Amount of substance

A mole of substance

Atoms are too small to be counted so chemists measure the amount of substance using moles. One mole of any substance contains the same number of particles as there are carbon atoms in 12.0 g of carbon-12.

MASS OF ONE MOLE OF A SUBSTANCE

The mass of one mole of any element is its relative atomic mass in grams.

One mole of carbon has a mass of 12.0 g.
One mole of sulfur has a mass of 32.1 g.
One mole of copper has a mass of 63.5 g.

PRACTICE QUESTION

1 Calculate the mass, in g, of:
- a 1.00 mole of magnesium
- b 1.00 mole of beryllium
- c 1.00 mole of aluminium.

MASS OF DIFFERENT AMOUNTS OF SUBSTANCE

To work out the mass of other amounts of substance, in moles, multiply the relative atomic mass in grams by the amount of substance required in moles.

2.00 moles of carbon has a mass of 12.0 g × 2 = 24.0 g
3.00 moles of sulfur has a mass of 32.1 g × 3 = 96.3 g
2.50 moles of copper has a mass of 63.5 g × 2.5 = 158.75 g

REMEMBER: The molar mass of a substance is the mass per mole of the substance. The units are $g\,mol^{-1}$.

PRACTICE QUESTION

2 Calculate the mass, in g, of:
- a 2.00 moles of magnesium
- b 1.50 moles of beryllium
- c 2.00 moles of aluminium
- d 5.00 moles of oxygen
- e 1.50 moles of neon.

EXAMPLES

The molar mass of carbon, C, is simply the relative atomic mass of carbon, which is 12.0 in units of $g\,mol^{-1}$. So the molar mass of carbon is $12.0\,g\,mol^{-1}$.

The molar mass of water, H_2O, is the total of the relative atomic masses of all the atoms that make up a relative formula unit of H_2O.

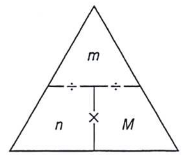
$(2 \times 1.0) + (1 \times 16.0) = 18.0\,g\,mol^{-1}$

PRACTICE QUESTION

3 Calculate the molar mass of:
- a sulfur, S
- b sulfur dioxide, SO_2
- c carbon dioxide, CO_2
- d boron trichloride, BCl_3
- e ammonia, NH_3.

LINKING MOLES, MASS AND MOLAR MASS

The number of moles (n), the mass of the substance (m) and the molar mass (M) are linked together using $n = \frac{m}{M}$



n = number of moles, in mol
m = mass of substance, in g
M = molar mass, in $g\,mol^{-1}$

1) Calculate the number of moles in 6.0 g of carbon.

$$n = \frac{m}{M} = \frac{6.0\,g}{12.0\,g\,mol^{-1}} = 0.5\text{ moles}$$

2) 0.5 moles of a substance has a mass of 22 g. Calculate the molar mass of the compound.

$$M = \frac{m}{n} = \frac{22.0\,g}{0.5\,mol} = 44.0\,g\,mol^{-1}$$

PRACTICE QUESTIONS

4 Calculate the amount, in mol, of:
- a 9.0 g of carbon, C
- b 36.45 g of magnesium, Mg
- c 76.0 g of fluorine, F_2.

5 Calculate the molar mass of a compound when:
- a 0.25 moles of a compound has a mass of 25.0 g
- b 0.10 moles of a compound has a mass of 4.4 g
- c 0.05 moles of a compound has a mass of 5.0 g.

Avogadro constant

The amount of substance

Chemists measure the amount of substance in moles. There is the same number of particles in one mole of any substance. This number of particles is known as the Avogadro constant. The Avogadro constant is a large number and is normally written in standard form as 6.02×10^{23} mol^{-1} to three significant figures). The symbol for the Avogadro constant is N_A.

FINDING THE NUMBER OF PARTICLES IN ONE MOLE

1) Calculate the number of atoms in 1.00 mole of carbon.
   1.00 mole of carbon contains 6.02×10^{23} atoms.

2) Calculate the number of atoms in 1.00 mole of sodium.
   1.00 mole of sodium contains 6.02×10^{23} atoms.

PRACTICE QUESTION

1 Calculate the number of atoms in:
- a 1.00 mole of lithium
- b 1.00 mole of tungsten
- c 1.00 mole of aluminium. Give your answer to three significant figures.

FINDING THE NUMBER OF ATOMS IN DIFFERENT AMOUNTS OF SUBSTANCE

To work out the number of atoms in other amounts of a substance, multiply the number of moles (n) by the Avogadro constant (N_A).

Number of atoms $= n \times N_A$

Calculate the number of atoms in 2.00 moles of magnesium atoms.

2.00 moles of magnesium atoms $= 2.00 \times 6.02 \times 10^{23} = 1.204 \times 10^{24}$ atoms

Notice that when you write the number in standard form the power is now $\times 10^{24}$.

PRACTICE QUESTION

2 Calculate the number of atoms in:
- a 0.10 moles of carbon
- b 2.50 moles of sulfur
- c 0.75 moles of magnesium. Give your answer to three significant figures.

