

Element number 4

Beryllium: emeralds, alloys and nuclear reactors

Atomic number: 4

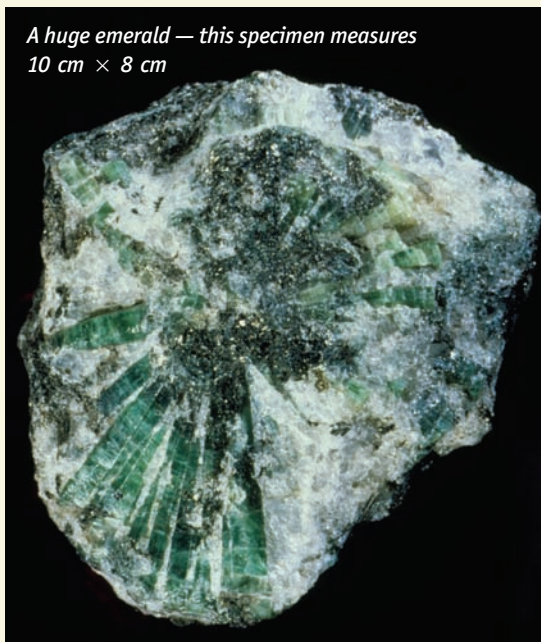
Electron configuration: $1s^2 2s^2$

Symbol: Be

Beryllium is a low-density metal in group 2 of the periodic table. It burns brightly when heated in air, but at room temperature is protected by a layer of oxide. The metal is amphoteric, reacting slowly with acid and rapidly with alkali.

Beryllium occurs in the mineral beryl, which is a beryllium aluminosilicate, $\text{Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}$. This mineral often occurs as huge crystals. The largest, found in South Dakota, was over 8 m long and yielded 61 tonnes of beryl.

*A huge emerald — this specimen measures
10 cm × 8 cm*



ROBERTO DE GUGLIELMO / SCIENCE PHOTO LIBRARY

Emeralds, aquamarine and other gemstones

Beryl is colourless, but impurities cause it to be coloured. Emeralds are crystals of beryl that contain small quantities of green chromium oxide. The best emeralds, found in Brazil, are clear and deep green. Poor quality emeralds are slightly cloudy and appear dull. Artificial emeralds can be made by melting beryl, adding traces of chromium(III) oxide and then allowing the liquid to crystallise slowly.

Aquamarine is a sea-green gemstone which is beryl with iron(III) oxide as an impurity. Rose-coloured beryl (morganite) owes its beautiful pale colour to traces of manganese(IV) oxide.

Beryllium in metallurgy

Beryllium is transparent to X-rays and is used as a window in X-ray tubes. Because of its low density, it is alloyed with other metals to make lightweight components in the space and aero industries. Beryllium and copper form an alloy with high tensile strength,

which is used in watch springs and springs that carry electric current.

Inhalation of beryllium compounds is dangerous. The high-charge density of the small +2 ion causes it to bind firmly to the active sites of some enzymes, inhibiting their activity.

Beryllium as a moderator

In nuclear fission, fissile nuclei (e.g. ^{235}U) are bombarded with slow-moving (thermal) neutrons. When the nucleus shatters, it emits several high-speed neutrons. To increase the chance of further fission and maintain the chain reaction, these neutrons have to be slowed down. When beryllium rods are inserted into a reactor, the neutrons are slowed down and the fission reaction speeds up; as they are withdrawn, the reaction slows down. Elements that slow down neutrons in a nuclear reactor are called moderators. Graphite and heavy water (deuterium oxide, D_2O) are also moderators.

Calculations from chemical equations

Chapter 4

Introduction

Calculations from chemical equations depend upon the ability to convert masses and volumes into amount of substance (moles) and to use the stoichiometry of an equation.

Calculations must be well laid out, with words explaining what is being calculated at each step. A series of numbers is not sufficient.

It is advisable to include units in calculations, rather than just adding them at the end. This makes it clear that you have used 'formulae' such as 'moles = concentration \times volume' correctly and that you have used comparable units, such as dm^3 and not cm^3 for the volume, when the concentration is in mol dm^{-3} . This method of adding units to the calculation is slower at first, but with practice becomes second nature and is more likely to avoid careless mistakes.

Significant figures

Answers to calculations should be given to an appropriate number of significant figures. The way to do this is to analyse the data:

- Count the number of significant figures in each quantity given in the question.
- Give the answer to the same number of significant figures as the data.
- If there are two pieces of data with different numbers of significant figures (e.g. one given to three significant figures and the other to four), give the answer to the lower number of significant figures.

- e** It is a good idea to check all your calculations.
- First check that you have worked out the molar masses correctly.
 - Then redo the calculations to make sure that you have not entered any wrong figures into your calculator.
 - Check your units and finally make sure that you have given the answers to the correct number of significant figures.
 - Check that your answer makes sense by doing a mental 'ball park' calculation as well.

