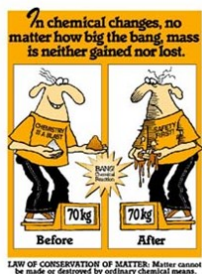


Home learning activities

Subject
Science
Year Group
Year 9
Unit of work / Knowledge organiser
Quantitative Chemistry – 2
Activities
<ul style="list-style-type: none"> • Complete the weekly 'Knowledge Check' through 'GCSEPod'. • Watch all 'GCSEPod' clips on the 'Quantitative Chemistry' Unit. • Complete the 'GCSEPod' Questions assigned for this Unit of work and any additional assignments which have been set by your teacher.
Where do you complete the work?
Use computer/phone for 'GCSEPod' or 'Seneca' and study materials.
What to do if you finish the work? (Extension activity)
<ul style="list-style-type: none"> • Sign up for 'Seneca Learning' using the 'Sign Up Guide' sheet and the special passcode: j5v9tvzq48. Complete the assignments which have been set.
These websites might help:
<ul style="list-style-type: none"> • BBC Bitesize -> Secondary -> GCSE -> Combined Science -> AQA Trilogy -> Chemistry -> Quantitative Chemistry • www.freesciencelessons.co.uk -> GCSE Videos -> Chemistry Paper 1 -> Quantitative Chemistry
If you are struggling with your work or if you have finished.
Please email your classroom teacher directly using the email list found in the Home Learning section of the website.

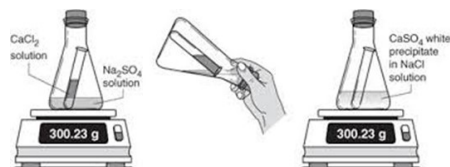
Year 9—Quantitative Chemistry

1. Conservations of mass.



The law of conservation of mass states that no atoms are lost or gained during a chemical reaction so the mass of the products equals the mass of the reactants.

Proving the conversion of mass:



Calcium chloride + sodium sulfate → calcium sulfate + sodium chloride



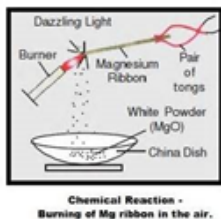
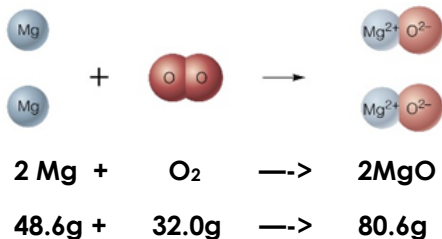
To check conservation of atoms:

Reactants use: 1 x Ca, 2 x Cl, 2 x Na, 1 x SO₄
 Products makes: 1 x Ca, 2 x Cl, 2 x Na, 1 x SO₄

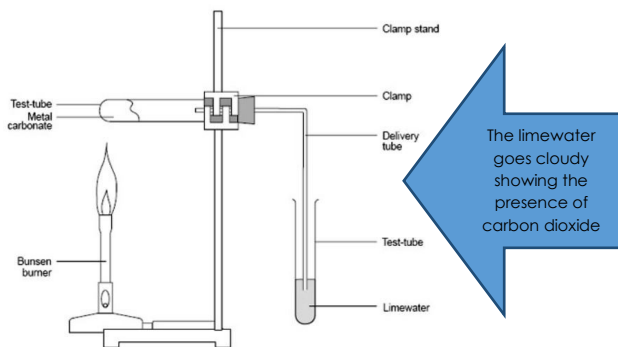
Some reactions may appear to involve a change in mass but this is explained because a reactant or product is a **gas** and its mass has not been taken into account.

The conservation of mass is explained using the balanced symbol equations:

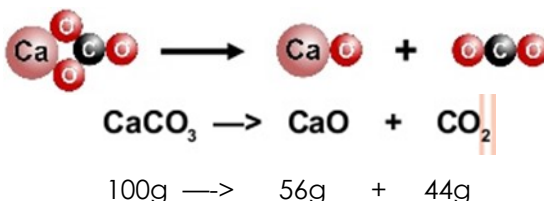
When a metal reacts with oxygen the mass of the oxide produced is **greater** than the mass of the metal



2. In thermal decomposition of metal carbonates carbon dioxide is produced and escapes into the atmosphere leaving the metal oxide as the only solid product.



The limewater goes cloudy showing the presence of carbon dioxide



3. Relative Atomic Mass

The **Relative Atomic Mass (RAM or Ar)** is calculated in comparison to CARBON 12. It is the sum of the protons and neutrons in the nucleus.

