SC9: Quantitative chemistry

Lesson sequence

1. Formula masses
2. Calculating empirical formulae
3. Conservation of mass
4. Calculating reacting masses
5. Moles (HT)
6. Stoichiometry of reactions (HT)

1. Formula masses	
Molecular formula	Gives the number of atoms of each element present in a molecule.
Empirical formula	Gives the number of atoms of each element present in a compound as the simplest whole number ratio.
Converting molecular to empirical formulae	Divide the number of each atom by the highest common factor of all of the atoms.
Molecular to empirical formula examples	C_2H_4 → CH_2 (divided by 2) $C_6H_{12}O_6$ → CH_2O (divided by 6) H_2O → H_2O (divided by 1)
Relative atomic mass, A_r	The mass of an atom relative to 1/12th the mass of carbon-12. No units.
Relative formula mass, M_r	The mass of one unit of a formula, found by adding the relative atomic masses of all of the atoms in it.

Worked examples W1

Calculate the M_r of carbon dioxide (CO_2).

$= A_r(C) + (2 \times A_r(O))$
$= 12 \quad + (2 \times 16)$

So, M_r of CO_2 = 44

Relative Formula Mass – RFM

The Relative Formula Mass (M_r) of a compound is the sum of the relative atomic masses of all its elements added together.

In order to calculate the RFM of a compound you must know the formula and the RAM's of each of the atoms involved (H = 1, O = 16).

Example: Find the M_r of water, H_2O

Step 1: Write the formula — H_2O = H + H + O

Step 2: Substitute the A_r's — 1 + 1 + 16

Step 3: Add them up to get the M_r — 1 + 1 + 16 = 18

H = 1 Hydrogen	C = 12 Carbon	N = 14 Nitrogen	O = 16 Oxygen	Na = 23 Sodium	Mg = 24 Magnesium	Al = 27 Aluminium
P = 31 Phosphorous	S = 32 Sulphur	Fe = 56 Iron	Cu = 63.5 Copper	Cl = 35.5 Chlorine	Ca = 40 Calcium	

Using the method shown and the A_r's above calculate the M_r's for the following:

Sodium chloride – $NaCl$
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Methane – CH_4
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Sodium hydroxide – $NaOH$
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Aluminium oxide – Al_2O_3
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Copper sulphate – $CuSO_4$
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Ethane – C_2H_6
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Ethene – C_2H_4
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Calcium hydroxide – $Ca(OH)_2$
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Iron sulphate – $Fe_2(SO_4)_3$
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Ammonium chloride – NH_4Cl
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Ammonium sulphate - $(NH_4)_2SO_4$
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2. Calculating empirical formulae	
To calculate empirical formulae from experimental data	- Write each element's symbol with a ratio (:) symbol between - Write out the amount of each element from the questions - Divide each amount by the A_r of the element - Divide each answer by the smallest answer to get a ratio - Write the empirical formula
To find a molecular formula from an empirical formula	- Calculate M_r for the empirical formula - Divide the M_r of the molecular formula by this number - Multiply the empirical formula by your answer

Symbol for element	Ca	Cl
Mass (g)	10.0	17.8
Relative atomic mass, A_r	40	35.5
Divide the mass of each element by its relative atomic mass	$\frac{10.0}{40} = 0.25$	$\frac{17.8}{35.5} = 0.5$
Divide the answers by the smallest number to find the simplest ratio	$\frac{0.25}{0.25} = 1$	$\frac{0.5}{0.25} = 2$
Empirical formula	$CaCl_2$	

D To calculate an empirical formula, each element needs its own column of working.

Empirical formula

Empirical formula – is the simplest formula which represents the ratio of atoms in a compound. There is one simple rule to follow: always divide the data you are given by the A_r of the element. Then simplify the ratio to give you the simplest formula.