Other types of particle

The Avogadro constant can be used to work out the number of particles in any type of substance.

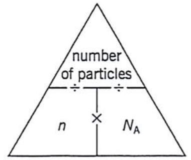
PRACTICE QUESTION

3 Calculate the number of particles in:
- a 1.00 mole of sodium ions
- b 1.00 mole of nitrogen molecules
- c 1.00 mole of magnesium ions. Give your answer to three significant figures.

FINDING THE NUMBER OF PARTICLES IN DIFFERENT AMOUNTS OF SUBSTANCE

To work out the numbers of particles in other amounts of a substance, multiply the amount of substance, in mol, by the Avogadro constant.

Number of particles $= n \times N_A$



Calculate the number of ions in 3.00 moles of aluminium ions.

Number of particles (ions) $= 3.00\,\text{mol} \times 6.02 \times 10^{23}\,\text{mol}^{-1} = 1.806 \times 10^{24}$ ions

PRACTICE QUESTION

4 Calculate the number of particles in:
- a 2.00 moles of electrons
- b 1.50 moles of oxide ions
- c 0.2 moles of lithium ions.

STRETCH YOURSELF

Linking the Avogadro constant to the number of grams of a substance

The number of particles in a given mass of a substance can be calculated using a two-step calculation.

First work out the number of moles using $n = \frac{m}{M}$

where n is the number of moles, m is the mass of the substance and M is the molar mass of the substance.

Next find the number of particles using:

number of particles $= n \times N_A$

Calculate the number of particles in 18.0 g of carbon.

$$n = \frac{m}{M} = \frac{18.0\,\text{g}}{12.0\,\text{g mol}^{-1}} = 1.50 \text{ moles}$$

Number of particles $= n \times N_A$
$= 1.50 \times 6.02 \times 10^{23} = 9.03 \times 10^{23}$

PRACTICE QUESTION

5 Calculate the number of particles in 48.6 g of magnesium atoms.

Relative atomic mass

The relative atomic mass A_r

Most elements consist of several different isotopes. Each isotope has a different mass and each isotope occurs in a different percentage abundance. The relative atomic mass, A_r, is the weighted mean mass of an atom compared with one-twelfth of the mass of an atom of carbon-12.

RELATIVE ATOMIC MASSES (TWO ISOTOPES)

Chlorine

A sample of chlorine contains 75.0% of chlorine-35 and 25.0% of chlorine-37.
Calculate the relative atomic mass of chlorine.

$$A_r = \underbrace{\left(\frac{75.0}{100} \times 35\right)}_{\text{the contribution from the chlorine-35 atoms}} + \underbrace{\left(\frac{25.0}{100} \times 37\right)}_{\text{the contribution from the chlorine-37 atoms}} = 35.5$$

Lithium

A sample of lithium contains 7.0% of lithium-6 and 93.0% of lithium-7.
Calculate the relative atomic mass of lithium.

$$A_r = \underbrace{\left(\frac{7.0}{100} \times 6\right)}_{\text{the contribution from the lithium-6 atoms}} + \underbrace{\left(\frac{93.0}{100} \times 7\right)}_{\text{the contribution from the lithium-7 atoms}} = 6.93$$

PRACTICE QUESTIONS

1 A sample of boron contains 20.0% of boron-10 and 80.0% of boron-11. Calculate the relative atomic mass of boron.

2 A sample of gallium contains 60.0% of gallium-69 and 40.0% of gallium-71. Calculate the relative atomic mass of gallium.

3 A sample of potassium contains 93.0% of potassium-39 and 7.0% of potassium-41. Calculate the relative atomic mass of potassium.

REMEMBER: Questions sometimes ask for the answer to be given to a certain number of decimal places or a certain number of significant figures. Try circling these parts of the question to help you to remember to check that you have given your final answer in the correct form.

CALCULATING RELATIVE ATOMIC MASSES (THREE ISOTOPES)

A sample of magnesium was analysed and the percentage abundances and relative mass of the isotopes recorded. Calculate the relative atomic mass of magnesium. Give your answer to three significant figures.

Percentage abundance/%	Relative mass
79	24
10	25
11	26

$$A_r = \underbrace{\left(\frac{79}{100} \times 24\right)}_{\text{the contribution from the magnesium-24 atoms}} + \underbrace{\left(\frac{10}{100} \times 25\right)}_{\text{the contribution from the magnesium-25 atoms}} + \underbrace{\left(\frac{11}{100} \times 26\right)}_{\text{the contribution from the magnesium-26 atoms}} = 24.32$$

The calculator answer of 24.32 has four significant figures so the answer is rounded down to 24.3 to three significant figures.

PRACTICE QUESTIONS

4 A sample of titanium contains three isotopes. Their percentage abundances and relative masses are shown on the right.
Calculate the relative atomic mass of titanium.
Give your answer to three significant figures.

Percentage abundance/%	Relative mass
18	45
7	46
75	47

5 A sample of sulfur contains three isotopes. Their percentage abundances and relative masses are shown on the right.
Calculate the relative atomic mass of sulfur.
Give your answer to three significant figures.