The way to work out the correct number of significant figures is based on these rules:

- For whole numbers and decimals of numbers greater than 1, all the figures, *including any zeros at the end*, are counted. For example:

e You *must* show all your working or you will not score full marks. Examiners mark consequentially, so if you make a mistake early on, you can still score the remaining marks, provided that you have carried out the subsequent steps correctly and your working is clear.

e If the question specifies a number of significant figures, that number should be given in your answer. Otherwise, give your answer to three significant figures.

e It is a good idea to keep all the figures on your calculator throughout the calculation or to write down all intermediate answers to four significant figures.

- the number 102 is written to three significant figures (3 s.f.)
- the number 1.02 is also written to 3 s.f.
- the number 1.020 is written to 4 s.f.
- the number 10.20 is also written to 4 s.f.
- the measurement 100 cm^3 is expressed to 3 s.f.
- For decimals in numbers less than 1, count all the numbers, except the zeros *before* and *immediately after* the decimal point. For example:
 - the number 0.102 is written to 3 s.f.
 - the number 0.1020 is written to 4 s.f.
 - the number 0.00102 is written to 3 s.f, not 5 s.f.
 - the number 0.01200 is written to 4 s.f.
- For values written in scientific notation, all the numbers before and after the decimal place are counted as significant figures. The power of 10 is not counted. For example:
 - the number 1.02×10^{-3} is written to 3 s.f.

Mass and volume calculations

Mass-to-mass calculations

A typical question involving mass-to-mass calculation is to ask for the mass of a product that could be obtained from a given mass of one of the reactants. For example, calculate the mass of barium sulfate that is precipitated when a solution containing 5.55 g of barium chloride, BaCl_2 , is reacted with excess magnesium sulfate. (Molar mass of barium chloride = $137.3 + (2 \times 35.5) = 208.3 \text{ g mol}^{-1}$)

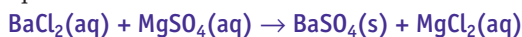
After the equation has been written, the calculation should be done in three steps:

Step 1: calculate the amount (moles) of reactant. In this example, the reactant is barium chloride.

Step 2: use the stoichiometry of the equation to calculate the amount (moles) of product.

Step 3: convert the moles of product to mass.

The equation is:



Step 1: amount of barium chloride = $\frac{\text{mass}}{\text{molar mass}} = \frac{5.55 \text{ g}}{208.3 \text{ g mol}^{-1}} = 0.02664 \text{ mol}$

Step 2: ratio $\text{BaSO}_4:\text{BaCl}_2$ is 1:1

amount of BaSO_4 produced = amount of $\text{BaCl}_2 = 0.02664 \text{ mol}$

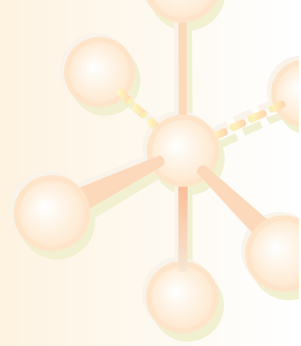
Step 3: molar mass of $\text{BaSO}_4 = 137.3 + 32.1 + (4 \times 16.0) = 233.4 \text{ g mol}^{-1}$

mass of barium sulfate produced = mol \times molar mass
= $0.02664 \text{ mol} \times 233.4 \text{ g mol}^{-1} = 6.22 \text{ g}$

e The answer must be given to three significant figures because the mass of barium chloride was given to three significant figures. Never round down to one or two significant figures in the middle of a calculation.

e The stoichiometry of an equation is the ratio of the numbers of each reactant and product in the balanced equation.

e When the stoichiometric ratio is 1:1, many candidates fail to make the important point that the amounts (in moles) of the two substances are the same. This *must* be stated clearly, as in the examples below.



The method can be illustrated by a flow diagram. The calculation is to find the mass of substance B produced from (or reacting with) substance A:

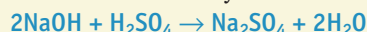


Steps 1 and 3 involve mass to moles and moles to mass conversions (pages 56–57). Step 2 uses the stoichiometry of the equation.

When the ratio in the equation is not 1:1, care must be taken in step 2.

Worked example

Calculate the mass of sodium sulfate produced when 3.45 g of sodium hydroxide is neutralised by dilute sulfuric acid. The equation is:



Answer

Step 1: molar mass of NaOH = 23.0 + 16.0 + 1.0 = 40.0 g mol⁻¹

$$\text{amount (moles) of sodium hydroxide} = \frac{\text{mass}}{\text{molar mass}} = \frac{3.45 \text{ g}}{40 \text{ g mol}^{-1}} = 0.08625 \text{ mol}$$

Step 2: ratio of Na₂SO₄:NaOH is 1:2, so *fewer* moles of sodium sulfate produced

$$\text{amount of sodium sulfate} = \frac{1}{2} \times 0.08625 \text{ mol} = 0.04313 \text{ mol}$$

Step 3: molar mass of Na₂SO₄ = (2 × 23.0) + 32.1 + (4 × 16.0) = 142.1 g mol⁻¹

$$\begin{aligned} \text{mass of sodium sulfate produced} &= \text{moles} \times \text{molar mass} \\ &= 0.04313 \text{ mol} \times 142.1 \text{ g mol}^{-1} = 6.13 \text{ g} \end{aligned}$$

e In calculations at AS, the equation will probably be given; you have to calculate the molar masses of two substances.

