C9: Quantitative chemistry Knowledge Organiser

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 1/12th 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 - 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

***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 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. 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 - 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 (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**	6.02×10^{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}

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) - Find the simplest whole-number ratio - Balance in the normal way to find the numbers of products (or reactants)