Percentage abundance/%	Relative mass
95	32
1	33
4	34

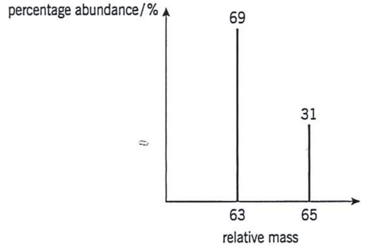
STRETCH YOURSELF

Extracting information from graphs

The percentage abundance of the different isotopes in a sample can be presented using a graph. The information from the graph can be used to deduce the percentage abundance and the relative mass of each isotope. This data can then be used to calculate the relative atomic mass of the element.

Example

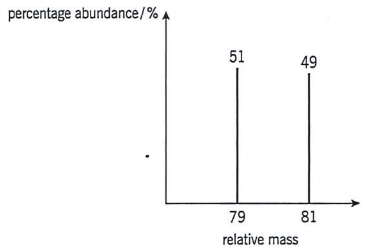
A sample of copper was analysed. Use the information in the graph to calculate the relative atomic mass of copper. Give your answer to three significant figures.



$$A_r = \left(\frac{69}{100} \times 63\right) + \left(\frac{31}{100} \times 65\right) = 63.62$$
$$= 63.6 \text{ (3 s.f.)}$$

PRACTICE QUESTION

6 The graph below shows the percentage abundance and relative mass of isotopes in a sample of bromine. Find the relative atomic mass of bromine. Give your answer to three significant figures.



Moles and gas volumes

Avogadro's law

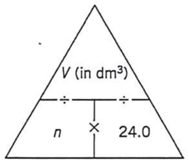
Avogadro's law states that equal volumes of gases at the same temperature and pressure contain an equal number of moles. At room temperature and pressure (RTP) one mole of any gas takes up a volume of 24 dm^3 or 24 000 cm^3. Room temperature and pressure are taken to be 25 °C and 1 atmosphere. This means that at room temperature and pressure one mole of helium atoms takes up a volume of 24 dm^3 or 24 000 cm^3
At room temperature and pressure one mole of oxygen molecules also takes up a volume of 24 dm^3 or 24 000 cm^3.

CALCULATIONS USING dm^3

Amount of substance from volume of gas at RTP

The amount of substance, in moles (n), and the volume of gas in dm^3 (V) are linked together by the equation:

$$n = \frac{V\text{ (in dm}^3\text{)}}{24.0}$$



Number of moles

Calculate the number of moles of gas in 12 dm^3 of nitrogen, N_2, at RTP:

$$n = \frac{V\text{ (in dm}^3\text{)}}{24.0}$$

$$n = \frac{12}{24.0} = 0.5\text{ mol}$$

PRACTICE QUESTION

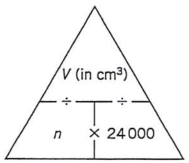
1 Calculate the number of moles of gas at RTP in:
- a 6 dm^3 of oxygen, O_2
- b 36 dm^3 of carbon dioxide, CO_2
- c 120 dm^3 of water vapour, H_2O
- d 72 dm^3 of carbon dioxide, CO_2
- e 12 dm^3 of water vapour, H_2O.

CALCULATIONS USING cm^3

Amount of substance from the volume of gas at RTP

The amount of substance, in moles (n), and the volume of gas in cm^3 (V) are linked together by the equation:

$$n = \frac{V\text{ (in cm}^3\text{)}}{24\,000}$$



Number of moles

Calculate the number of moles of gas in 600 cm^3 of oxygen, O_2 at RTP.

$$n = \frac{V\text{ (in cm}^3\text{)}}{24\,000}$$

$$n = \frac{600}{24\,000} = 0.025\text{ mol}$$

PRACTICE QUESTION

2 Calculate the number of moles of gas at RTP in:
- a 8000 cm^3 of carbon dioxide, CO_2
- b 72 000 cm^3 of carbon monoxide, CO
- c 1800 dm^3 of sulfur dioxide, SO_2
- d 12 000 cm^3 of carbon monoxide, CO
- e 3600 dm^3 of sulfur dioxide, SO_2.

CALCULATING THE VOLUME IN dm^3 THAT AN AMOUNT OF GAS OCCUPIES AT RTP

The volume of gas can be calculated using

$$V\text{ (in dm}^3\text{)} = n \times 24.0$$

Calculate the volume, in dm^3, of 2.75 moles of a gas at RTP.

$$V\text{ (in dm}^3\text{)} = 2.75 \times 24.0 = 66\text{ dm}^3$$

PRACTICE QUESTION

3 Calculate the volume, in dm^3, of these gases at RTP:
- a 0.10 moles of carbon dioxide, CO_2
- b 2.50 moles of sulfur dioxide, SO_2
- c 0.20 moles of water vapour, H_2O
- d 15.0 moles of sulfur dioxide, SO_2
- e 0.05 moles of water vapour, H_2O.