Volume of gas calculations

These are based on the fact that the volume occupied by 1 mol of all gases at a given temperature and pressure is the same. This volume is called the molar volume.

The molar volume of a gas is the volume occupied by 1 mol of the gas under specified conditions of temperature and pressure.

- Under typical laboratory conditions, molar volume is usually given as 24 dm³ mol⁻¹ = 24 000 cm³ mol⁻¹.
- At 0°C and 1 atm pressure it is 22.4 dm³ mol⁻¹ = 22 400 cm³ mol⁻¹.

The relationship between amount of gas (moles) and volume is given by:

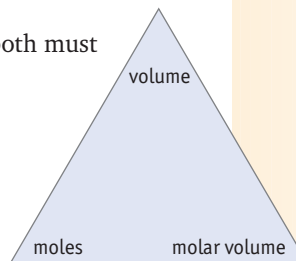
$$\text{amount (moles) of gas} = \frac{\text{volume of gas}}{\text{molar volume}}$$

Make sure that the units of the two volumes are the same; both must be either dm³ or cm³.

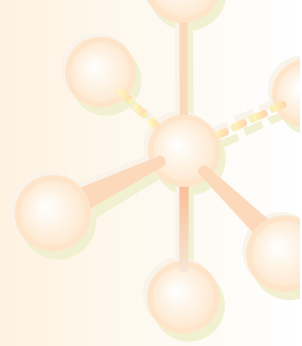
Again a 'mole triangle' can be used:

$$\text{volume of gas} = \text{moles} \times \text{molar volume}$$

$$\text{moles} = \frac{\text{volume of gas}}{\text{molar volume}}$$



e The value of the molar volume will always be given in the question.



Mass to volume of gas calculations

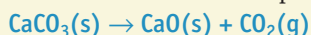
The route for this type of calculation is similar to that for mass-to-mass calculations.



Steps 1 and 3 use the conversions mass to moles and moles to volume of gas. Step 2 uses the stoichiometry of the equation.

Worked example 1

Calculate the volume of carbon dioxide produced when 2.68 g of calcium carbonate is heated and decomposes according to the equation:



Under the conditions of the experiment, 1 mol of gas occupies 24 000 cm³.

Answer

Step 1: molar mass of CaCO₃ = 40.1 + 12.0 + (3 × 16.0) = 100.1 g mol⁻¹

$$\begin{aligned} \text{amount (moles) of calcium carbonate} &= \frac{\text{mass}}{\text{molar mass}} \\ &= \frac{2.68 \text{ g}}{100.1 \text{ g mol}^{-1}} = 0.0268 \text{ mol} \end{aligned}$$

Step 2: ratio of CO₂ to CaCO₃ is 1:1

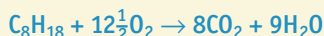
$$\text{amount (moles) of carbon dioxide} = 0.0268 \text{ mol}$$

Step 3: volume of carbon dioxide gas = moles × molar volume

$$= 0.0268 \text{ mol} \times 24\,000 \text{ cm}^3 \text{ mol}^{-1} = 643 \text{ cm}^3$$

Worked example 2

The average family motorist uses about 1000 kg of petrol each year. Assume that the molecular formula of petrol is C₈H₁₈ and that the equation for the reaction is:



Calculate:

- the yearly volume of carbon dioxide produced at room temperature and pressure
(The molar volume of gas at room temperature and pressure is 24 dm³ mol⁻¹.)
- the mass of carbon dioxide (the carbon footprint) produced by the car in a year

Answer

a Step 1: molar mass of C₈H₁₈ is (8 × 12.0) + (18.0 × 1.0) = 114.0 g mol⁻¹

$$\text{amount of petrol} = \frac{\text{mass}}{\text{molar mass}} = \frac{1\,000\,000 \text{ g}}{114.0 \text{ g mol}^{-1}} = 8772 \text{ mol}$$

Step 2: ratio of CO₂:C₈H₁₈ = 8:1 (so more moles of CO₂ produced)

$$\text{amount of CO}_2 \text{ produced} = 8 \times 8772 \text{ mol} = 7.02 \times 10^4 \text{ mol}$$

$$\begin{aligned} \text{Step 3: volume of carbon dioxide produced} &= 7.02 \times 10^4 \text{ mol} \times 24 \text{ dm}^3 \text{ mol}^{-1} \\ &= 1.68 \times 10^6 \text{ dm}^3 \end{aligned}$$

- b** 1 mol of CO_2 has a mass of 44.0 g
 7.02×10^4 mol of CO_2 has a mass of $7.02 \times 10^4 \times 44.0 \text{ g} = 3.09 \times 10^6 \text{ g}$
 $= 3.09$ tonne
 The family car has, on average, a carbon footprint of just over 3 tonnes.

Volume of gas to volume of another gas calculations

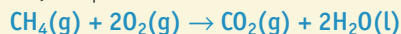
There are two alternative methods for performing this type of calculation.

It can be done in three steps, similarly to mass-to-volume calculations.

