydroxide calculated in (b)(ii).

Give an appropriate unit.

The equation for the reaction is

$$2LiOH + H_2SO_4 \rightarrow Li_2SO_4 + 2H_2O$$

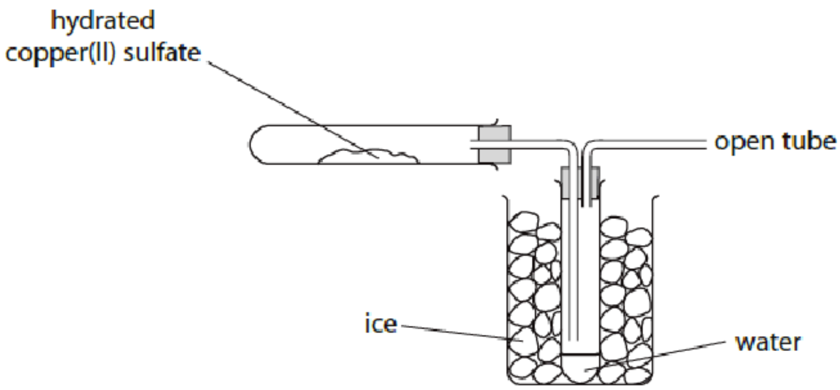
(3)

volume of sulfuric acid = ... unit ....

8 The apparatus in the diagram is used to heat a sample of hydrated copper(II) sulfate crystals, $CuSO_4.5H_2O$

The equation for the reaction that takes place is

$$CuSO_4.5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(l)$$



(a) Draw an arrow on the diagram to show where heat is applied.

(1)

(b) What is the purpose of the ice?

(1)

....

(c) Calculate the maximum mass of water that could be collected when a sample of hydrated copper(II) sulfate of mass 2.50 g is heated.
[M_r of $CuSO_4.5H_2O$ is 250]

(3)

mass of water = ... g

14 Potassium hydrogencarbonate ($KHCO_3$) decomposes on heating.

Three possible equations for the decomposition are

equation 1 $2KHCO_3(s) \rightarrow K_2O(s) + 2CO_2(g) + H_2O(g)$

equation 2 $KHCO_3(s) \rightarrow KOH(s) + CO_2(g)$

equation 3 $2KHCO_3(s) \rightarrow K_2CO_3(s) + CO_2(g) + H_2O(g)$

When 8.00 g of potassium hydrogencarbonate is heated until it is fully decomposed, 5.52 g of solid is formed.

(a) Complete the table by calculating the amount, in moles, of each solid.

(2)

Solid	M_r of solid	Mass of solid in g	Amount of solid in mol
$KHCO_3$	100	8.00	
K_2O	94	5.52	
KOH	56	5.52	
K_2CO_3	138	5.52	

(b) Use the information in the table to explain which equation, 1, 2 or 3, represents the decomposition of potassium hydrogencarbonate.

(2)

8 This is a recipe for making Irish soda bread.

- add 170 g of wholemeal flour, 170 g of plain flour, 10 g of salt and 10.5 g of bicarbonate of soda (sodium hydrogencarbonate, $NaHCO_3$) to a bowl and stir
- pour in 290 cm^3 of buttermilk and stir quickly to form a soft dough
- form the dough into a round ball and slightly flatten it
- cut a cross in the top and bake for 30 minutes in an oven at 200 °C

When sodium hydrogencarbonate is heated, it forms carbon dioxide gas.

$$2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$$

(a) Calculate the mass, in grams, of carbon dioxide that would be produced by completely decomposing 10.5 g of sodium hydrogencarbonate.
[M_r of $NaHCO_3$ = 84]

(2)

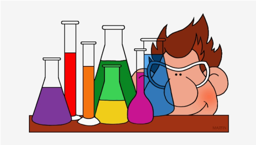
mass of carbon dioxide = ... g

(b) Use your answer from part (a) to calculate the volume, in cm^3, at room temperature and pressure, of carbon dioxide that would be produced by completely decomposing 10.5 g of sodium hydrogencarbonate.