Example: Find the empirical formulae of an oxide of hydrogen, produced by reacting 1g of hydrogen with 8g of oxygen

Step 1: Write down the relative atomic masses of the elements involved - A_r of H = 1 and A_r of O = 16

Step 2: Divide the masses given in the question by the A_r's of the elements –

H = 1/1 = 1 : O = 8/16 = 0.5

Step 3: Identify the ratio of atoms in the compound and simplify it, the easiest way to do this is to divide both sides by the smallest number and then make sure both sides are whole numbers–

1/0.5 : 0.5/0.5 = 2:1

Step 4: Convert your answer to the empirical formula, by substituting the numbers for the atomic symbols and adding the required number, representing the number of atoms, after the symbol – 2:1 = H_2O

1) 4g of Titanium reacting with 1g of Carbon (Ti = 48, C = 12)

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empirical formula = ……….……….……

2) 1.12g of Iron reacting with 0.48g of Oxygen (Fe = 56, O = 16)

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empirical formula = ……….……….……

3) 0.31g of Phosphorous reacting with 1.07g of Chlorine (P = 31, Cl = 35.5)

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empirical formula = ……….……….……

4) 6g of Magnesium reacting with 4g of Oxygen (Mg = 24, O = 16)

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empirical formula = ……….……….……

3. Conservation of mass	
Conservation of mass	The total mass of products must equal the total mass of reactants.
Precipitation reaction	A reaction that produces a solid precipitate by mixing two solutions.
Closed system	A system in which no chemicals can enter or leave, such as a sealed test tube.
Open system	A system in which chemicals can enter or leave – such as an open test tube.
Conservation of mass in a closed system	No atoms are able to enter or leave, so the total mass stays the same – for example a precipitation reaction in a closed flask.
Conservation of mass in an open system	For example, a carbonate reacting with acid producing CO_2 bubbles: the mass appears to decrease because you can't weigh the gas that goes into the air, however it is still there.

3. Concentration	
Concentration	The amount of a solute dissolved in a certain volume of solvent.
Calculating concentration (g/dm^{-3})	$\frac{\text{Mass of solute (g)}}{\text{Volume of solution (dm}^3\text{)}}$
Decimetre (dm^3)	A unit of volume equivalent to $1000cm^3$, to covert from cm^3 to dm^3 divide the volume by 1000.

To work out concentrations you need to know the following formula:

$$Concentration\ (mol\ dm^{-3}) = \frac{mass\ (g)}{volume\ of\ solution\ (dm^{-3})}$$

You also need to know that 1 dm^3 = $1000cm^3$, which means that to convert to cm^3 to dm^3 you should divide by 1000.

Work out the concentrations (in g/dm^{-3}) of the following solutions:

1. 20g of NH_3 in $500cm^3$
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2. 10g of Br_2 in $2000cm^3$
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3. 18g of NaOH on $300cm^3$
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Work out the mass of the solute in the following solutions

4. $250cm^3$ of a $200g/dm^{-3}$ solution of $Ca(OH)_2$
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5. $50cm^3$ of a 0.5 M solution of HNO_3
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B The total mass of the reactants always equals the total mass of products.

Use the balanced equations to answer the following questions.

1) 12.4g of copper carbonate was heated and formed 8.0g of copper oxide. Calculate the mass of carbon dioxide produced.

$2CuCO_{3\ (s)} \rightarrow 2CuO_{(s)} + CO_{2\ (g)}$

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2) 1.27g of copper was heated in air and formed 1.59g of copper oxide. Calculate the mass of oxygen that reacted with the copper.

$2Cu_{(s)} + O_{2\ (g)} \rightarrow 2CuO_{(s)}$

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4. Calculating reacting masses	
Excess reactant	Any reactant which is not used up completely in a reaction because there is more of it than needed.
Limiting reactant	Any reactant of which is completely used up in a reaction. The limiting reactant determines how much product is made.
Calculating reacting masses	- Write out the balanced equation - Write the mass of the chemical you are given, and 'm' for the mass you are finding under their symbols - Draw a line underneath the masses to make it a division - Calculate the M_r of each, multiply by the big numbers and write under the line. - Put an equals sign between the two to form an equation. - Solve for 'm'

Calculating the mass of a product

Sometimes, we need to be able to work out how much of a substance is produced in a chemical reaction.

Example: What mass of hydrogen is produced by the electrolysis of 4g of water?