STRETCH YOURSELF

Linking the volume of gas to its mass

The mass of a gas in a given volume can be calculated using a two-step calculation.

First work out the number of moles using $n = \frac{V\text{ (in dm}^3\text{)}}{24.0}$

Next work out the mass of the gas using:

$$m = n \times M$$

Calculate the mass of 0.60 dm^3 of carbon dioxide, CO_2.

$$n = \frac{V\text{ (in dm}^3\text{)}}{24.0}$$

$$n = \frac{0.60\text{ (in dm}^3\text{)}}{24.0} = 0.025\text{ mol}$$

$$m = n \times M$$

$$m = 0.025\text{ mol} \times 44\text{ g mol}^{-1} = 1.1\text{ g}$$

PRACTICE QUESTION

4
- a Calculate the mass of 2.00 dm^3 of sulfur dioxide, SO_2.
- b Calculate the mass of 12.00 dm^3 of carbon dioxide, CO_2.
- c Calculate the mass of 120.00 dm^3 of oxygen, O_2.

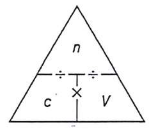
Concentration

Making a solution

A solution is made when a solute dissolves in a solvent. The concentration of a solution is a way of saying how much solute, in moles, is dissolved in 1 dm^3 or 1 litre of solution. Concentration is usually measured using units of mol dm^{-3}.

The concentration of the amount of substance dissolved in a given volume of a solution is given by the equation:

$$c = \frac{n}{V\text{ (in dm}^3)}$$



Where n is the amount of substance in moles, c is the concentration and V is the volume in dm^3.

CALCULATING CONCENTRATION

1) Calculate the concentration in mol dm^{-3} of a solution formed when 0.5 moles of a solute is dissolved in 2 dm^3 of solution.

$$c = \frac{n}{V\text{ (in dm}^3)}$$

$$c = \frac{0.5\text{ mol}}{2\text{ dm}^3} = 0.25\text{ mol dm}^{-3}$$

2) Calculate the concentration in mol dm^{-3} of a solution formed when 0.1 moles of a solute is dissolved in 500 cm^3 of solution.

$$500\text{ cm}^3 = 0.5\text{ dm}^3$$

$$c = \frac{n}{V\text{ (in dm}^3)}$$

$$c = \frac{0.1\text{ mol}}{0.5\text{ dm}^3} = 0.2\text{ mol dm}^{-3}$$

PRACTICE QUESTIONS

1 Calculate the concentration, in mol dm^{-3}, of a solution formed when 0.2 moles of a solute is dissolved in 50 cm^3 of solution.

2 Calculate the concentration, in mol dm^{-3}, of a solution formed when 0.25 moles of a solute is dissolved in 0.1 dm^3 of solution.

3 Calculate the concentration, in mol dm^{-3}, of a solution formed when 0.05 moles of a solute is dissolved in 2.0 dm^3 of solution.

Calculating the amount of substance from the concentration and volume of the solution

The equation can be rearranged to calculate the amount of substance, in moles, from a known volume and concentration of solution.

$$n = c \times V\text{ (in dm}^3)$$

WORKED EXAMPLE

Calculate the number of moles in a solution of sodium hydroxide, NaOH, in 25 cm^3 of aqueous solution of concentration 0.5 mol dm^{-3}.

$n = c \times V\text{ (in dm}^3)$

$n = 0.5\text{ mol dm}^{-3} \times 0.025\text{ dm}^3 = 0.0125\text{ mol}$

PRACTICE QUESTION

4 Calculate the number of moles of NaOH in an aqueous solution of:

a 36 cm^3 of 0.1 mol dm^{-3}

b 26 cm^3 of 0.5 mol dm^{-3}

c 50 cm^3 of 0.05 mol dm^{-3}.

Calculating the volume of a solution from a given amount of substance and concentration

The equation can be rearranged to calculate the volume of a solution from a known amount of substance, in moles, and the concentration of the solution.

$$V\text{ (in dm}^3) = \frac{n}{c}$$

WORKED EXAMPLE

Calculate the volume, in dm^3, of a solution of concentration 0.5 mol dm^{-3} that contains 0.05 moles of the solute.

$V\text{ (in dm}^3) = \frac{n}{c}$

$V = \frac{0.05\text{ mol}}{0.5\text{ mol dm}^{-3}} = 0.1\text{ dm}^3$

PRACTICE QUESTIONS

5 Calculate the volume, in dm^3, of a solution of concentration 0.10 mol dm^{-3} that contains 0.01 moles of the solute.

6 Calculate the volume, in dm^3, of a solution of concentration 0.05 mol dm^{-3} that contains 0.25 moles of the solute.

STRETCH YOURSELF

Linking the mass of a solute to the volume and concentration of a solution

The mass of a solute in a solution of a given volume and concentration can be calculated using a two-step calculation.

First work out the number of moles using $n = c \times V$ (in dm^3)

Next work out the mass of the solute using:

$$m = n \times M$$

where m is the mass of the substance, n is the amount of substance in moles and M is the molar mass.

