Edexcel Chemistry A-level

Topic 5: 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:

$$\text{Number of particles} = nL$$

(n = moles) (L = Avogadro constant)

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

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

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

The **molar mass** of a substance is its mass in grams per mole and has the units **g mol^{-1}**.

Mr and Ar

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

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

For **ionic compounds**, relative molecular mass is known as **relative formula mass**.

Empirical and Molecular Formula

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 = $2(P_2O_5) = P_4O_{10}$

The Ideal Gas Equation

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**:

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

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 J $K^{-1}mol^{-1}$**.

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.**

Examples:

$$Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \rightarrow AgCl_{(s)}$$

$$Cl_{2(aq)} + 2Br^{-}_{(aq)} \rightarrow Br_{2(aq)} + 2Cl^{-}_{(aq)}$$

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**.

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

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.

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

Acid-Base Titrations

A titration is a practical method where a **standard solution** of known concentration is reacted with a solution of **unknown concentration** in order to determine its concentration. For this, there is a standard method to make up the standard solution and carry out the titration.

Volumetric Solution - Simple Method

1. Weigh the sample bottle containing the solid on a (2 dp) balance.
2. Transfer solid to beaker and reweigh sample bottle.
3. Record the difference in mass.
4. Add distilled water and stir with a glass rod until all the solid has dissolved.
5. Transfer to a volumetric flask with washings.
6. Make up to the 250cm^3 mark with distilled water.
7. Shake flask.

Common errors in this method include **systematic errors** on the balance, **lost substance** in transfer processes and **overfilling** of the volumetric flask. These can be reduced using **washing** methods and by reading volumes from the **bottom of the meniscus**.

Titration - Simple Method

1. Fill the burette with the standard solution of known concentration, ensuring the jet space in the burette is filled and doesn't contain air bubbles.
2. Using a pipette filler and pipette to transfer 25cm^3 of the solution with unknown concentration into a conical flask.
3. Add two to three drops of indicator.
4. Record the initial burette reading.
5. Titrate the contents of the conical flask by adding solution to it from the burette until the indicator undergoes a definite, permanent colour change.
6. Record the final burette reading and calculate the titre volume.
7. Repeat until at least two concordant results are obtained (within 0.1cm^3 of each other).

Uncertainty

The equipment used in a titration all comes with their own **uncertainty values**. These must be combined to find the overall uncertainty in the final answer.
The percentage uncertainty of a measurement can be calculated if the uncertainty of the instrument is known.

$$\text{Percentage uncertainty} = \frac{\text{instrument uncertainty}}{\text{measurement}} \times 100$$

Example:

Calculate the percentage uncertainty when a 25 cm^3 volume was recorded using a burette which has a 0.5 cm^3 uncertainty.

$$\text{Percentage uncertainty} = \frac{0.5}{25} \times 100$$
$$= 0.02 \times 100$$
$$= 2\%$$

The best way of reducing uncertainties in a titration is to **increase the titre volume needed** for the reaction. This can be done by increasing the volume and concentration of the substance in the conical flask or by decreasing the concentration of the substance in the burette.

It is also important to carry out a **risk assessment** before undertaking any practical work. This should analyse **equipment, the lab environment** and the **chemicals** being used and suggest methods for **reducing the risk** and what should be done if an accident occurs.