Volume of gas A $\xrightarrow{\text{step 1}}$ Moles of gas A $\xrightarrow{\text{step 2}}$ Moles of gas B $\xrightarrow{\text{step 3}}$ Volume of gas B

Worked example

Calculate the volume of oxygen needed to react with 123 cm^3 of gaseous methane, CH_4 .



Answer

Step 1: amount of methane = $\frac{\text{volume}}{\text{molar volume}} = \frac{123 \text{ cm}^3}{24\,000 \text{ cm}^3 \text{ mol}^{-1}} = 0.005125 \text{ mol}$

Step 2: the ratio of oxygen to methane is 2:1

amount of oxygen = $2 \times 0.005125 \text{ mol} = 0.01025 \text{ mol}$

Step 3: volume of oxygen = moles \times molar volume
 $= 0.01025 \text{ mol} \times 24\,000 \text{ cm}^3 \text{ mol}^{-1} = 246 \text{ cm}^3$

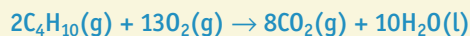
The alternative, and simpler method, is to use Avogadro's hypothesis.

Equal volumes of gases, measured at the same temperature and pressure, contain the same number of molecules.

This means that if there is twice the number of moles, the volume occupied is doubled — as long as both substances are gases.

Worked example

Calculate the volume of oxygen gas needed to burn completely 200 cm^3 of gaseous butane.



Answer

Both butane and oxygen are gases, therefore:

$$\frac{\text{volume of oxygen}}{\text{volume of butane}} = \frac{\text{moles of oxygen}}{\text{moles of butane}} = \frac{13}{2}$$

$$\begin{aligned} \text{volume of oxygen} &= \frac{13}{2} \times \text{volume of butane} \\ &= \frac{13}{2} \times 200 \text{ cm}^3 = 1300 \text{ cm}^3 \end{aligned}$$

e You may be given the molar volume for the gases. If not, assume that it is $24\,000 \text{ cm}^3 \text{ mol}^{-1}$ or give it a value of $V \text{ cm}^3$. The value, V , will cancel.

e This method can only be used if, under the conditions of the experiment, both substances are gases. You do not need to know the value of the molar volume.

Percentage yield calculations

Many reactions do not produce the same amount of product as that calculated from the chemical equation. This is caused by the reaction being reversible (so equilibrium is reached) or because of competing reactions. This is especially true in organic chemistry.

Percentage yield is defined as:

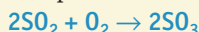
$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

The actual yield is the measured mass of the product obtained in the experiment.

The theoretical yield is the mass that is calculated from the equation for the reaction, assuming that all the reactant is converted into the product.

Worked example

When 1000 g of sulfur dioxide is reacted with excess oxygen, 1225 g of sulfur trioxide is produced:



Calculate the percentage yield.

Answer

First, calculate the theoretical yield, using the method on page 280.

Step 1: molar mass of $\text{SO}_2 = 32.1 + (2 \times 16.0)$

$$= 64.1 \text{ g mol}^{-1}$$

$$\text{amount of sulfur dioxide} = \frac{\text{mass}}{\text{molar mass}}$$

$$= \frac{1000 \text{ g}}{64.1 \text{ g mol}^{-1}} = 15.60 \text{ mol}$$

Step 2: ratio of $\text{SO}_2:\text{SO}_3$ is 2:2 = 1:1

$$\begin{aligned} \text{theoretical amount of sulfur trioxide produced} &= \text{amount of } \text{SO}_2 \text{ reacted} \\ &= 15.60 \text{ mol} \end{aligned}$$

Step 3: molar mass of $\text{SO}_3 = 32.1 + (3 \times 16.0) = 80.1 \text{ g mol}^{-1}$

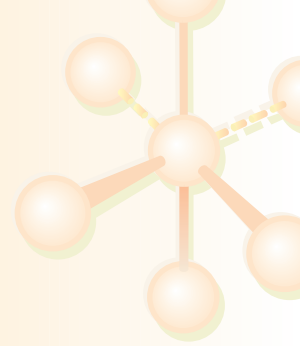
$$\begin{aligned} \text{theoretical yield} &= \text{moles} \times \text{molar mass} \\ &= 15.60 \text{ mol} \times 80.1 \text{ g mol}^{-1} \\ &= 1250 \text{ g} \end{aligned}$$

Second, use the theoretical yield to calculate the percentage yield:

$$\begin{aligned} \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 &= \frac{1225}{1250} \times 100 \\ &= 98.0\% \end{aligned}$$

Calculation of reaction stoichiometry

If the masses of both reactants (or of one reactant and one product) are known, the stoichiometry of the equation can be worked out. This is done by converting the masses to amounts and examining the ratio of these amounts.

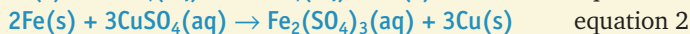
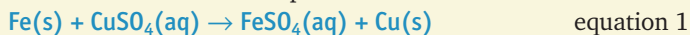


e The actual and theoretical yields are usually masses (grams), but *both* could be in moles.

e Do *not* calculate the percentage yield as:
$$\frac{\text{mass of product} \times 100}{\text{mass of reactant}}$$

Worked example 1

3.48 g of pure iron was placed in excess copper(II) sulfate solution and stirred until all reaction had ceased. The residue of copper was filtered off, washed and dried. It had a mass of 3.95 g. Use these data to work out which of the two reactions below took place.