Find the mass of sodium hydroxide, NaOH, required to prepare 250 cm^3 of an aqueous solution with a concentration of 0.10 mol dm^{-3}.

$n = c \times V$ (in dm^3)

$n = 0.1\text{ mol dm}^{-3} \times 0.25 = 0.025\text{ mol}$

$m = n \times M$

$m = 0.025\text{ mol} \times 40\text{ g mol}^{-1} = 1\text{ g}$

PRACTICE QUESTION

7 Find the mass of sodium hydroxide, NaOH, required to prepare 100 cm^3 of an aqueous solution with a concentration of 0.20 mol dm^{-3}.

Titrations

Acid–base titrations

In acid–base titrations a solution of an acid reacts with a solution of a base. The concentration of one of the solutions is known. A titration can be used to work out the concentration of the other solution. The amount of substance in moles (n), the volume of solution in cm^3 (V) and the concentration (c) are linked together by the equation:

$$n = \frac{V\text{ (in cm}^3)}{1000} \times c$$

FINDING THE CONCENTRATION OF A SOLUTION

A student carries out a titration to find the concentration of some hydrochloric acid. The student finds that 22.50 cm^3 of hydrochloric acid was required to neutralise 25.00 cm^3 of 0.10 mol dm^{-3} aqueous sodium hydroxide solution.

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

a) Calculate the number of moles of sodium hydroxide used.

$$n = \frac{V\text{ (in cm}^3)}{1000} \times c = \frac{25.00}{1000} \times 0.1 = 0.0025\text{ moles}$$

b) Calculate the number of moles of hydrochloric acid used.
From the stoichiometric equation, the number of moles of hydrochloric acid used = 0.0025 moles.

c) Calculate the concentration of the hydrochloric acid. Give your answer to two decimal places.

$$c = \frac{n \times 1000}{V} = \frac{0.0025 \times 1000}{22.50} = 0.11\text{ mol dm}^{-3}$$

PRACTICE QUESTIONS

1 A student carries out a titration to find the concentration of some nitric acid. The student finds that 50 cm^3 of 0.125 mol dm^{-3} aqueous sodium hydroxide solution was neutralised by 22.50 cm^3 of the nitric acid. Give your answers to two decimal places.

$$HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

a Calculate the number of moles of sodium hydroxide used.
b Calculate the number of moles of nitric acid used.
c Calculate the concentration of the nitric acid.

2 A chemist carries out a titration to find the concentration of some hydrochloric acid. 25.0 cm^3 of a standard solution of 0.20 mol dm^{-3} sodium hydroxide was placed into conical flask. The chemist found that 22.0 cm^3 of hydrochloric acid was required for neutralisation.

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

a Calculate the number of moles of sodium hydroxide used.
b Calculate the number of moles of hydrochloric acid used.
c Calculate the concentration of the hydrochloric acid.

WORKED EXAMPLE

A student carries out a titration to find the concentration of some sodium hydroxide solution. The student finds that 25.00 cm^3 of aqueous sodium hydroxide solution was neutralised by 28.00 cm^3 of 0.08 mol dm^{-3} sulfuric acid.

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

a) Calculate the number of moles of sulfuric acid used.

$$n = \frac{V\text{ (in cm}^3)}{1000} \times c = \frac{28.00}{1000} \times 0.08 = 0.00224\text{ moles}$$

b) Calculate the number of moles of sodium hydroxide used.
From the stoichiometric equation, the number of moles of sodium hydroxide used = 0.00448 moles.

c) Calculate the concentration of the sodium hydroxide. Give your answer to two decimal places.

$$c = \frac{n \times 1000}{V} = \frac{0.00448 \times 1000}{25.00} = 0.18\text{ mol dm}^{-3}$$

PRACTICE QUESTIONS

3 A student carries out a titration to find the concentration of some sulfuric acid. The student finds that 28.00 cm^3 of sulfuric acid was required to neutralise 25.00 cm^3 of 0.02 mol dm^{-3} aqueous potassium hydroxide solution.

$$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$$

a Calculate the number of moles of potassium hydroxide used.
b Calculate the number of moles of sulfuric acid used.
c Calculate the concentration of the sulfuric acid. Give your answer to four decimal places.

4 A student carries out a titration to find the concentration of some nitric acid. The student finds that 28.00 cm^3 of nitric acid was required to neutralise 50.00 cm^3 of 2.00 mol dm^{-3} aqueous potassium hydroxide solution.

$$HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$$

a Calculate the number of moles of potassium hydroxide used.
b Calculate the number of moles of nitric acid used.
c Calculate the concentration of the nitric acid. Give your answer to four decimal places.

STRETCH YOURSELF

Finding the volume of solution used

In titrations, chemists must find out exactly how much solution is required to neutralise a measured volume of a second solution. This means that they must repeat the titration until they are confident that they have found the correct volume. Chemists know they have found that volume when they repeat the titration and get two very similar (concordant) results. The mean average of the concordant results is used to find the volume that is then used in calculations.

Use the results table below to work out the mean average volume of sulfuric acid used.