Step 1: Write down the balanced equation, and underline the substances mentioned in the question:

$$\underline{2H_2O} \rightarrow \underline{2H_2} + O_2$$

Step 2: Work out the relative formula mass (M_r) of each substance:

$$2 \times ((2 \times 1) + 16) \rightarrow 2 \times (2 \times 1) + (2 \times 16)$$

Step 3: Since the question only mentions water and hydrogen, you can ignore the oxygen. You just need the ratio of mass of H_2O to mass of H_2:

So, 36g of water produces 4g of hydrogen.

1g of water produces (4 ÷ 36) g of hydrogen (Divide both sides by 36)

4g of water produces (4 ÷ 36) x 4 g of hydrogen

= 0.44g of hydrogen

Worked example

Calculate the mass of chlorine needed to make 53.4 g of aluminium chloride.

Write the balanced equation	$2Al + 3Cl_2 \rightarrow 2AlCl_3$
Calculate relative formula masses of the substances needed	M_r Cl_2 = 2 × 35.5 = 71 M_r $AlCl_3$ = 27 + (3 × 35.5) = 133.5
Calculate ratio of masses (multiply M_r values by the balancing numbers shown in the equation).	$3Cl_2$ makes $2AlCl_3$ so 3 × 71 = 213 g Cl_2 makes 2 × 133.5 = 267 g $AlCl_3$
Work out the mass for 1 g of reactant or product. (Here we want 1 g of the product because that's the mass we know already.)	÷ 267: $\frac{213}{267}$ g Cl_2 makes $\frac{267}{267}$ g $AlCl_3$ ÷ 267 0.798 g Cl_2 makes 1 g $AlCl_3$
Scale up or down (from 1 g to the mass you are given)	× 53.4: 42.6 g Cl_2 makes 53.4 g $AlCl_3$ × 53.4

Q1) What mass of aluminium is produced from 100 tonnes of aluminium oxide? (Al = 27, O = 16)

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

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Q2) What mass of ethanol (C_2H_5OH) is produced from the reaction of 14 tonnes of ethane (C_2H_4)? (C = 12, H = 1, O = 16)

$$C_2H_{4\,(g)} + H_2O_{(g)} \rightarrow C_2H_5OH_{(l)}$$

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Triple Science - Chemistry

SC9 Knowledge organiser

5. Moles (HT)	
Moles	The unit of measurement of chemicals – one mole of any chemical is the same amount.
One mole	An amount of a chemical such that one mole has a mass in grams that is the same as its relative formula mass.
Avogadro's constant	6.02×10^{23}: the number of atoms/molecules present in one mole of a substance.
Calculating moles from mass	Quantity in moles = mass / relative formula mass
Calculating moles from a number of particles	Quantity in moles = number of particles / 6.02×10^{23}
Calculating the number of particles from a mass of substance	Number of particles = (mass / relative formula mass) x 6.02×10^{23}

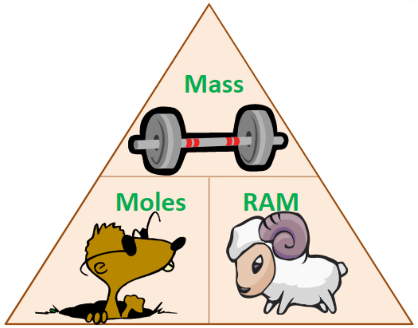
Worked example W2

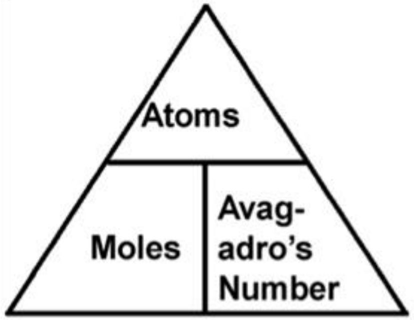
10.8 g of aluminium reacted with 42.6 g of chlorine, Cl_2, to produce aluminium chloride, $AlCl_3$. Deduce the balanced equation for the reaction.

