FACTFILE:
GCSE DAS CHEMISTRY: UNIT 1.7

Rewarding Learning

Quantitative chemistry

Learning outcomes

Student should be able to:

1.7.1 recall that the relative atomic mass (A_r) of an atom is the mass of the atom compared with that of the carbon-12 isotope, which has a mass of exactly 12, and demonstrate knowledge and understanding that it is a weighted mean of the mass numbers (linked to 1.1.12);

1.7.2 calculate relative formula mass (M_r) (relative molecular mass) of a compound and percentage of an element, by mass, in a compound;

1.7.3 demonstrate knowledge and understanding that chemical amounts are measured in moles and that the mass of one mole of a substance in grams is numerically equal to the relative formula mass;

1.7.4 convert the given mass of a substance to the amount of the substance in moles (and vice versa) by using the relative atomic or formula masses;

1.7.5 **calculate the reacting masses of reactants or products, given a balanced symbol equation and using moles and simple ratio, including examples where there is a limiting reagent;**

1.7.7 **calculate the theoretical yield, and the percentage yield, of a chemical reaction given the actual yield;**

1.7.8 **recognise possible reasons why the percentage yield of a product is less than 100 %, including loss of product in separation from the reaction mixture, as a result of side reactions or because the reaction is reversible and may not go to completion;**

Relative atomic mass (A_r)

The relative atomic mass is a weighted mean mass of the isotopes of an element compared with that of the carbon-12 isotope, which has a mass of exactly 12.

Relative formula mass (M_r)

The **relative formula mass** (M_r) (**relative molecular mass** for covalent substances) is the sum of the relative atomic mass of all the atoms present in the formula of a substance.

Example

What is the relative formula mass of $Ca(NO_3)_2$?

$M_r = 1 \times A_r\,Ca + 2 \times A_r\,N + 6 \times A_r\,O$

$= 40 + 2 \times 14 + 6 \times 16 = 164$

The mole

Chemical amounts are measured in moles. The mass of one mole of a substance in grams is numerically equal to the relative formula mass

$$\text{moles} = \frac{\text{mass (g)}}{\text{relative formula mass } (M_r)}$$

Example
Calculate the number of moles in 100g of $Ca(OH)_2$.

Answer
M_r $Ca(OH)_2$ = 40 + 2 × 16 + 2 × 1 = 74

$$\text{Number of moles} = \text{mass (g)} = \frac{\text{mass (g)}}{M_r} = \frac{100}{74} = 1.35 \text{ mol}$$

Mass in g = moles × M_r

Example
Calculate the mass, in grams of 20 moles of CO_2.

Answer
M_r CO_2 = 12 + 2 × 16 = 44

Mass = moles × M_r = 20 × 44 = 880 g

Mole ratios in equations

The balancing numbers in a chemical equation give the **ratio** of moles which react together.

For example 2 moles of calcium react with one mole of oxygen to produce 2 moles of calcium oxide.

$2Ca + O_2 \rightarrow 2CaO$

The ratio is 2 moles Ca : 1 mole O_2 : 2 moles CaO

The ratio can be used to calculate the number of moles which would react, and be produced in any one equation.

Example
In the reaction

$2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2 + O_2$

How many moles of copper(II) nitrate are needed to produce 0.4 moles of oxygen O_2?

Answer
$Cu(NO_3)_2$: O_2
Ratio 2 : 1
? : 0.4

There are twice as many moles of copper(II) nitrate as oxygen so multiply by 2.

$0.4 \times 2 = 0.8$ moles of $Cu(NO_3)_2$

Calculating reacting masses of reactant and products.

- Write the information (mass) given in the question, underneath the equation.
- Calculate the relative formula mass (M_r) of the substance you have information about and calculate the number of moles. (Remember that the balancing numbers in the equation are **not** part of the formulae).
- Use the ratio from the balanced equation to calculate the number of moles of the substance you need to find the mass of.
- Calculate the mass of the substance. (using mass = moles × M_r)

Example

What mass of magnesium oxide reacts with 3.65 g of HCl?