Assume one mole of carbon dioxide has a volume of 24 000 cm^3 at room temperature and pressure.

(2)

volume of carbon dioxide = ... cm^3



9 A student prepares a sample of copper(II) sulfate crystals using this reaction.

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

He obtains the crystals from the solution formed.

(a) He records this information about the reactants he uses.

mass of copper(II) oxide = 6.3 g

volume of sulfuric acid = $52\ cm^3$

concentration of sulfuric acid = $1.1\ mol/dm^3$

(i) Calculate the amount, in moles, of copper(II) oxide used.

(2)

amount of copper(II) oxide = ... mol

(ii) Calculate the amount, in moles, of sulfuric acid used.

(2)

amount of sulfuric acid = ... mol

(iii) Why is it important for the amount of copper(II) oxide to be greater than the amount of sulfuric acid?

(1)

....

....

(iv) Draw a diagram of the apparatus that the student should use to remove the excess copper(II) oxide from the reaction mixture.

(1)

(b) In a similar preparation the student uses 0.12 mol of copper(II) oxide to obtain crystals of copper(II) sulfate, $CuSO_4.5H_2O$

Calculate the maximum mass of $CuSO_4.5H_2O$ that he could obtain using this preparation.

(2)

maximum mass = ... g

7 Magnesium chloride can be made by reacting excess magnesium carbonate with dilute hydrochloric acid.

The equation for the reaction is

$$MgCO_3 + 2HCl \rightarrow MgCl_2 + H_2O + CO_2$$

(a) (i) In one experiment, a sample of 0.050 mol of $MgCO_3$ is added to 0.080 mol of HCl.

Show, by calculation, that the $MgCO_3$ is in excess.

(2)

(ii) Calculate the maximum volume, in cm^3, of carbon dioxide, measured at room temperature and pressure, that would be obtained when 0.080 mol of HCl react completely with $MgCO_3$.

[One mole of any gas occupies 24 000 cm^3 at room temperature and pressure.]

(2)

maximum volume of carbon dioxide = ... cm^3

(b) In another experiment 0.050 mol of $MgCO_3$ reacts with excess HCl.

A yield of 5.5 g of $MgCl_2.6H_2O$ is obtained.

(i) Calculate the percentage yield of $MgCl_2.6H_2O$

(2)

percentage yield = ... %

(ii) Suggest why the percentage yield is less than 100%.

(1)

....

....

....

I have completed the questions in this section	
I have read the relevant sections of the College Website	
I have made some revision material (mind-map, key-words & definitions etc)	
Prep Grade	
Test Grade	

Target	Pupil Signature

10.4 - Percentage Yield & Atom Economy

Syllabus

- Even though no atoms are gained or lost in a chemical reaction, it is not always possible to obtain the calculated amount of a product because:

1. the reaction ***may not go to completion*** because it is ***reversible***
2. some of the product may be ***lost*** when it is ***separated*** from the reaction mixture
3. some of the reactants may ***react in ways different to the expected reaction***.

- The amount of a product obtained is known as the ***yield***.
- When compared with the maximum theoretical amount as a percentage, it is called the percentage yield.

% Yield = (Mass of product actually made / Maximum theoretical mass of product) × 100

- Students should be able to:

1. calculate the percentage yield of a product from the actual yield of a reaction
2. (HT only) calculate the theoretical mass of a product from a given mass of reactant and the balanced equation for the reaction.

- The ***atom economy*** (atom utilisation) is a measure of the amount of starting materials that end up as ***useful*** products.
- It is important for ***sustainable*** development and for economic reasons to use reactions with high atom economy.
- The percentage atom economy of a reaction is calculated using the balanced equation for the reaction as follows:

(M_r of desired product from equation / Sum of M_rs of all reactants from equation) × 100

- Students should be able to:

1. calculate the atom economy of a reaction to form a desired product from the balanced equation
2. (HT only) explain why a particular reaction pathway is chosen to produce a specified product given appropriate data such as atom economy (if not calculated), yield, rate, equilibrium position and usefulness of by-products.