Isotope – This is an element with the **same** number of protons – **but a different** number of neutrons in its nucleus.

If you look at chlorine on the periodic table its RAM is **35.5** this is because it exists as the two isotopes of:



4. To calculate the RAM you need to know:

The abundance (amount) of each isotope;
 The RAM of each isotope.
 If the relative abundances are 75% of Cl₃₅ and 25% of Cl₃₇.
 The equation can then be used:

$$\frac{(\% \text{ of Cl}_{35} \times \text{RAM of Cl}_{35}) + (\% \text{ of Cl}_{37} \times \text{RAM of Cl}_{37})}{100}$$

$$(75 \times 35) + (25 \times 37) = 2625 + 925 = \frac{3550}{100} = \mathbf{35.5}$$

The **Relative Molecular Mass (RMM or Mr)** is calculated using the RAM/Ar of the atoms making up the molecule.

$$\text{Mr of CaCO}_3 = \text{Ar of Ca} + \text{Ar of C} + 3\text{Ar of O}$$

$$= 40 + 12 + 3 \times 16 = 48$$

$$= 100$$

Remember – in a balanced equation, the sum of the Mr of the reactants **equals** the sum of the Mr of the products – this shows conservation of mass.

5. Moles and reacting mass (HT ONLY)

Avogadro's number **6.02 x 10²³ atoms** is the number of atoms in the relative atomic mass of an atom.

So, a 24 g piece of magnesium contains 6.02 x 10²³ atoms.

This also refers to **one mole of a substance**.
 The relative molecular mass of a compound also refers to Avogadro's number.

Year 9—Quantitative Chemistry

6. Calculating molar mass

Unit is g/mol or g mol^{-1} .

The mass of one mole of a substance is calculated by adding up the relative atomic masses of the atoms in the formula.

Eg for H_2O

$\text{H} + \text{H} + \text{O} = \text{H}_2\text{O}$

$1 + 1 + 16 = 18\text{g}$

One mole of water = 18g/mol

Eg for formula containing brackets, these must be considered in the calculation:

$\text{Mg}(\text{NO}_3)_2$ atoms in the brackets must be multiplied by 2

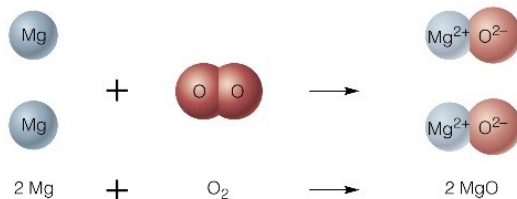
$\text{Mg} = 24, \text{N} = 14, \text{O} = 16$

So $\rightarrow 24 + (2 \times 14) + (2 \times (3 \times 16)) = 148 \text{ g/mol}$.

7. Amounts of substances in equations (HT ONLY)

Using balanced symbol equations the masses of reactants and products can be calculated.

For the reaction:



The equation also shows us that 2 moles of magnesium will react with 1 mole of oxygen to produce 2 moles of magnesium oxide.

$\text{Mg} = 24 \quad \text{O} = 16$

So \rightarrow

$48\text{g} + 32\text{g} \rightarrow 80\text{g}$

To calculate different masses the equation is needed:

$\frac{\text{molar mass of substance A}}{\text{mass of A}} = \frac{\text{molar mass of substance B}}{\text{mass of B}}$

Worked example – calculate the mass of MgO made from 6.0g of Mg .

Rearrange the equation to become:

Mass of B = mass A x $\frac{\text{molar mass of B}}{\text{Molar mass of A}}$

Substitute in numbers

Mass of $\text{MgO} = 6.0 \times \frac{80}{48}$

Calculate = **10g** Don't forget the units

8. Using moles to balance equations

Number or moles = $\frac{\text{mass of the chemical}}{\text{Molar mass}}$

Worked example – Aluminium oxide Al_2O_3 produces aluminium, Al and oxygen O_2 .

If 204g of Al_2O_3 produces 108g of Al work out the number of moles of Al_2O_3 , Al and O_2 involved hence write out the full balanced equation.

Use the equation:

Number or moles = $\frac{\text{mass of the chemical}}{\text{Molar mass}}$

$\text{Al}_2\text{O}_3 \rightarrow \text{Al} + \text{O}_2$

$204\text{g} \rightarrow 108\text{g} + \text{??g} (204 - 108 = 96\text{g})$

Number of moles of aluminium = $\frac{204}{102} = 2$
oxide

Number of moles of aluminium = $\frac{108}{27} = 4$

Number of moles of oxygen = $\frac{96}{32} = 3$

Balanced equation:

$2\text{Al}_2\text{O}_3 \rightarrow 4\text{Al} + 3\text{O}_2$

9. Concentration of solutions

CONCENTRATION – the amount of a chemical dissolved in a certain volume of a solution. It is calculated using:

Concentration = $\frac{\text{mass of solute}}{\text{volume}}$

The units for volume is dm^3 this is equal to 1000cm^3 to convert cm^3 to dm^3 **divide by 1000**.