$\underline{MgO} + 2HCl \rightarrow MgCl_2 + H_2O + CO_2$

3.65g

M_r HCl $= 1 + 35.5 = 36.5$

moles HCl $= \dfrac{\text{Mass (g)}}{M_r} = \dfrac{3.65}{36.5} = 0.1$ mol

ratio 2 HCl : 1 MgO

$0.1 : \dfrac{0.1}{2}$

$0.1 : 0.05$

M_r MgO $= 24 + 16 = 40$

mass MgO = moles × M_r $= 0.05 \times 40 = 2$ g

Limiting reactant

In a chemical reaction between two reactants if one reactant is in excess the other reactant is the **limiting reactant** and is completely used up.

Example

Iron(III) oxide reacts with carbon monoxide as shown below to produce iron. 10 moles of iron oxide (Fe_2O_3) is reacted with 40 moles of carbon monoxide (CO).

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Which reactant is in excess? Calculate the number of moles of iron formed.

Answer

If there are 10 moles of iron(III) oxide, then according to the ratio (3 × 10) = 30 moles of CO are needed to completely react with it. There are 40 moles of CO present so there are 10 moles of CO in excess. The iron oxide is the limiting reactant and is completely used up. One mole of iron(III) oxide produces two moles of iron, so 10 moles of iron(III) oxide produces 20 moles of iron.

Percentage yield

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

There are many reasons why the percentage yield is less than 100%.

1. Some of the product may be lost when it is separated from the reaction mixture. It may be left on the apparatus. This is often called loss by mechanical transfer.
2. Side reactions may occur.
3. Some reactions are reversible and do not go to completion.

Example
In a reaction to produce a compound, the theoretical yield was 20 g, but the mass of compound produced was 5 g. Calculate the percentage yield of the compound.

Answer

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{5}{20} \times 100 = 20\ \%$$

The **theoretical yield** can be worked out using the method of reacting mass calculations.

Example
Calculate the theoretical yield of calcium chloride when 5.6 g of calcium oxide reacts with excess hydrochloric acid. If 6 g of calcium chloride is produced calculate the percentage yield.

Answer

$$CaO + 2HCl \rightarrow CaCl_2 + H_2O$$

5.6g

M_r CaO = 40 + 16 = 56

$$\text{Moles CaO} = \text{moles} = \frac{\text{mass (g)}}{M_r} = \frac{5.6}{56} = 0.1\ \text{mol}$$

Ratio 1 CaO : 1 $CaCl_2$

0.1 : 0.1

M_r $CaCl_2$ = 40 + 35.5 × 2 = 111

mass = moles × M_r = 0.1 × 111 = 11.1 g

Theoretical yield = 11.1 g

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{6}{11.1} \times 100 = 54.1\ \%$$

Revision Questions

1.

Magnesium compounds have many important and wide-ranging uses. Magnesium nitrate is used as a fertiliser and is also present in many cosmetics including hair conditioner.

On heating, magnesium nitrate breaks down according to the equation below:

$$2Mg(NO_3)_2(s) \rightarrow 2MgO(s) + 4NO_2(g) + O_2(g)$$

(a) What term is used to describe a reaction in which a substance breaks down when heated?

____ [2]

(b) Calculate the mass of nitrogen dioxide, NO_2, produced when 4.44 g of magnesium nitrate are heated.

[5]

2. Today, most teeth whitening kits contain the chemical carbamide peroxide which breaks down in the mouth into urea and hydrogen peroxide. During the bleaching process the hydrogen peroxide decomposes to produce water and oxygen.

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

Calculate the mass of oxygen produced from 5.1 g of hydrogen peroxide.

[5]

3. Magnesium sulfate is manufactured from magnesium carbonate and sulfuric acid.

$$MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + CO_2 + H_2O$$

1.26 g of magnesium carbonate are reacted with 1.75 g of sulfuric acid and 1.35 g of magnesium sulfate were obtained.

(a) Calculate the moles of magnesium carbonate and sulfuric acid.

[2]

(b) Which reactant is in excess?

____ [1]

(c) Calculate the mass of magnesium sulfate formed.

[1]

(d) Calculate the percentage yield of magnesium sulfate.

[2]