CC9: Quantitative chemistry

1. Relative Formula masses

Molecular formula	Gives the number of atoms of each element present in a molecule.
Empirical formula	The **simplest ratio** of the atoms of each element present in a compound.
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_4H_8 ← write the formula 4 : 8 ← write as a ratio $\frac{4}{4}:\frac{8}{4}$ ←divide by small number 1 : 2 ← simplest ratio CH_2 ← write as formula
Relative atomic mass, Ar	The mass of an atom relative to 1/12th the mass of carbon-12. No units.
Relative formula mass, Mr	The mass of one unit of a formula, found by adding the relative atomic masses of all of the atoms in it.

2. Calculating empirical formulae

Steps to calculate empirical formulae from experimental data	1) Write each element's symbol with a ratio (:) symbol between 2) Write out the amount of each element from the questions 3) Divide each amount by the $\mathbf{A_r}$ of the element 4) Divide each answer by the **smallest number** to get a ratio 5) Write the empirical formula
To find a molecular formula from an empirical formula	1) Calculate M_r for the empirical formula 2) Divide the M_r of the molecular formula by this number 3) Multiply the empirical formula by your answer

Empirical formula example

A compound contains 14.3% hydrogen by mass and 85.7% carbon. Determine its empirical formula.

Symbols:	C	:	H
Amounts:	85.7%		14.3%
by A_r:	85.7 ÷ 12 = 7.14		14.3 ÷ 1 = 14.3
÷ by smallest:	7.14 ÷ 7.14 = 1		14.3 ÷ 7.14 = 2
Write formula:	CH_2		

The relative formula mass of the compound is 28, determine its molecular formula.

M_r of empirical: $M_r(CH_2)$ = 12 x 1 + 1 x 2 = 14

÷ molecular M_r by empirical M_r: 28 ÷ 14 = 2

Multiply empirical formula: CH_2 x 2 = C_2H_4

3. Magnesium Oxide Experiment

Equipment	Crucible (small pot capable of withstanding high heat) Clay triangle (to put the crucible on because a gauze would melt)
Method	1) Weigh small amount of magnesium ribbon 2) Heat in a **crucible** to react with air 3) Reweigh once cool to find new mass.
Results	It gets heavier because the oxygen has been added to the solid
Analysis	Find the mass of oxygen added by doing **new mass – old mass.** **Then do the empirical formula calculation**
Magnesium Oxide	Is MgO

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 An insoluble 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 - total mass stays the same. Example: precipitation in a closed flask.
Conservation of mass in an open system	Atoms can leave – total mass appears to change. 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.

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 which is completely used up in a reaction. The limiting reactant determines how much product is made because it will run out of this then stop.
Stoichiometry	Means the balancing of an equation. Use the **limiting reactant** to work out how much is made from balancing.
Calculating reacting masses	1) Write out the balanced equation 2) Calculate the RFMs 3) Write the RFMs as a **ratio** 4) Divide both sides of the ratio by the RFM of the chemical you know the mass of 5) Scale up or down

Calculate concentration	$\text{Concentration} = \frac{\text{mass in g}}{\text{volume in dm}^3}$
Convert cm^3 to dm^3	$\frac{\text{cm}^3}{1000} = \text{dm}^3$

Reacting masses example

What mass of iron can be produced from 50 g of iron oxide (Fe_2O_3)?

$2Fe_2O_3 + 3C$	→	$4Fe + 3CO_2$
320	:	224
$\frac{320}{320}$	:	$\frac{224}{320}$
1	:	0.7
1 × 50	:	0.7 × 50
50g	:	**<u>35g</u>**

RFM calcs: **2 Fe_2O_3:** 2 x (2 x 56 + 3 x 16) = 320

4 Fe: 4 x 56 = 224

5. Moles (HIGHER ONLY)

Moles	Measures **amount of substance** – one mole of any chemical is the same amount.
One mole is…	The Avogadro number of particles (atoms, ions or molecules)
One mole is also…	The mass in grams of 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	$\text{moles} = \frac{\text{mass}}{\text{relative formula mass}}$
Calculating moles from a number of particles	$\text{Quantity in moles} = \frac{\text{no.particles}}{6.02 \times 10^{23}}$
Calculating the number of particles from moles	No. particles = moles x 6.02×10^{23}

Lesson	Memorised?
1. Relative Formula Masses	
2. Calculating Empirical Formulae	
3. Conservation of mass	
4. Reacting masses	
5. Moles	