Worked example:

A solution has a concentration of 4.2g/dm^3 . Calculate the mass of solute dissolved in 250cm^3 of solution.

1. Use the equation. Substitute in known values.

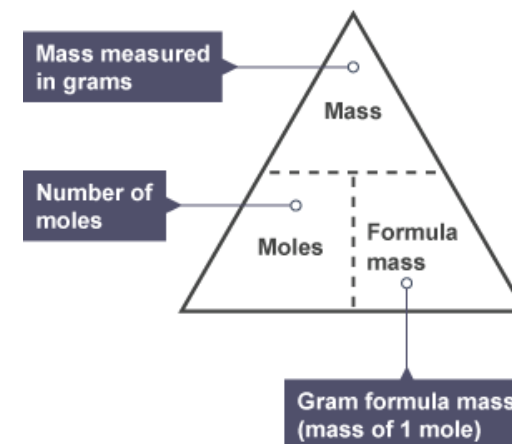
Concentration = $\frac{\text{mass of solute}}{\text{Volume}}$

$4.2 \text{ g/dm}^3 = \frac{??}{(250/1000)}$

2. Rearrange

Mass of solute = $4.2 \times (250/1000)$

= **1.05 g** Don't forget the units



10. Using moles to calculate concentration of solutions

You can work out the concentration of a solution using this equation:

$$\text{concentration in mol dm}^{-3} = \frac{\text{number of moles of solute}}{\text{volume of solution in dm}^3}$$

Worked example:

25.00cm³ of sodium hydroxide solution was titrated against 0.10mol dm⁻³ hydrochloric acid. An average of 20.00cm³ of the acid was needed to react completely. What is the concentration of the sodium hydroxide solution?

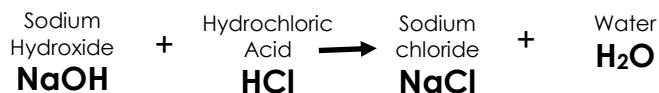
Step 1:

$$\begin{array}{l} \text{Number of moles of} \\ \text{hydrochloric} \\ \text{acid} \end{array} = \begin{array}{l} \text{Concentration of} \\ \text{hydrochloric acid} \\ \text{(mol dm}^{-3}\text{)} \end{array} \times \begin{array}{l} \text{Volume} \\ \text{used (dm}^{-3}\text{)} \end{array}$$

$$\begin{array}{l} \text{So, number of} \\ \text{moles of} \\ \text{hydrochloric} \\ \text{acid} \end{array} = 0.1 \times \frac{20.0}{1000} = 0.002 \text{ mol}$$

Step 2: Write the balanced equation for the reaction and use this to work out how many moles of sodium hydroxide reacted with this number of moles of acid:

The equation shows that 1mol of hydrochloric acid reacts



with 1mol of sodium hydroxide. So, 0.002mol of hydrochloric acid reacts with 0.002mol of sodium hydroxide

Step 3:

$$\begin{array}{l} \text{Concentration of} \\ \text{sodium hydroxide} \end{array} = \frac{\text{Moles of sodium hydroxide}}{\text{Volume of sodium hydroxide solution (dm}^{-3}\text{)}} \\ = \frac{0.002}{0.025} = 0.08 \text{ mol dm}^{-3}$$

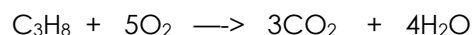
11. Amounts of substances in volumes of gases

We find that is the **formula mass** of a gas is measured in grams and the **volume** it occupies is measured (at **room temperature and pressure**) then a gas occupies **24dm³**.

$$\text{Number of moles of gas} = \frac{\text{volume of gas in dm}^3}{24}$$

Volumes from balanced equations

If propane reacts with oxygen the equation is:



If 24dm³ of C₃H₈ is used then 5 x 24dm³ of O₂ is needed (120 dm³)

If 2.4dm³ of C₃H₈ is used then 3 x 2.4dm³ of CO₂ is made (7.2 dm³)

This is calculated using the equation:

$$\frac{\text{Volume 1 (v1)}}{\text{Moles 1 (n1)}} = \frac{\text{volume 2 (v2)}}{\text{Moles 2 (n2)}}$$

Rearranged to become:

$$\frac{v1 \times n2}{n1} = v2$$

worked example:

Calculate the volume of water vapour made at rtp when 1.5 dm³ of C₃H₈ reacts with oxygen in the equation above.