**Answer**

$$\begin{aligned} \text{amount of iron reacted} &= \frac{\text{mass}}{\text{molar mass}} \\ &= \frac{3.48 \text{ g}}{55.8 \text{ g mol}^{-1}} = 0.0624 \text{ mol} \end{aligned}$$

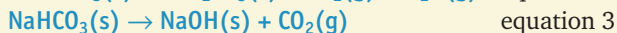
$$\begin{aligned} \text{amount of copper produced} &= \frac{\text{mass}}{\text{molar mass}} \\ &= \frac{3.95 \text{ g}}{63.5 \text{ g mol}^{-1}} = 0.0622 \text{ mol} \end{aligned}$$

Within experimental error, these numbers are in the ratio of 1:1. Therefore, equation 1 is correct. (Equation 2 has a ratio of moles of iron:moles of copper of 2:3.)

Worked example 2

When heated, sodium hydrogencarbonate, NaHCO_3 , decomposes leaving a solid residue. When 4.20 g of sodium hydrogencarbonate was heated until there was no further loss in mass, 2.65 g of solid was left.

Three equations have been suggested for this reaction:



Which of these equations is consistent with the experimental data?

Answer

$$\text{molar masses: NaHCO}_3 = 23.0 + 1.0 + 12.0 + (3 \times 16.0) = 84.0 \text{ g mol}^{-1}$$

$$\text{Na}_2\text{O} = (2 \times 23.0) + 16.0 = 62.0 \text{ g mol}^{-1}$$

$$\text{Na}_2\text{CO}_3 = (2 \times 23.0) + 12.0 + (3 \times 16.0) = 106.0 \text{ g mol}^{-1}$$

$$\text{NaOH} = 23.0 + 16.0 + 1.0 = 40.0 \text{ g mol}^{-1}$$

$$\text{amount of NaHCO}_3 \text{ taken} = \frac{4.20 \text{ g}}{84.0 \text{ g mol}^{-1}} = 0.0500 \text{ mol}$$

$$\text{equation 1: amount of Na}_2\text{O produced} = 0.5 \times 0.0500 = 0.0250 \text{ mol}$$

$$\text{mass of Na}_2\text{O produced} = 0.0250 \text{ mol} \times 62.0 \text{ g mol}^{-1} = 1.55 \text{ g}$$

This does not equal 2.65 g, so equation 1 is incorrect

$$\text{equation 2: amount of Na}_2\text{CO}_3 \text{ produced} = 0.5 \times 0.0500 = 0.0250 \text{ mol}$$

$$\text{mass of Na}_2\text{CO}_3 \text{ produced} = 0.0250 \text{ mol} \times 106.0 \text{ g mol}^{-1} = 2.65 \text{ g}$$

This is the same as the experimental mass produced, so equation 2 is in agreement with the experimental data.

Chemicals react in simple whole number ratios by moles. Therefore, 0.0624:0.0622 is a mole ratio of 1:1.

The ratio of NaHCO_3 to solid product in equations 1 and 2 is 2:1, so the number of moles of solid formed is half the number of moles of reactant.

Equation 3 predicts that 2.00 g of solid would be left, so it cannot be the correct equation.

Limiting reagent

When a reaction is carried out in the laboratory, the reactants are not always present in the exact stoichiometric ratio determined by the equation. As a result, one reactant is used completely; some of the other reactant is left over. The reagent left over is said to be in excess; the one used completely is said to be the **limiting reagent**.

A limiting reagent is the substance that determines the theoretical yield of product in a reaction.

An analogy is a factory producing sunglasses (S). Every frame (F) needs two lenses (L). The 'equation' for the process is:



If the factory owner buys 144 frames and 280 lenses, the maximum number of sunglasses that can be produced is limited by the number of lenses. One hundred and forty-four frames need 288 lenses. There are only 280 lenses, so the lenses are the limiting factor and only 140 sunglasses can be made.

The lenses are the limiting 'reagent' and the frames are the 'reagent' in excess.

To identify the limiting reagent and hence calculate the theoretical yield, use the following method.

Step 1: calculate the amount (in moles) of one reagent and use the reaction stoichiometry to calculate the amount (in moles) of product that could be formed from this reagent.

Step 2: calculate the amount (in moles) of the second reagent and use the reaction stoichiometry to calculate the amount (in moles) of product that could be formed from this second reagent.

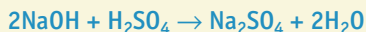
Step 3: the reagent that produces the *least* amount (in moles) of product is the limiting reagent.

Step 4: calculate the theoretical yield of the product from the *least* amount (in moles) of product calculated in steps 1 and 2, i.e. from the limiting reagent.