Experiment	Volume of sulfuric acid used/cm^3
1	22.80
2	22.50
3	22.60

The mean average should only include the concordant results = $\frac{(22.50 + 22.60)}{2}$ = 22.55 cm^3

PRACTICE QUESTION

5 Use the results table below to work out the mean average volume of sodium hydroxide used.

Experiment	Volume of sodium hydroxide used/cm^3
1	26.40
2	26.90
3	26.30

Mole calculations 1

Reacting masses and gas volumes

The balanced equation for a reaction shows how many moles of each reactant and product are involved in a chemical reaction.

Sodium reacts with chlorine to form sodium chloride:

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

The balanced equation shows that two moles of sodium react with one mole of chlorine molecules to form two moles of sodium chloride.

The molar reacting quantities can be calculated using the balanced equation.

If the amount, in moles, of one of the reactants or products is known, the number of moles of the other reactants and products can be calculated.

The number of moles (n), the mass of the substance (m) and the molar mass (M) are linked together using $n = \frac{m}{M}$

The amount of substance in moles (n) and the volume of gas in dm^3 (V) are linked together by the equation:

$$n = \frac{V\text{ (in dm}^3)}{24.0}$$

USING BALANCED EQUATIONS

Magnesium chloride

A chemist reacted 0.243 g of magnesium with chlorine to produce magnesium chloride.

$$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$$

Molar mass of $MgCl_2$ is 95.3 g mol^{-1}.

a) Calculate the amount, in mol, of magnesium that reacts.

$n = \frac{m}{M} = \frac{0.243}{24.3} = 0.01$ mol

b) Calculate the amount, in mol, of magnesium chloride that was made.

From the balanced equation, the number of moles of magnesium = number of moles of magnesium chloride = 0.01 mol

c) Calculate the mass, in grams, of magnesium chloride made. Give your answer to three decimal places.

$m = n \times M = 0.01 \times 95.3 = 0.953$ g

PRACTICE QUESTIONS

1 In a reaction, 0.486 g of magnesium was added to oxygen to produce magnesium oxide.

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

- a Calculate the amount, in moles, of magnesium that reacted.
- b Calculate the amount, in moles, of magnesium oxide made.
- c Calculate the mass, in grams, of magnesium oxide made.

2 In a reaction, 0.115 g of sodium was added to chlorine to produce sodium chloride.

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

- a Calculate the amount, in moles, of sodium that reacted.
- b Calculate the amount, in moles, of sodium chloride made.
- c Calculate the mass, in grams, of sodium chloride made.

3 In a reaction, 0.1955 g of potassium was added to chlorine to produce potassium chloride.

$$2K(s) + Cl_2(g) \rightarrow 2KCl(s)$$

- a Calculate the amount, in moles, of potassium that reacted.
- b Calculate the amount, in moles, of potassium chloride made.
- c Calculate the mass, in grams, of potassium chloride made.

WORKED EXAMPLE

A student heated 2.50 g of calcium carbonate, which decomposed as shown in the equation.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

a) Calculate the amount, in mol, of calcium carbonate that decomposes.

$n = \frac{m}{M} = \frac{2.50}{100.1} = 0.025$ mol

b) Calculate the amount, in mol, of carbon dioxide that forms.

From the balanced equation, the number of moles of calcium carbonate = number of moles of carbon dioxide = 0.025 mol

c) Calculate the volume, in dm^3, of carbon dioxide made.

$V\text{ (in dm}^3) = n \times 24.0 = 0.025 \times 24 = 0.60\text{ dm}^3$

PRACTICE QUESTIONS

4 A student heated 4.25 g of sodium nitrate. The equation for the decomposition of sodium nitrate is given below.

$$2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$$

- a Calculate the amount, in moles, of sodium nitrate that reacted.
- b Calculate the amount, in moles, of oxygen made.
- c Calculate the volume, in dm^3, of oxygen made at RTP.

5 A 0.2764 g sample of potassium carbonate decomposes on heating to form potassium oxide and carbon dioxide.

$$K_2CO_3(s) \rightarrow K_2O(s) + CO_2(g)$$

- a Calculate the amount, in moles, of potassium carbonate that reacted.
- b Calculate the amount, in moles, of carbon dioxide made.
- c Calculate the volume, in dm^3, of carbon dioxide made at RTP.

6 A chemist heated 2.022 g of potassium nitrate. The equation for the decomposition of potassium nitrate is:

$$2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$$

- a Calculate the amount, in moles, of potassium nitrate that reacted.
- b Calculate the amount, in moles, of oxygen made.
- c Calculate the volume, in dm^3, of oxygen made at RTP.

7 0.500 kg of magnesium carbonate decomposes on heating to form magnesium oxide and carbon dioxide. Give your answers to three decimal places.

$$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$$

- a Calculate the amount, in mol, of magnesium carbonate used.
- b Calculate the amount, in mol, of carbon dioxide produced.
- c Calculate the volume, in dm^3, of carbon dioxide produced at RTP.

Mole calculations 2

Reacting masses and volumes

The balanced symbol equation for a reaction can be used to work out the quantities of reactant and products involved in a reaction.