1.5 dm³ C₃H₈ is used to produce 4x1.5dm³ of water = **6dm³**

12. Percentage yield

The more reactant used the more product is made.

There may not be 100% of the product because:

Loss of filtration – small amounts stay on the filters

Loss in evaporation – some chemicals evaporate

Loss in transferring liquids – it sticks to the glass vessels

The **percentage yield** of a reaction is a way of comparing the mass of product made (the **actual yield**) to the mass we expect to make (the **theoretical mass**).

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Worked example:

For example, in the reaction between hydrogen and oxygen the theoretical maximum yield of water which could be produced from reacting 32 g of oxygen is 36 g. The actual yield obtained was 28 g. So:

Actual yield (in grams)	28
Theoretical maximum yield (in grams)	36
Percentage yield = $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100$	$\frac{28}{36} \times 100 = 78.8\%$

13. Atom economy:

Atom economy is a way of measuring how many of the starting materials end up as **useful products** in a chemical reaction. It is measured in terms of the atoms taking part in the reaction.

It is calculated using:

$$\% \text{ atom economy} = \frac{\text{Mr desired product}}{\text{Sum of Mr of all reactants}} \times 100$$

A company makes magnesium sulfate MgSO_4 for use as bath salts. They need to find the best method.

A_r : Mg = 24, O = 16, H = 1, S = 32

M_r of MgSO_4 (desired) = $24 + 32 + (16 \times 4) = 24 + 32 + 64 = 120$

Method 1:	Method 2:
$\text{MgO} + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2\text{O}$	$\text{MgCO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2\text{O} + \text{CO}_2$
M_r reactants:	M_r reactants:
$24 + 16 + 2 + 32 + 64 = 138$	$24 + 12 + 48 + 2 + 32 + 64 = 182$
% atom economy = $120/138 \times 100 = 86.9\%$	% atom economy = $120/182 \times 100 = 65.9\%$

Which method is best? The higher the atom economy, the fewer atoms are in the wasted product, so the first method is a less wasteful process.

Calculating theoretical yields

The reactant used to calculate the theoretical maximum should be the limiting factor of the reaction

14. Choosing reactions pathways

All chemicals are produced following an extensive period of research and development. Chemicals made in the laboratory need to be "scaled up" to be manufactured on the plant.

To make a process viable industry tries:

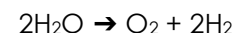
to find suitable conditions – compromise between rate and equilibrium

to find a suitable catalyst – increases rate and cost effective as not used up in the process.



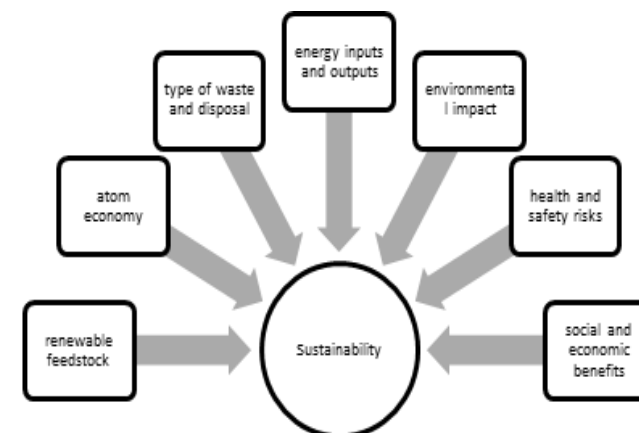
The hydrogen fuel cell car:

Looking at by-products, some reactions can give a low atom economy, e.g., hydrogen for vehicles made from water is:



Using the atom economy formula we find this atom economy is 12.5%. However, if oxygen were the desired product, this reaction would have an atom economy of 87.5%.

15. Considerations for reaction pathways – extension



'Seneca Learning' Sign-Up Guide

Passcode: j5v9tvzq48

Step 1: Open an internet browser - *Any browser except Internet Explorer will work.*

Step 2: Go to SenecaLearning.com

Step 3: Click on "Get Started" or "Sign Up"

Step 4: Create your account - *If you don't know your parent email, then type: N/A.*

Step 5: Click on "Classes & Assignments" - *You'll find this in the top menu.*

Step 6: Click on "Join Class" - *It's the green button in the top right corner.*

Step 7: Type the code from your teacher - *If you received a link instead, then open the link.*