Topic 1: Stoichiometric relationships

1.1 Introduction to the particulate nature of matter and chemical change

U1	Atoms of different elements combine in fixed ratios to form compounds, which have different properties from their components elements
U2	Mixtures contain more than one element and/or compound that are not chemically bonded together and so retain their individual properties
U3	Mixtures are either homogeneous or heterogeneous.
A1	Deduction of chemical equations when reactants and products are specified.
A2	Application of the state symbols (s), (l), (g) and (aq) in equations.
A3	Explanation of observable changes in physical properties and temperature during changes of state.

State of matter

Solid	Liquid	Gas
Fixed volume, fixed shape	Fixed volume, no fixed shape	No fixed shape, no fixed volume
Cannot be compressed	Cannot be compressed	Can be compressed
Attractive forces between particles hold the particles in a close-packed arrangement	Forces between particles are weaker than in solids	Forces between particles are taken as zero
Particles vibrate in a fixed position	Particles move around(vibrate, rotate, translate)	Particles move around freely, much faster than a liquid

The way the particles move depends on the **temperature**.

Temperature (K) = temperature (C) + 273.15

Elements and compounds

- Elements: contains atoms of only one type
- Compounds: atoms of elements combines in a fixed ratio
- **Mixtures** contain more than one element and/or compound that are not chemically bonded together and so retain their individual properties
- **Homogeneous mixture**: has uniform composition and uniform properties (metal alloys)
- **Heterogeneous mixture**: has non-uniform composition and non-uniform properties (paints)

1.2 The mole concept

U1	The mole is a fixed number of particles and refers to the amount, n, of substance
U2	Masses of atoms are compared on a scale relative to ^{12}C and are expressed as relative atomic mass (Ar) and relative formula/molecular mass (Mr).
U3	Molar mass (M) has the units g mol^{-1}.
U4	The empirical formula and molecular formula of a compound give the simplest ratio and the actual number of atoms present in a molecule respectively
A1	Calculation of the molar masses of atoms, ions, molecules and formula units.
A2	Solution of problems involving the relationships between the number of particles, the amount of substance in moles and the mass in grams.
A3	Interconversion of the percentage composition by mass and the empirical formula.
A4	Determination of the molecular formula of a compound from its empirical formula and molar mass.
A5	Obtaining and using experimental data for deriving empirical formulas from reactions involving mass changes.

Moles

A **mole** (mol) of a substance (atoms / molecules / ions):

- contains as many elementary entities (particles, atoms, molecules etc.) as there are atoms in 12 grams of pure carbon-12
- is the relative atomic (or molecular or formula) mass expressed in grams
- 6.02×10^{23}elementary entities (**Avogadro's constant**)

Molar mass

The **molar mass** (g mol^{-1}) is the mass in grams of one mole of a substance.

Number of particles = number of moles × Avogadro's constant

Relative abundance

Isotopes are atoms of the same element that have the same number of protons in the nucleus but different numbers of neutrons.

Relative abundance 1 × atomic mass 1 + Relative abundance 2 × atomic mass 2 = Relative atomic mass of this substance

e.g. 75% [^{35}Cl] × 35.0 + 25% [^{37}Cl] × 37.0 = 35.5

Relative mass

Definitions

Relative atomic mass (A_r) – the relative atomic mass is the average mass of an atom, taking into account the relative abundances of all the naturally occurring isotopes of the element, relative to one atom of C-12. It is a relative term so it has no units.

Relative molecular mass (M_r)–the relative molecular mass is the average mass of a molecule, calculated by adding the relative atomic masses of its constituent atoms. It is a relative term so it has no units.

For formula units (the single units of ionic compounds), the term **relative formula mass** is used. The mass of a mole of a species is the relative mass expressed in grams. The number of moles is given by:

$$\text{Moles (mol)} = \frac{\text{Mass (g)}}{\text{Molar mass (g mol}^{-1}\text{)}}$$

Formulas

The **molecular formula** is the chemical formula of a substance showing the actual number of atoms of each element in a molecule. The **empirical formula** is the formula in which the ratio is simplified into the smallest integers.

Problem solving

Finding the empirical formula from percent composition:

Just simplify to the smallest integer ratio (round with 0.1 difference $O_{1.9}$– O_2)

Finding the empirical formula from percent mass:

Divide the percentages by the relative atomic masses and then simplify to the smallest integer ratio

Finding the molecular formula given the empirical formula and relative molecular mass:

1. Use the empirical formula to find the empirical mass
2. Divide the molecular mass by the empirical mass, and round to an integer, *n*
3. Multiply the empirical formula by *n*

Experimental Analysis

- Qualitative analysis: determine which elements are in the compound, verify the purity of the compound
- Quantitative analysis: relative mass of elements, work out exact composition

1.3 Reacting masses and volumes

U1	Calculate theoretical yields from chemical equations.
U2	The experimental yield can be different from the theoretical yield.
U3	Avogadro's law enables the mole ratio of reacting gases to be determined from volumes of the gases.
U4	The molar volume of an ideal gas is a constant at specified temperature and pressure.

