# 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)<br>$C_6H_{12}O_6 \rightarrow CH_2O$ (divided by 6)<br>$H_2O \rightarrow H_2O$ (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. |

| 2. Calculating empirical formulae | |
|---|---|
| ***To calculate empirical formulae from experimental data** | - Write each element's symbol with a ratio (:) symbol between<br>- Write out the amount of each element from the questions<br>- Divide each amount by the $A_r$ of the element<br>- Divide each answer by the smallest answer to get a ratio<br>- Write the empirical formula |
| ****To find a molecular formula from an empirical formula** | - Calculate $M_r$ for the empirical formula<br>- Divide the $M_r$ of the molecular formula by this number<br>- 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:** | CH2 | | |

**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 ÷ 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. |

| 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<br>- Write the mass of the chemical you are given, and 'm' for the mass you are finding under their symbols<br>- Draw a line underneath the masses to make it a division<br>- Calculate the $M_r$ of each, multiply by the big numbers and write under the line.<br>- Put an equals sign between the two to form an equation.<br>- Solve for 'm' |

****Reacting masses example**

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

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

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

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

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

***2 $Fe_2O_3$:** 2 x (2 x 56 + 3 x 16) = 320
***4 Fe:** 4 x 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** | $6.02 \times 10^{23}$: 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 / $6.02 \times 10^{23}$ |
| *****Calculating the number of particles from a mass of substance** | Number of particles = (mass / relative formula mass) x $6.02 \times 10^{23}$ |

| 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** | - Calculate the number of moles present in each of the reactants (or products)<br>- Find the simplest whole-number ratio<br>- Balance in the normal way to find the numbers of products (or reactants) |