

C9: 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	$C_2H_4 \rightarrow CH_2$ (divided by 2) $C_6H_{12}O_6 \rightarrow CH_2O$ (divided by 6) $H_2O \rightarrow H_2O$ (divided by 1)
*Relative atomic mass, A_r	The mass of an atom relative to $1/12^{th}$ 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.

2. Calculating empirical formulae

*To calculate empirical formulae from experimental data	<ul style="list-style-type: none"> - 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
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**To find a molecular formula from an empirical formula	<ul style="list-style-type: none"> - 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
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*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:	$85.7 \div 12 = 7.14$		$14.3 \div 1 = 14.3$
÷ by smallest:	$7.14 \div 7.14 = 1$		$14.3 \div 7.14 = 2$
Write formula:	CH ₂		

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

M_r of empirical: $M_r(CH_2) = 12 \times 1 + 1 \times 2 = 14$
÷ molecular M_r by empirical M_r : $28 \div 14 = 2$
Multiply empirical formula: $CH_2 \times 2 = C_2H_4$

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.

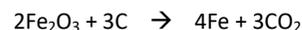
**Conservation of mass in an open system	For example, a carbonate reacting with acid producing CO ₂ bubbles: the mass appears to decrease because you can't weigh the gas that goes into the air, however it is still there.
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4. Calculating reacting masses

***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	<ul style="list-style-type: none"> - 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'

**Reacting masses example

What mass of iron can be produced from 50 g of iron oxide (Fe₂O₃)?



$$\frac{50}{320^*} = \frac{m}{224^*}$$

$$\frac{50}{320} \times 224 = m$$

$$35 \text{ g} = m$$

***2 Fe₂O₃:** $2 \times (2 \times 56 + 3 \times 16) = 320$

***4 Fe:** $4 \times 56 = 224$

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) $\times 6.02 \times 10^{23}$

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	<ul style="list-style-type: none"> - 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)