The content on this sheet is common to Paper 1 & Paper 2

Mass Calculations and Quantitative Analysis

* Indicates that these are some examples only: you could be asked about any substance / reaction.

1.43—Calculate relative atomic masses

Use the relative atomic masses given in the question and the formula. Replace each symbol with the relative atomic mass, and add all of the values together.

Calculate the relative formula mass of:

$CaCO_3$ A_r: Ca = 40, C = 12, O = 16

40 + 12 + 16 + 16 + 16 = 100

Na_2SO_4 A_r: Na = 23, S = 32, O = 16

23 + 23 + 32 + 16 + 16 + 16 + 16 = 142

1.44—Calculate empirical formulae from masses or percentages

The empirical formula is the simplest whole number *ratio* of atoms in a compound. It can be calculated using the following method (you **must** show working in these):

Step 1: $\frac{mass\ or\ percentage}{relative\ atomic\ mass}$ for each element in the question.

Step 2: Divide each value from step 1 by the smallest value from step 1. You'll probably have whole numbers at this point. If you don't, either a) round them if they are very close to a whole number or b) multiple all the values by a bigger number (probably 2), so all the values are whole numbers.

Step 3: Write the formula.

Calculate the empirical formula of an iron oxide containing 1.12 g of iron and 0.48 g of oxygen. A_r: Fe = 56, O = 16.

Fe	O	
$\frac{1.12}{56}$	$\frac{0.48}{16}$	
$= \frac{0.02}{0.02}$	$= \frac{0.03}{0.02}$	
= 1	= 1.5	
= 2	= 3	Fe_2O_3

1.45a—Deduce empirical formulae from molecular formulae

Here, all you need to do is simplify the *ratio* of atoms. Divide all numbers by the highest common factor to do this.

Deduce the empirical formula of P_4O_{10}.

P_2O_5

1.45b—Deduce the molecular formula from empirical formula & formula mass

Step 1: work out the formula mass of the empirical formula.

Step 2: work out the *ratio* between the empirical and molecular formula masses.

Step 3: multiply the empirical formula by this value.

Calculate the molecular formula of a compound with a relative formula mass of 180 and an empirical formula of CH_2O. A_r: C = 12, H = 1, O = 16.

12 + 1 + 1 + 16 = 30

30 : 180 = 1 : 6

CH_2O x 6 = $C_6H_{12}O_6$

1.46—Describe an experiment to determine empirical formula

This is a classic practical involving **heating** magnesium in a **crucible** to a constant **mass**. Outline the method below. Think about the equipment you will need.

- Record the mass of a piece of metal.
- Place it in a crucible. Record the total mass.
- Heat the metal strongly. Lift the lid to allow oxygen in.
- Continue heating until the mass remains constant.
- Calculate the mass of oxygen that reacted.
- Use the mass of oxygen & magnesium to calculate the empirical formula.

What are the safety concerns, and how would you manage them?

Hot equipment—allow to cool before handling / use tongs to move the hot crucible.

Hot products leaving crucible—wear goggles to protect eyes.

1.47a—Conservation of mass: closed system

In a closed system, the total mass of the reactants is equal to the total mass of the products, as atoms cannot be created or destroyed in a chemical reaction.

1.47b—Conservation of mass: open flask

If a gas is being made in a reaction, it may appear as though mass is being lost. However, this mass is in the gas that has been released by the reaction, so mass is conserved.

1.48—Calculating reacting masses

Step 1: Do the calculation (right) for the substance you have this data for. $\frac{mass}{relative\ atomic\ or\ formula\ mass}$

Step 2: Use the equation to work out the *ratio* of the substance you need to find the mass for to the substance you worked out in step 1.

Step 3: *Ratio x relative atomic or formula mass x answer from step 1.*

Calculate the maximum mass of iron that can be extracted from 320 g of iron oxide. A_r: Fe = 56, O = 16.

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

$\frac{320}{56 + 56 + 16 + 16 + 16} = 2$

Fe : Fe_2O_3 = 4 : 2

4/2 x 56 x 2 = 224 g

1.49—Calculating the concentrations of solutions in g dm^{-3}

First thing to remember: you'll probably need to convert from cm^3 to dm^3. To do this, divide the value in cm^3 by 1000.

The formula is: $\frac{mass\ (in\ grams)}{volume\ (in\ dm^3)}$

Calculate the concentration, in g dm^{-3} of a solution of 12 g of solute in 250 cm^3 of water.

$\frac{250}{1000} = 0.25$ $\frac{12}{0.25} = 48$ g dm^{-3}

1.50a—Moles and Avogadro's constant (HT only)

1 mole = 6.02 x 10^{23} particles (atoms, ions, molecules or formulae).

1.50b—Moles and mass

1 mole = the relative mass of a substance in grams.

1.51a—Moles of particles from mass & vice versa (HT only)

$Number\ of\ moles = \frac{mass}{relative\ mass}$

Calculate the number of moles in 220 g of carbon dioxide. A_r: C = 12, O = 16.

$\frac{220}{12 + 16 + 16} = 5$ moles

1.51b—Number of particles from moles & vice versa (HT only)

$Number\ of\ particles = moles\ x\ Avogadro's\ number$

Calculate the number of molecules in 2 moles of carbon dioxide.

2 x 6.02 x 10^{23} = 1.204 x 10^{24} molecules

1.51c—Particles of a substance in a given mass (HT only)

Combine 1.51a and 1.51b to calculate these.

Calculate the number of molecules in 8.8 g of carbon dioxide.

$\frac{8.8}{12 + 16 + 16} = 0.2$ 0.2 x 6.02 x 10^{23} = 1.204 x 10^{22} molecules

1.52—Limiting reactants (HT only)

In a reaction, the maximum mass of a product is limited by the reactant which is not in excess. This is because once all of the reactant has reacted, no additional product can be formed.

1.53—Calculating a balanced symbol equation [stoichiometry] (HT only)

Step 1: Calculate the number of moles of the reactants from the mass and relative masses in the question. $Number\ of\ moles = \frac{mass}{relative\ mass}$

Step 2: Convert these values to whole numbers—these are the values for the reactants.

Step 3: The product will (most likely) be all of the reactants forming one product—simply write one formula with the total number of each atom.

In an experiment, 2.4 g of carbon reacted with 3.2 g of oxygen, O_2 to form an oxide of carbon. Determine the balanced symbol equation. A_r: C = 12, O = 16.

C: $\frac{2.4}{12} = 0.2$ O_2: $\frac{3.2}{32} = 0.1$

2   1

$2 C + O_2 \longrightarrow C_2O_2$ (although this is, in reality, 2 CO)