The number of moles (n), the mass of the substance (m) and the molar mass (M) are linked together using $n = \frac{m}{M}$

The amount of substance in moles (n), the volume of solution in cm^3 (V) and the concentration (c) are linked together by the equation:

$$n = \frac{V\ (\text{in } cm^3)}{1000} \times c$$

In these calculations, first work out the amount in moles of one of the substances involved in the reaction. Then use the balanced symbol equation to work out the amount in moles of the desired substances. Then use this information to solve the last part of the question.

CALCULATING THE VOLUME OF SOLUTIONS

Calcium carbonate

Calcium carbonate reacts with 0.25 mol dm^{-3} hydrochloric acid to make calcium chloride. Water and carbon dioxide are also produced. 1.25 g of calcium carbonate is used in the reaction.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

a) Calculate the amount, in mol, of calcium carbonate that reacts.

$$n = \frac{m}{M} = \frac{1.25}{100.1} = 0.0125\ \text{mol}$$

b) Calculate the amount, in mol, of hydrochloric acid that reacts.
From the balanced equation the number of moles of hydrochloric acid = number of moles of calcium carbonate × 2 = 0.0250 mol.

c) Calculate the volume, in cm^3, of hydrochloric acid used. Give your answer to three significant figures.

$$V = \frac{n \times 1000}{c} = \frac{0.0250 \times 1000}{0.25} = 100\ cm^3$$

PRACTICE QUESTIONS

1 A student reacted 2.47 g of copper carbonate with 0.1 mol dm^{-3} sulfuric acid.

$$CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$$

- a Calculate the amount, in moles, of copper carbonate that reacted.
- b Calculate the amount, in moles, of sulfuric acid used in the reaction.
- c Calculate the volume, in cm^3, of sulfuric acid used.

2 In a reaction 1.68 g of magnesium carbonate was reacted with 2.0 mol dm^{-3} hydrochloric acid.

$$MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l) + CO_2(g)$$

- a Calculate the amount, in moles, of magnesium carbonate that reacted.
- b Calculate the amount, in moles, of hydrochloric acid used in the reaction.
- c Calculate the volume, in cm^3, of hydrochloric acid used.

3 In a reaction 1.254 g of zinc carbonate was reacted with 1.0 mol dm^{-3} hydrochloric acid.

$$ZnCO_3(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l) + CO_2(g)$$

- a Calculate the amount, in moles, of zinc carbonate that reacted.
- b Calculate the amount, in moles, of hydrochloric acid used in the reaction.
- c Calculate the volume, in cm^3, of hydrochloric acid used.

CALCULATING THE CONCENTRATION OF SOLUTIONS

Reacting masses and concentrations

A student reacted 0.403 g of magnesium oxide with 25.0 cm^3 of nitric acid to form magnesium nitrate and water.

$$MgO(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2O(l)$$

a) Calculate the amount, in mol, of magnesium oxide that reacted.

$$n = \frac{m}{M} = \frac{0.403}{40.3} = 0.01\ \text{mol}$$

b) Calculate the amount, in mol, of nitric acid used.
From the balanced equation, the amount of nitric acid = 0.02 mol

c) Calculate the concentration, in mol dm^{-3}, of the nitric acid used.

$$c = \frac{n \times 1000}{v} = \frac{0.02 \times 1000}{25.0} = 0.08\ \text{mol dm}^{-3}$$

PRACTICE QUESTIONS

4 A student reacted 4.00 g of calcium carbonate with 50 cm^3 of hydrochloric acid. Calcium chloride, water and carbon dioxide were produced in the reaction.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

- a Calculate the amount, in mol, of calcium carbonate that reacted.
- b Calculate the amount, in mol, of hydrochloric acid that reacted.
- c Calculate the concentration of hydrochloric acid used.

5 A student reacted 1.686 g of magnesium carbonate with 50 cm^3 of hydrochloric acid. Magnesium chloride, water and carbon dioxide were produced in the reaction.

$$MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l) + CO_2(g)$$

- a Calculate the amount, in mol, of magnesium carbonate that reacted.
- b Calculate the amount, in mol, of hydrochloric acid that reacted.
- c Calculate the concentration of the hydrochloric acid used.

6 A student reacted 5.00 g of calcium carbonate with 50 cm^3 of sulfuric acid. Calcium sulfate, water and carbon dioxide were produced in the reaction.

$$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2O(l) + CO_2(g)$$

- a Calculate the amount, in mol, of calcium carbonate that reacted.
- b Calculate the amount, in mol, of sulfuric acid that reacted.
- c Calculate the concentration of the sulfuric acid used.

7 A chemist prepared an aqueous solution of 0.10 mol dm^{-3} potassium hydroxide. What is the concentration of the solution in g dm^{-3}?

8 A chemist prepared an aqueous solution of 0.20 mol dm^{-3} sodium hydroxide. What is the concentration of the solution in g dm^{-3}?

