Edexcel IAL Chemistry A-Level

Topic 1: Formulae, Equations and Amounts of Substance

Detailed Notes

Moles and the Avogadro Constant

The mole is a **unit of measurement** for substances. It always contains the **same number of particles**.

$$L = 6.022 \times 10^{23} \text{ particles}$$

This number is the **Avogadro Constant** (L) and allows the number of particles present in a sample of a substance with known mass to be found:

$$\text{Number of particles} = n \times L$$

(n = moles) (L = Avogadro constant)

The mole is a **very important unit of measurement** in many calculations:

$$\text{Moles} = \frac{\text{mass}}{\text{Mr}} = \frac{\text{concentration} \times \text{volume}}{1000}$$

(where concentration is in mol dm^{-3} and volume is in cm^3)

Key terms

Types of species

An **atom** is the **smallest part of an element** that can exist. All substances are made up of atoms.

An **ion** is what's formed when an atom **loses or gains electrons**. This gives it an overall **charge** - a positive charge if it has lost at least one electron and a negative charge if it has gained at least one electron.

A **molecule** consists of two or more atoms that have been **bonded** together chemically.

A **compound** is a substance that combines **two or more different elements** through the formation of **chemical bonds**.

Ar and Mr

Relative atomic mass (**Ar**) is defined as:

The mean mass of an atom of an element, divided by one twelfth of the mass of an atom of the carbon-12 isotope.

Relative molecular mass (**Mr**) is defined as:

The mean mass of a molecule of a compound, divided by one twelfth of the mass of an atom of the carbon-12 isotope.

The molecular mass (**Mr**) of a compound or molecule can be calculated by adding together the atomic masses (**Ar**) of all the atoms in that compound.

Example:

To calculate the Mr of the compound C_2H_5OH the Ar's must be used:
C = 12
O = 16
H = 1

$C_2H_5OH = (2 \times 12) + (6 \times 1) + (16 \times 1)$
Mr = 46

Relative formula mass refers to compounds that have a giant structure.

Empirical and Molecular Formula

An empirical formula is the **simplest whole number ratio** of atoms of each element in a compound. It is found using **molar ratios** of each element.

Molecular formula is the **actual number of each atom in the molecule**. It can be determined using the **Mr of the empirical formula** and the **true Mr** of the molecule. This gives a **multiplier** value which can be used to scale up the empirical formula.

True Mr = Mr of empirical formula x multiplier

Example

The empirical formula of a molecule containing 5 atoms of oxygen for every 2 atoms of phosphorus has an Mr of 284. What is its molecular formula?   Empirical formula = P_2O_5 Mr of empirical formula = (31x2) + (16x5) = 142   Multiplier = 284 ÷ 142 = 2   Molecular formula = $2(P_2O_5) = P_4O_{10}$

Molar mass

The **molar mass** of a substance is its mass in grams per mole and has the units **g mol^{-1}**. It can be calculated using the following equation:

$$\text{Molar mass} = \frac{\text{Mass}}{\text{Number of moles}}$$

Parts per million (ppm)

Concentration can be given in **parts per million (ppm)**. This gives the units of mass of that particular species within 1,000,000 total units of mass. It is most commonly used to represent the **concentrations of gases.**

Volumes of gases

Molar volume of gases

One mole of any gas at **room temperature and pressure** will take up the **same volume**, regardless of its composition. This volume is 24,000 cm^3, or 24 dm^3, and is known as the **molar volume of gases**.

This relationship gives the following equation that can be used to work out the volume of a gas if its amount (number of moles) is known and vice versa.

$$\text{Volume of gas (dm}^3\text{)} = 24 \times \text{Number of moles}$$

(At room temperature & pressure)

The ideal gas law

When under **standard conditions**, gases and volatile liquids follow certain trends:

Pressure is proportional to Temperature
Volume is proportional to Temperature
Pressure and Volume are inversely proportional

These relationships can be combined to give the **ideal gas equation**:

$$\boxed{pV = nRT = \frac{mRT}{Mr}}$$

In order to use this equation, the variables must be in the correct **standard units**:

p = pressure in Pascals
V = volume in m^3
T = temperature in Kelvin
n = moles
m = mass in grams

R is the **ideal gas constant**, equal to **8.31 $JK^{-1}mol^{-1}$**.

Equations and Calculations

Full or ionic chemical equations must be **balanced** before they can be used in calculations. This is because the **reacting ratios** must be correct. For a chemical equation to be balanced, it must have the **same number and type** of each atom present on both sides of the equation.

It can be useful to also include **state symbols** so it is clear what might be observed during the reaction, for example, **bubbles** of gas, a **precipitate** forming, or a **colour change** that may infer a **displacement reaction.**
There are four state symbols:

- (s) - solid
- (l) - liquid
- (g) - gas
- (aq) - aqueous (dissolved in water)

These balanced equations can then be used to calculate **reacting masses, percentage yield** and **atom economy**.

Ionic equations

Ionic equations show just the **reacting particles** that undergo a change during the reaction and not the **spectator species**. As with normal chemical equations, it must be balanced. The reacting species are shown as **dissociated ions.**

Percentage Yield

The percentage yield indicates how much of the maximum amount of product you obtained during an **experiment**. A **low** percentage yield could indicate an **incomplete reaction**, or the loss of product during **purification**.

$$\% \text{ yield} = \frac{\text{Experimental mass} \times 100}{\text{Theoretical mass}}$$

Atom Economy

The atom economy is a measure of **efficiency** since it measures the **proportion** of reactant atoms which are converted into the **desired product**.

$$\% \text{ atom economy} = \frac{\text{Mr of desired product} \times 100}{\text{Total Mr of all products}}$$

In industrial chemical processes, it is desirable to have a **high atom economy** for a reaction. This means there is **little or no waste product**, only the desired product. Therefore it means the process is more **economically viable** for industrial-scale manufacture.

Concentration calculations

The concentration of a solution can be measured in **mol dm^{-3}** and **g dm^{-3}** which can be calculated using the following equations:

$$\textbf{Concentration (mol dm}^{-3}\textbf{)} = \frac{\text{Number of moles (mol)}}{\text{Volume (dm}^3\text{)}}$$

$$\textbf{Concentration (g dm}^{-3}\textbf{)} = \frac{\text{Mass (g)}}{\text{Volume (dm}^3\text{)}}$$

Experimental data

Experimental data can be used to work out **empirical and molecular formulas** and **reaction stoichiometries**. These calculations require use of the equations given in this section, along with some others. To summarise, these include:

Mol = volume x concentration

True Mr = Mr of empirical formula x multiplier

Volume of gas (dm^3) = 24 x Number of moles
(At room temperature & pressure)

Number of particles = n x L

Mass = Mr x mol