Percentage Yield

When you make a new substance, it's possible to calculate how many moles you ***should*** make as long as you know the equation and the number of moles of reactants used.

But you never quite make as much as you expect because some substances evaporate, solids get stuck in filters, more than one reaction happens or some of the product remains in the flask as drops.

It's often possible to make the same substance in more than one way. So we need to know which method is most efficient.

To do this we calculate the ***Percentage Yield***.

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Examples

1. You expect to make 2.5g but make only 2.2g. What is the percentage yield?

   Percentage Yield = 2.2/2.5 x 100 = 88%

2. A scientist makes 0.8 moles of Ammonia when he expected 1.2 moles. Another makes 3.4 moles when she expected 3.9 moles. Which method is most efficient?

   Method 1 %Yield = 0.8/1.2 x 100 = 67%

   Method 2 % Yield = 3.4/3.9 x 100 = 87%

   Method 2 is the most efficient method.

3. You calculate that you could make 120g of product but the method has a maximum yield of 65%. How much will you actually make?

   65 = Actual yield/120 x 100

   Actual yield/120 = 65/100

   Actual Yield = 65/100 x 120 = 78g

Percentage Yield Questions

1. Complete the table below

Actual Yield (g)	Predicted Yield (g)	Percentage Yield (g)
50	100	
25	50	
8	40	
80	120	
33	100	

2. If the theoretical yield of Carbon Dioxide from burning an amount of Methane is 75g and the actual yield is 50g, what is the percentage yield?

3. 46g of Sodium is dropped into water. It produces 70g of Sodium Hydroxide when you had expected 80g. What is the percentage yield?

Percentage Yield Questions

4. 250g of Calcium Carbonate is heated to produce 100g of calcium oxide instead of the 140g predicted. What is the percentage yield?

5. 125g of Zinc Carbonate ($ZnCO_3$) reacts with Sulphuric Acid to produce 100g of Zinc Sulphate ($ZnSO_4$). It had been predicted that 161 g would be made. What is the percentage yield?

6. 96g of Methane (CH_4) reacts with Oxygen to produce 88g of Carbon Dioxide rather than the expected 264g. What is the percentage yield?

Harder examples

Most GCSE questions are of the variety above.

But you might have to work out the expected yield yourself on Paper 2.

A student drops 1.4g of Lithium in Water. She finds that 3.6g of Lithium Hydroxide is made. Find the percentage yield.

$$Li_{(s)} + H_2O_{(l)} \rightarrow LiOH_{(aq)} + \frac{1}{2} H_{2(g)}$$

First we need to find out how many moles of Li reacted.

Moles = Mass ÷ M_r = 1.4 / 7 = 0.2 moles

Then we can say how many moles of LiOH were expected.

Ratio Li : LiOH = 1:1 so expect 0.2 moles

Then we can say what mass of LiOH was expected.

This lesson involves:

*Logical Thinking
*Precision
*Multi-step problem-solving

Mass = Moles x M_r = 0.2 x 24 = 4.8g

Finally we can work out the percentage yield

% yield = (actual ÷ expected) x 100

= 3.6/4.8 x100 = 75%

A student adds 3.4g of Ammonia (NH_3) to Sulphuric Acid. He finds that 12.6g of Ammonium Sulphate $(NH_4)_2SO_4$ is made. Find the percentage yield.

$$2\ NH_{3(g)} + H_2SO_{4(aq)} \rightarrow (NH_4)_2SO_{4\ (aq)}$$

First we need to find out how many moles of Ammonia reacted.

Moles = Mass ÷ M_r = 3.4 / 17 = 0.2 moles

Then we can say how many moles of $(NH_4)_2SO_4$ were expected.

Ratio NH_3: $(NH_4)_2SO_4$ = 2:1

so expect 0.1 moles

Then we can say what mass of $(NH_4)_2SO_4$ was expected.