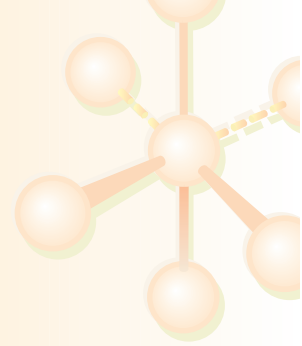
In steps 1 and 2 you must be able to convert mass or volume of reactant into moles and then use the ratio of numbers of moles of reactant to product in the equation (the stoichiometry) to calculate the amount of product. Step 4 is the conversion of amount (moles) of product into mass or volume of product.

Worked example 1

Solutions containing 12.8 g of sulfuric acid (molar mass 98.1 g mol⁻¹) and 10.0 g of sodium hydroxide (molar mass 40.0 g mol⁻¹) are mixed and produce sodium sulfate and water according to the following equation:



Calculate the mass of sodium sulfate (molar mass 142.1 g mol⁻¹) produced.



Note the analogy to the mole here. The number of moles of a chemical is a measure of the (very large) number of molecules of that chemical.

Answer

$$\text{Step 1: amount of NaOH} = \frac{\text{mass}}{\text{molar mass}} = \frac{10.0 \text{ g}}{40.0 \text{ g mol}^{-1}} = 0.250 \text{ mol}$$

ratio of Na_2SO_4 to $\text{NaOH} = 1:2$

$$\text{theoretical amount of Na}_2\text{SO}_4 \text{ that would be produced} = \frac{1}{2} \times 0.250 = 0.125 \text{ mol}$$

$$\text{Step 2: amount of H}_2\text{SO}_4 = \frac{\text{mass}}{\text{molar mass}} = \frac{12.8 \text{ g}}{98.1 \text{ g mol}^{-1}} = 0.130 \text{ mol}$$

ratio Na_2SO_4 to $\text{H}_2\text{SO}_4 = 1:1$

theoretical amount of Na_2SO_4 that would be produced = 0.130 mol

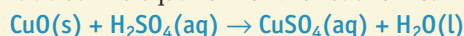
Step 3: the reagent that produces the least product (0.125 mol) is sodium hydroxide, so that is the limiting reagent.

$$\begin{aligned} \text{Step 4: mass of Na}_2\text{SO}_4 \text{ produced} &= \text{moles} \times \text{molar mass} \\ &= 0.125 \text{ mol} \times 142.1 \text{ g mol}^{-1} \\ &= 17.8 \text{ g} \end{aligned}$$

Even though there are more moles of sodium hydroxide, it is the limiting reagent. This is because of the stoichiometry of the equation — 2 moles of sodium hydroxide are required to react with 1 mole of sulfuric acid.

Worked example 2

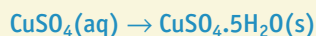
7.95 g of copper(II) oxide is mixed with a solution containing 7.35 g of sulfuric acid. The equation for the reaction is:



a Determine which is the limiting reagent.

b Hence, calculate the mass of CuSO_4 produced.

c On evaporation of the solution, 16.3 g of crystals of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ are formed:



Calculate the percentage yield.

Answer

$$\text{a molar mass of CuO} = 63.5 + 16.0 = 79.5 \text{ g mol}^{-1}$$

$$\text{amount of CuO} = \frac{7.95 \text{ g}}{79.5 \text{ g mol}^{-1}} = 0.100 \text{ mol}$$

The ratio $\text{CuO}:\text{CuSO}_4$ is 1:1, so 0.100 mol of CuSO_4 would be produced.

$$\text{molar mass of H}_2\text{SO}_4 = (2 \times 1.0) + 32.1 + (4 \times 16.0) = 98.1 \text{ g mol}^{-1}$$

$$\text{amount of H}_2\text{SO}_4 = \frac{7.35 \text{ g}}{98.1 \text{ g mol}^{-1}} = 0.0750 \text{ mol}$$

The ratio $\text{H}_2\text{SO}_4:\text{CuSO}_4$ is 1:1, so 0.0750 mol of CuSO_4 would be produced.

This is the smaller amount, so sulfuric acid is the limiting reagent.

$$\text{b amount of CuSO}_4 \text{ produced} = 0.0750 \text{ mol}$$

$$\text{molar mass of CuSO}_4 = 63.5 + 32.1 + (4 \times 16.0) = 159.6 \text{ g mol}^{-1}$$

$$\text{mass of CuSO}_4 \text{ produced} = 0.0750 \text{ mol} \times 159.6 \text{ g mol}^{-1} = 12.0 \text{ g}$$

c 1 mol of $\text{CuSO}_4(\text{aq})$ produces 1 mol $\text{CuSO}_4 \cdot 5\text{H}_2\text{O(s)}$.

$$\text{molar mass of CuSO}_4 \cdot 5\text{H}_2\text{O} = 159.6 + (5 \times 18.0) = 249.6 \text{ g mol}^{-1}$$

$$\text{theoretical yield} = 0.0750 \text{ mol} \times 249.6 \text{ g mol}^{-1} = 18.7 \text{ g}$$

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{16.3}{18.7} \times 100 = 87.2\%$$

Concentration

A **solution** consists of a substance that is dissolved — the **solute** — and the substance that is doing the dissolving — the **solvent**. For example, when a salt is dissolved in water, the salt is the solute and the water is the solvent. Solutions may be coloured or colourless, but they are always clear and never cloudy.