U5	The molar concentration of a solution is determined by the amount of solute and the volume of solution.
U6	A standard solution is one of known concentration.
A1	Solution of problems relating to reacting quantities, limiting and excess reactants, theoretical, experimental and percentage yields.
A2	Calculation of reacting volumes of gases using Avogadro's law.
A3	Solution of problems and analysis of graphs involving the relationship between temperature, pressure and volume for a fixed mass of an ideal gas.
A4	Solution of problems relating to the ideal gas equation.
A5	Explanation of the deviation of real gases from ideal behaviour at low temperature and high pressure
A6	Obtaining and using experimental values to calculate the molar mass of a gas from the ideal gas equation.
A7	Solution of problems involving molar concentration, amount of solute and volume of solution.
A8	Use of the experimental method of titration to calculate the concentration of a solution by reference to a standard solution.

Theoretical yield

Theoretical yields assume that the **limiting reagent** is 100% used up. The actual yield is called the **experimental yield**.

$$\text{Percentage yield }(\%) = \frac{\text{experimental yield}}{\text{theoretical yield}} \times 100\%$$

Percentage Purity

$$\text{Percentage purity }(\%) = \frac{\text{mass of the pure compound}}{\text{total mass of the mixture}} \times 100\%$$

Limiting and excess reactants

Limiting reagent as a means of controlling the amount of product obtained, will be completely consumed during the experiment. Limiting reagents determines amount of products produced

Problem solving

If A + B $\rightarrow$ C

1. Calculate how much of reactant B (in moles) is needed to react with the amount of reactant A provided.
2. If the amount needed is less than that provided, B is in **excess** and vice versa. A is therefore the **limiting reagent**.
3. The difference between the amount of B needed and the amount of B provided is equivalent to the moles of excess B.
4. The amount of product formed if A reacts completely is the theoretical yield.

Avogadro's law

Avogadro's law states that equal volumes of all gases under identical conditions of pressure and temperature contain the same number of molecules i.e. the molar volume of gas A is the same as that of gas B if the conditions (pressure and temperature) are the same. Under STP conditions **(273K and 1.013 × 10^5 Pa or 1 atm) the volume per mole is 22.7 dm^3. (STP)**

Ideal gases

The ideal gas equation is:

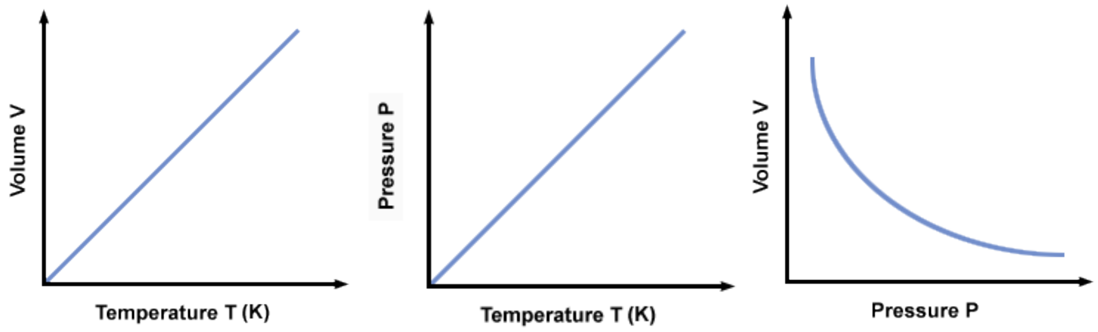
$$PV = nRT$$

Where R is the gas constant which has a value of 8.31 J K^{-1} mol^{-1}

The ideal gas equation implies **Avogadro's law**. We can also use it to derive laws for specific cases by moving the variables to the left hand side and the constants to the right hand side.

Law	Result	Formula
Combined gas law	$\frac{PV}{T} = k$	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Gay-Lussac's law	$\frac{P}{T} = k$	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
Boyle's law	$PV = k$	$P_1V_1 = P_2V_2$
Charles's law	$\frac{V}{T} = k$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$



Definitions

Solute – the solute is the substance dissolved in a solvent in forming a solution.

Solvent – the solvent is the liquid that dissolves another substance or substances to form a solution.

Solution – a solution is a homogeneous mixture of a liquid (the solvent) with another substance (the solute). There is usually some interaction between the solvent and solute molecules.

Concentration – concentration is the amount of solute in a known volume of solution. It can be expressed either as g dm^{-3} or as mol dm^{-3}. Concentration in mol dm^{-3} is often represented by square brackets around the substance.

$$c = \frac{n}{V}$$

Molarity – is the concentration of the solution in terms of mol/L Unit: mol dm^{-3}

Standard Conditions for temperature and pressure (RTP)(STAP)

298K, 1.01× 10^5Pa ---- 24.0 dm^3/mol

STP Condition

273K, 1.01× 10^5Pa ---- 22.7 dm^3/mol