	Al	Cl_2
Calculate the number of moles (= mass/A_r or M_r)	$\frac{10.8}{27} = 0.4$	$\frac{42.6}{2 \times 35.5} = 0.6$
Divide by the smaller	$\frac{0.4}{0.4} = 1$	$\frac{0.6}{0.4} = 1.5$
Simplest whole number ratio	$1 \times 2 = 2$	$1.5 \times 2 = 3$

So 2 mol of Al react with 3 mol of Cl_2. The equation is completed by adding the formula of the product and balancing in the normal way.

$$2Al + 3Cl_2 \rightarrow 2AlCl_3$$





1 Calculate the number of **moles** of water molecules, H_2O, in 9 g of water

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Calculate the number of **moles** of ethanol molecules, C_2H_5OH, in 9.2 g of ethanol.

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2 Calculate the mass of 2.5 mol of potassium iodide, KI

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Calculate the mass of 0.125 mol of calcium sulfate, $CaSO_4$.

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Avagadro's constant = 6.02×10^{23}

1 Calculate the number of molecules in 0.5 mol of carbon dioxide, CO_2

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Calculate the number of molecules in 2 mol of oxygen, O_2.

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2 Calculate the number of moles in 1.505×10^{23} atoms of sodium

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Calculate the number of molecules in 1.806×10^{24} atoms of copper.

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3 Calculate the number of molecules in 9 g of hydrogen, H_2

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Calculate the number of molecules in 48 g of oxygen, O_2.

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6. Stoichiometry (HT)	
Stoichiometry	The ratio of the number of moles of each substance involved in a reaction.
Stoichiometric coefficient	The 'big' numbers written in a balanced equation.
Deducing stoichiometry	- Calculate the number of moles present of each of the reactants (or products) - Find the simplest whole-number ratio - Balance in the normal way to find the numbers of products (or reactants)

Worked example W1

1.50 g of ammonium chloride and 4.00 g of calcium hydroxide are heated together to form ammonia.

$2NH_4Cl + Ca(OH)_2 \rightarrow 2NH_3 + CaCl_2 + 2H_2O$

a Which is the limiting reactant?

b Calculate the mass of ammonia formed.

a The equation shows that 2 mol of NH_4Cl reacts with 1 mol of $Ca(OH)_2$

number of moles of $Ca(OH)_2$ = 4.00 g/(40 + 2(16 + 1)) = 0.0541 mol

We need: 2 × 0.0541 = 0.108 mol NH_4Cl to react with 0.0541 mol of $Ca(OH)_2$.

We have: 1.50 g/(14 + (4 × 1) + 35.5) = 0.0280 mol

We have less than the 0.0541 mol of NH_4Cl needed; NH_4Cl = limiting reactant.

b The equation shows that the number of moles of NH_3 made equals the number of moles of NH_4Cl used.

So, 0.0280 mol of NH_4Cl forms 0.0280 mol of NH_3

mass of NH_3 formed = mol × M_r = 0.0280 × (14 + (3 × 1)) = 0.476 g

$$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$$

1. Identify the stoichiometric ratio between magnesium and sulphuric acid in the reaction above.
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2. If 5 moles of magnesium is reacted with 7 moles of sulphuric acid, which reagent is in excess? How many moles of each product is formed?
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$$Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$$

3. Identify the stoichiometric ratio between iron oxide and carbon monoxide in the reaction above.
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4. If 480 tonnes of iron oxide is reacted with 308 tonnes of carbon monoxide, which reagent is in excess? How many moles of each product is formed?
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1. 2.76 g sodium reacts with 5.70 g titanium chloride, $TiCl_4$ to form titanium and sodium chloride, NaCl. Use this data to deduce the balanced equation for the reaction.

   Relative atomic masses:
   Na = 23   Cl = 35.5   Ti = 48

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2. $TiCl_4 + 2Mg \rightarrow Ti + 2MgCl_2$

   What mass of Titanium (Ti) can be made when 1.9g of Titanium chloride ($TiCl_4$) reacts with 6g of magnesium (Mg?

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3. $WO_{3\,(s)} + 3H_{2\,(g)} \rightarrow W_{(s)} + 3H_2O_{(l)}$

   What mass of Tungsten (W) can be made when 23.2g of Tungsten oxide (WO_3) reacts with 20g of hydrogen (H_2)?

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