Units of concentration

Solutions can contain different amounts of solute up to a maximum value, which is called the **solubility** of the solute. Therefore, the same volume of solutions of a solute can contain different amounts of that solute. To enable the amount or the mass of solute in a given volume to be determined, the concentration of the solution must be known.

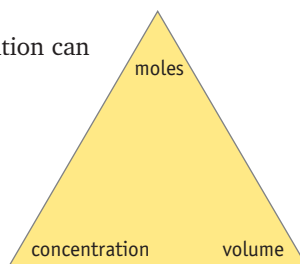
- The most common unit of concentration is **mol dm⁻³**.
- Mol dm⁻³ is sometimes called molarity (symbol, *M*).
- To calculate a concentration in mol dm⁻³, the amount (moles) of solute is divided by the volume of the solution in dm³.

Conversions such as amount of solute to concentration can be performed using a version of the 'mole triangle':

$$\text{moles} = \text{concentration} \times \text{volume}$$

$$\text{concentration} = \frac{\text{moles}}{\text{volume}}$$

$$\text{volume} = \frac{\text{moles}}{\text{concentration}}$$



- Another unit of concentration that is sometimes used is **g dm⁻³**.
- This is the mass of the solute divided by the volume (in dm³) of the solution.
- The conversion of concentrations in mol dm⁻³ to g dm⁻³ is the same as converting mol to mass, which is to multiply by the molar mass of the solute.
- To convert concentrations in g dm⁻³ to mol dm⁻³, divide by the molar mass of the solute.

Worked example 1

- a** Calculate the concentration in mol dm⁻³ of a solution made by dissolving 0.123 mol of sodium hydroxide in water and making up the solution to a total volume of 250 cm³.
- b** Calculate the concentration of this solution in g dm⁻³.

Answer

a volume of solution = $\frac{250}{1000} = 0.250 \text{ dm}^3$

$$\text{concentration} = \frac{\text{moles of solute}}{\text{volume of solution}} = \frac{0.123 \text{ mol}}{0.250 \text{ dm}^3} = 0.492 \text{ mol dm}^{-3}$$

b molar mass of NaOH = 23.0 + 16.0 + 1.0 = 40.0 g mol⁻¹

$$\text{concentration of solution} = 0.492 \text{ mol dm}^{-3} \times 40.0 \text{ g mol}^{-1} = 19.7 \text{ g dm}^{-3}$$

◀ The volume is the volume of the solution, not of the solvent.

Mol cancels with mol⁻¹, leaving the final unit as g dm⁻³. This is a useful check to make sure that you have carried out the conversion correctly.

Worked example 2

Calculate the mass of hydrated sodium carbonate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, that is required to make up 250 cm^3 of a $0.100 \text{ mol dm}^{-3}$ solution.

Answer

amount of sodium carbonate required = concentration \times volume of solution

$$\text{amount} = 0.100 \text{ mol dm}^{-3} \times \frac{250 \text{ dm}^3}{1000} = 0.0250 \text{ mol}$$

$$\begin{aligned} \text{molar mass of } \text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} &= (2 \times 23.0) + 12.0 + (3 \times 16.0) + (10 \times 18.0) \\ &= 286.0 \text{ g mol}^{-1} \end{aligned}$$

$$\text{mass required} = \text{moles} \times \text{molar mass} = 0.0250 \text{ mol} \times 286.0 \text{ g mol}^{-1} = 7.15 \text{ g}$$

In the question, the volume of the solution was given in cm^3 . Therefore, it had to be converted to dm^3 .

Worked example 3

Calculate the amount of sulfuric acid in 22.4 cm^3 of $0.0502 \text{ mol dm}^{-3}$ solution.

Answer

amount (in moles) of sulfuric acid = concentration \times volume

$$\text{amount} = 0.0502 \text{ mol dm}^{-3} \times \frac{22.4 \text{ dm}^3}{1000} = 0.00112 \text{ mol}$$

Worked example 4

What volume of sodium hydroxide solution of concentration $0.100 \text{ mol dm}^{-3}$ contains 0.00250 mol of sodium hydroxide?

Answer

$$\text{volume of sodium hydroxide solution} = \frac{\text{moles}}{\text{concentration}}$$

$$\text{volume} = \frac{0.00250 \text{ mol}}{0.100 \text{ mol dm}^{-3}} = 0.0250 \text{ dm}^3 = 25.0 \text{ cm}^3$$

Concentrations of very dilute solutions

There are two units used commonly:

- nanomoles per dm^3 (nmol dm^{-3}); $1 \text{ nmol} = 10^{-9} \text{ mol}$
- parts per million (ppm)

'Parts per million' is normally the mass of a solute per million grams of solvent. (The assumption is made that the masses of the solvent and the solution are the same because the solution is very dilute.)

