C3 Quantitative Chemistry – Knowledge organiser
Higher Separate Chemistry GCSE

4.3.1	**Chemical measurements**
Conservation of mass and **balanced chemical equations**	- **no atoms are lost or made** during a chemical reaction so **the mass of the products equals the mass of the reactants.** - This is why symbol equations must be **balanced**, so that there is the same number of atoms on both sides of the equation. e.g. CH_4 + **2**O_2 → CO_2 + **2**H_2O (now each side has 1 C, 4 H's and 4 O's)
Relative formula mass (M_r)	- This is the **sum of all the atomic masses** (A_r)of the atoms in a molecule/compound. e.g. M_r of H_2O= 2x1 + 1x16 = 18; M_r of $CaCO_3$= 1x40 + 1x12 + 3x16 = 100 - In a balanced symbol equation: **M_r of the reactants = M_r of the products**
Explaining mass changes (when a reactant or product is a gas)	- **State symbols** in an equation show the state of the substance: **(s)** solid; **(l)** liquid; **(g)** gas; **(aq)** aqueous (=dissolved in water) - If there appears to be a change in mass in a reaction this is usually because **one of the reactants or products is a gas** whose mass has not been included. e.g. $2Mg_{(s)} + O_{2\,(g)}$ → $2MgO_{(s)}$ **:** the **mass** will appear to **increase**, because the magnesium is **gaining** the atoms of oxygen **gas**. e.g $CaCO_{3\,(s)}$ → $CO_{2\,(g)} + CaO_{(s)}$: the **mass** appears to **decrease**, because carbon dioxide **gas** is made and **escapes into the air**.
Uncertainty in chemical measurements	- Whenever a measurement is made there is always some **uncertainty** about the result obtained. This can be represented on a graph by plotting the **range** of repeats as well as the **mean**. **The bigger the range around the mean, the more uncertainty there is** about the results.
4.3.2	**Moles, masses and concentrations**
Moles the symbol is 'mol'	- 1 mole of a substance always contains the same number of atoms/molecules of that substance: this is **the Avogadro constant = 6.02×10^{23}** atoms/molecules per mole. - The **mass of 1 mole of a substance in grams** in the same number as its **relative formula mass**. e.g - the mass of 1 mole of carbon = 12g (because the A_r of carbon is 12) - the mass of 1 mole of H_2O = 18g (because the M_r of H_2O is 18) - **number of moles = mass of a substance ÷ relative formula mass**
Using moles to balance equations Remember: moles = mass ÷ M_r	- if you work out the number of moles of each reactant and product in a reaction (using moles = mass ÷ relative formula mass) this gives you the **ratio of reactants and products**, so you can write the **balanced equation**. e.g 48g of Mg reacts with 32g of O_2 to produce 80g of MgO so: 48÷24 = **2mol of Mg**; 32÷(2x16)= **1mol of O_2**; 80÷(24+16)= **2mol of MgO** this is a **ratio** of **2:1:2** (Mg: O_2: MgO)

	so the **balanced equation** is: **2**Mg + **(1)**O_2 → **2**MgO
Calculating masses of reactants and products Remember: moles = mass ÷ M_r	- The masses of reactants and products can be calculated from balanced symbol equations, because these give you the ratio of moles of each one. e.g Mg + 2HCl → $MgCl_2$ + H_2 this mean you have: **1mol of Mg** that reacts with **2mol of HCl** to produce **1mol of $MgCl_2$** and **1mol of H_2**: the ratio is **1:2:1:1** so if you have **48g of Mg** (48g÷24 = 2mol) this will react with: 4mol×36.5 = **146g of HCl** to produce 2mol×95 = **190g of $MgCl_2$** and 2mol×2 = **4g of H_2**
Limiting reactants	- In a reaction with with 2 reactants, it is common to use an **excess** of one reactant to make sure that **all** of the other reactant is used up. This reactant that is completely used up is called the **limiting reactant**, as it **limits the amount of the products** that can be made. - if you know that number of moles or the mass in grams of this limiting product, you can **calculate the moles or mass of the products formed**.
Concentrations of solutions	- The concentration of a solution (aq) can be measured in **g/dm^3** (mass/volume) **Concentration = mass ÷ volume** - The concentration of the **solution** depends on the mass of the **solute** and the volume of the **solvent**. Increasing mass increases concentration, increasing volume decreases concentration.
4.3.4	**Using mol/dm^3 for concentration**
Concentration in mol/dm^3 Remember: moles = mass ÷ M_r	- You can use the concentration in mol/dm^3 to calculate the number of moles or the mass in grams of a solute in a given volume. **Concentration = number of moles ÷ volume** - You can use this equation and a balanced symbol equation to work out the **concentration of a reactant** if you have its volume and the volume and concentration of the other reactant.
4.3.5	**Volume of gases**
Volumes of gases Remember: moles = mass ÷ M_r	- 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^{o}C and 1 atmosphere pressure) is **24dm^3**. - You can calculate the **volume of a gas at room temperature and pressure** from its mass and relative formula mass (because mass ÷ M_r = number of moles, and volume of 1 mole is 24dm^3 at room temperature and pressure) - You can calculate **volumes of gaseous reactants and products** from a **balanced equation** and a given volume of a gaseous reactant or product (again by using moles)
4.3.3	**Yield and atom economy**
Percentage Yield	**% yield = actual mass of product made ÷ expected mass of product (×100)** - It is not always possible to get all of the expected amount of a product because:

	• the reaction may not finish completely because it is a **reversible reaction** • **some of the product may be lost** when it is separated from the reaction mixture • some of the **reactants may react differently** to the expected reaction
Theoretical mass	- The theoretical mass of a product is how much of it you **expect** to get in a reaction. This can be calculated if you know the mass of one of the reactants and the balanced symbol equation (see the '**Calculating masses of reactants and products**' section above)
Atom economy	**% atom economy = M_r of useful product ÷ total M_r of all reactants (x100)** - This is a measure of the percentage of starting reactants that end up as useful products. - It is important for **economic reasons** and for **sustainable development** to use reactions with a **high atom economy**. - Decisions about which reaction pathway to choose will depend on the atom economy, yield, rate of reaction, equilibrium position (in reversible reactions) and usefulness of by-products,