9 A chemist prepared an aqueous solution of 0.05 mol dm^{-3} potassium hydroxide. What is the concentration of the solution in g dm^{-3}? Give your answer to three significant figures.

Water of crystallisation

Hydrated salts

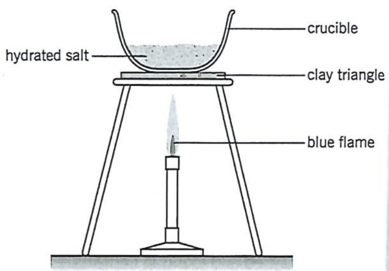
Hydrated salts are crystalline compounds that contain water molecules, for example, hydrated copper(II) sulfate. Hydrated copper(II) sulfate has the chemical formula $CuSO_4.5H_2O$. The last part of the name refers to the water of crystallisation. This is the water molecules that are found within the crystalline structure of the hydrated salt. Copper(II) sulfate also exists in an anhydrous form. The anhydrous form of the compound does not contain water molecules.

The molar mass of anhydrous copper(II) sulfate = 159.6 g mol^{-1}

The molar mass of hydrated copper(II) sulfate = $159.6 + (5 \times 18.0) = 249.6$ g mol^{-1}

Determining the chemical formula of a hydrated salt

When a sample of a hydrated salt is heated strongly, the water of crystallisation can be driven off and evaporated. As this happens the mass of the sample decreases. The sample can be repeatedly heated and then its mass measured until, eventually, the mass remains constant. When this happens all of the water of crystallisation has been removed from the hydrated salt.



This method is only suitable for salts where the anhydrous form of the compound does not decompose on heating.

The table below shows a sample set of readings from an experiment to find the formula of hydrated sodium carbonate.

Mass of sample before heating/g	5.720
Mass of sample after heating/g	2.120
Mass of water lost during heating/g	3.600

Notice how all the readings are given to the same number of decimal places.

The mass of the sample before heating is the mass of the hydrated salt.

The mass of the sample after heating is the mass of the anhydrous salt.

FINDING THE FORMULA OF A HYDRATED SALT

Use the information in the table above to find the formula of the hydrated salt.

Number of moles (n), the mass of the substance (m) and the molar mass (M) are linked together using $n = \frac{m}{M}$

The amount, in mol, of anhydrous sodium carbonate $= \frac{2.120}{106} = 0.02$ mol

The amount, in mol, of water $= \frac{3.600}{18.0} = 0.20$ mol

The molar ratio of $Na_2CO_3 : H_2O$ is:

	Na_2CO_3	H_2O
molar ratio	0.02	0.20
next divide by the smallest number	$\frac{0.02}{0.02}$	$\frac{0.20}{0.02}$
	= 1	= 10

The formula of hydrated sodium carbonate is $Na_2CO_3.10H_2O$.

PRACTICE QUESTIONS

1 A sample of hydrated magnesium sulfate was strongly heated. The mass of the magnesium sulfate before and after heating were recorded in the table below.

Mass of sample before heating/g	1.232
Mass of sample after heating/g	0.602
Mass of water lost during heating/g	

a Copy and complete the table to show the mass of water lost during heating.
b Calculate the amount, in mol, of anhydrous magnesium sulfate produced.
c Calculate the amount, in mol, of water.
d Calculate the formula of hydrated magnesium sulfate.

2 A sample of hydrated zinc sulfate was strongly heated. The mass of the zinc sulfate before and after heating were recorded in the table below.

Mass of sample before heating/g	2.875
Mass of sample after heating/g	1.615
Mass of water lost during heating/g	

a Copy and complete the table to show the mass of water lost during heating.
b Calculate the amount, in mol, of anhydrous zinc sulfate produced.
c Calculate the amount, in mol, of water.
d Calculate the formula of hydrated zinc sulfate.

3 A sample of hydrated magnesium chloride was strongly heated. The mass of the magnesium chloride before and after heating were recorded in the table below.

Mass of sample before heating/g	3.706
Mass of sample after heating/g	1.906
Mass of water lost during heating/g	

a Copy and complete the table to show the mass of water lost during heating.
b Calculate the amount, in mol, of anhydrous magnesium sulfate produced.
c Calculate the amount, in mol, of water.
d Calculate the formula of hydrated magnesium chloride.

STRETCH YOURSELF

Empirical formula and dot formula

As well as being found in waters of crystallisation, oxygen is also found in sulfate, SO_4^{2-}, carbonate, CO_3^{2-} and nitrate, NO_3^- ions.
Chemists use the hydrogen atoms to work out the number of waters of crystallisation.

Worked example

A compound of a hydrated salt has an empirical formula of $CuSH_{10}O_9$.
Write the dot formula for the compound.
If there are 10 hydrogen atoms, there are 5 waters of crystallisation ($5H_2O$).
This leaves $CuSO_4$ so the dot formula is $CuSO_4.5H_2O$.

PRACTICE QUESTION

4 Write the dot formula for a hydrated salt with the empirical formula:
a $CoCl_2H_{12}O_6$
b $CaN_2H_8O_{10}$
c $Na_2SH_{20}O_{14}$.