Worked example 1

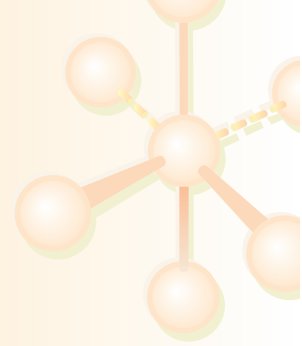
Aluminium sulfate is used as a coagulant in domestic water. Analysis of some water found that the aluminium ion concentration was 6.3 nmol dm^{-3} . Calculate the mass of aluminium ions in 1 dm^3 of the treated water.

Answer

$$6.3 \text{ nmol} = 6.3 \times 10^{-9} \text{ mol}$$

$$\text{molar mass of } \text{Al}^{3+} = 27.0 \text{ g mol}^{-1}$$

$$\begin{aligned} \text{mass of } \text{Al}^{3+} \text{ ions in the water} &= 6.3 \times 10^{-9} \text{ mol} \times 27.0 \text{ g mol}^{-1} \\ &= 1.70 \times 10^{-7} \text{ g} (= 1.70 \times 10^{-4} \text{ mg} = 170 \text{ ng}) \end{aligned}$$



Worked example 2

A bottle of mineral water was labelled as containing 2.0 ppm of sodium ions.

- a** Calculate its concentration in nmol dm^{-3} .
b Calculate the number of sodium ions in 1.0 dm^3 of the mineral water.
(Avogadro constant = $6.02 \times 10^{23} \text{ mol}^{-1}$)

Answer

- a** 1 000 000 g of water contain 2.0 g of sodium ions

1000 g of water has a volume of 1.0 dm^3 and contains

$$\frac{2.0}{1000 \text{ g}} = 0.0020 \text{ g of sodium ions}$$

$$\begin{aligned} \text{amount (moles) of sodium ions} &= \frac{0.0020 \text{ g}}{23.0 \text{ g mol}^{-1}} \\ &= 8.7 \times 10^{-5} \text{ mol} = 8.7 \times 10^4 \text{ nmol} \end{aligned}$$

The concentration of sodium ions is $8.7 \times 10^4 \text{ nmol dm}^{-3}$.

- b** number of ions = moles \times Avogadro constant
 $= 8.7 \times 10^{-5} \text{ mol} \times 6.02 \times 10^{23} \text{ mol}^{-1} = 5.2 \times 10^{19}$

1 mol = 1×10^9 nmol,
so $8.7 \times 10^{-5} \text{ mol}$
 $= 8.7 \times 10^{-5} \times 1 \times 10^9$
 $= 8.7 \times 10^4 \text{ nmol}$

Questions

- 1** Calculate the mass of iron(III) hydroxide precipitated in the reaction between 12.7 g of iron(III) sulfate and excess sodium hydroxide solution.
$$\text{Fe}_2(\text{SO}_4)_3(\text{aq}) + 6\text{NaOH}(\text{aq}) \rightarrow 2\text{Fe}(\text{OH})_3(\text{s}) + 3\text{Na}_2\text{SO}_4(\text{aq})$$
- 2** Copper reacts with silver nitrate solution according to the equation:
$$\text{Cu}(\text{s}) + 2\text{AgNO}_3(\text{aq}) \rightarrow \text{Cu}(\text{NO}_3)_2(\text{aq}) + 2\text{Ag}(\text{s})$$

Calculate the mass of copper needed to react with a solution containing 12.6 g of silver nitrate.
- 3** Calculate the volume of oxygen produced when 33.3 g of sodium nitrate is heated. Sodium nitrate decomposes according to the equation:
$$2\text{NaNO}_3(\text{s}) \rightarrow 2\text{NaNO}_2(\text{s}) + \text{O}_2(\text{g})$$

Under the conditions of the experiment, 1 mol of gas occupies a volume of 25.0 dm^3 .
- 4** When 3000 g of hydrogen is reacted with excess nitrogen under high pressure and temperature in the presence of an iron catalyst, 2550 g of ammonia is produced.
$$\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$$

Calculate the theoretical yield and hence the percentage yield in the process.
- 5** Calculate the amount of sulfuric acid in 23.4 cm^3 of a $0.0545 \text{ mol dm}^{-3}$ solution of the acid.
- 6** Calculate the volume of a $0.106 \text{ mol dm}^{-3}$ solution of sodium hydroxide, which contains 0.00164 mol.
- 7** Calculate the mass of ethanedioic acid, $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$, that is needed to make 500 cm^3 of a $0.0500 \text{ mol dm}^{-3}$ solution.
- 8** 4.50 g of iron powder was added to 50.0 cm^3 of 2.00 mol dm^{-3} copper(II) sulfate solution. The equation for the reaction is:
$$\text{Fe}(\text{s}) + \text{CuSO}_4(\text{aq}) \rightarrow \text{Cu}(\text{s}) + \text{FeSO}_4(\text{aq})$$

a Calculate which reagent is limiting.
b Calculate the mass of copper produced.
c What observation would tell you that you were correct in identifying the limiting reagent?
- 9** 21.5 g of sodium hydroxide was added carefully to 500.0 cm^3 of 1.00 mol dm^{-3} sulfuric acid solution.
a Write the equation for the reaction.
b Calculate which reagent is limiting.
c Calculate the mass of sodium sulfate produced.
d What would be observed if pieces of red and blue litmus paper were added to the solution after the reaction?