Mass = Moles x M_r = 0.1 x 132 = 13.2g

Finally we can work out the percentage yield

% yield = (actual ÷ expected) x 100

= 12.6/13.2 x100 = 95.5

More Questions.

1. In an experiment, 11.5g of sodium is reacted with an excess of chlorine. The mass of sodium chloride made is 23.4g.

   $Na + ½\ Cl_2 \rightarrow NaCl$

   a) What is the maximum amount of product that could be made?

b) What is the percentage yield?

2. 24g of Magnesium reacts with Chlorine to make what mass of Magnesium Chloride. $Mg + Cl_2 \rightarrow MgCl_2$

a) What is the maximum amount of $MgCl_2$ that could be made?

b) What is the percentage yield if 40g of $MgCl_2$ was made?

Yet More Questions.

3. 46g of Sodium reacts with Oxygen to produce what mass of Sodium Oxide. $4Na + O_2 \rightarrow 2Na_2O$

a) What is the maximum amount of Na_2O that could be made?

b) What is the percentage yield if 32g of Na_2O was made?

3. 28g of Lithium reacts with Hydrochloric acid to produce what mass of Lithium Chloride. $2Li + 2HCl$ □ $2LiCl + H_2$

a) What is the maximum amount of LiCl that could be made?

b) What is the percentage yield if 100g of LiCl was made?

Atom Economy

Percentage yield is concerned with how ***efficient*** a reaction is at making what it's supposed to make.

Atom economy isn't about efficiency. It's about whether the process makes sense ***ecologically***.

It's possible to have a high yield and a low atom economy, or vice versa – they're just not related.

In theory, the ***more waste*** a process makes the ***less sustainable*** it is.

Atom economy quantifies this as…

Atom economy = (M$_r$ of desired product / Sum of M$_r$s of all reactants) × 100

Eg One way of making Iron is.. $2Fe_2O_3 + 3C \longrightarrow 4Fe + 3CO_2$

This makes the CO2 a waste product.

- **PRODUCT:** The A_r of Iron is 56 g/mol. 4 x 56 = 224 g/mol
- **REACTANT #1:** The M_r of Fe_2O_3 = 160 g/mol 2 x 160 = 320 g/mol
- **REACTANT #2:** The A_r of C = 12 g/mol 3 x 12 = 36 g/mol
- **COMBINED REACTANTS:** 320 + 36 = 356 g/mol
- ATOM ECONOMY : 224/356 x100 = 62.9%

Q1, Calculate the atom economy in $2Na_3PO_4 + 3MgCl_2 \longrightarrow 6NaCl + Mg_3(PO_4)_2$

if the desired product is Magnesium Phosphate.

b) If we could also sell the NaCl, why would the Atom Economy be irrelevant?

Q2. $N_2 + 3H_2 \longrightarrow 2NH_3$

Why isn't it necessary to carry out a calculation to be sure that the atom economy of this process is 100%?

Q3. a.What is the atom economy for producing Aluminium Sulphate using the process below?

$$2Al + 3H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3H_2$$

b. If this process were used to make Hydrogen what would the atom economy be?

This lesson involves ...

- Logical Thinking
- Precision
- Multi-step problem-solving

Q4. What's the atom economy of the following way of making CCl_4?

$$CH_4 + 4Cl_2 \longrightarrow CCl_4 + 4HCl$$

(c) Tungsten can also be obtained by reacting tungsten fluoride with hydrogen.

The equation for this reaction is

$$WF_6 + 3H_2 \rightarrow W + 6HF$$

(i) In an experiment, a chemist used 59.6 g of tungsten fluoride.

What is the maximum mass of tungsten he could obtain from 59.6 g of tungsten fluoride?

Relative formula mass of tungsten fluoride = 298

(2)

maximum mass = ... g

(ii) Starting with a different mass of tungsten fluoride, he calculates that the mass of tungsten formed should be 52.0 g. In his experiment he actually obtains 47.5 g of tungsten.

What is the percentage yield of tungsten in this experiment?

(2)

percentage yield = ... %