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 $12.0\,g \times 2 = 24.0\,g$
3.00 moles of sulfur has a mass of $32.1\,g \times 3 = 96.3\,g$
2.50 moles of copper has a mass of $63.5\,g \times 2.5 = 158.75\,g$

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.

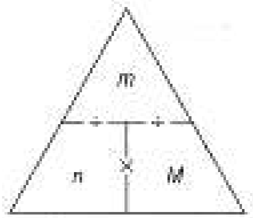
$(2 \times 1.0) + (1 \times 16.0) = 18.0\,g\,mol^{-1}$

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 $6.02 \times 10^{23}\ mol^{-1}$ 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

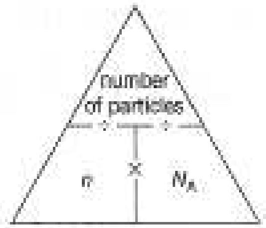
3 Calculate the number of particles in:

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) $= 3.00\ mol \times 6.02 \times 10^{23}\ mol^{-1} = 1.806 \times 10^{24}$ 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.

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

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

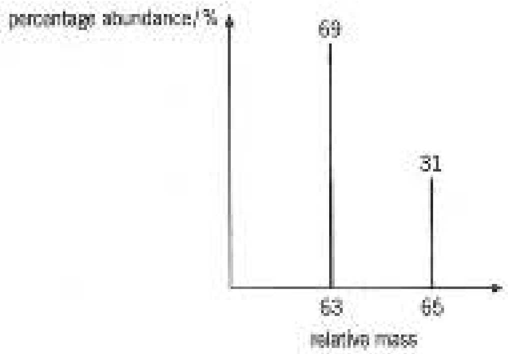
STRETCH YOURSELF

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

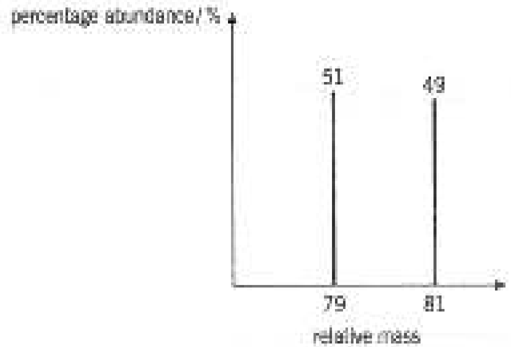
A sample of copper was analysed. Use the information in the graph to calculate the relative atomic mass of copper. Give your answer to three significant figures.



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

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

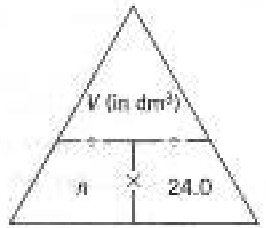
Avogadro's law states that equal volumes of gases at the same temperature and pressure contain an equal number of moles. At room temperature and pressure (RTP) one mole of any gas takes up a volume of 24 dm^3 or 24 000 cm^3. Room temperature and pressure are taken to be 25 °C and 1 atmosphere. This means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

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



Number of moles

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

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

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

PRACTICE QUESTION

1 Calculate the number of moles of gas at RTP in:

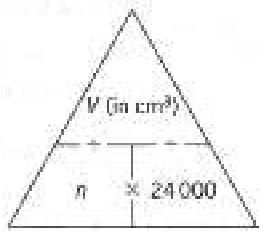
a 6 dm^3 of oxygen, O_2
b 36 dm^3 of carbon dioxide, CO_2
c 120 dm^3 of water vapour, H_2O
d 72 dm^3 of carbon dioxide, CO_2
e 12 dm^3 of water vapour, H_2O.

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:

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



Number of moles

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

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

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

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

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

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

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

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

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.

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

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

$$m = n \times M$$

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

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

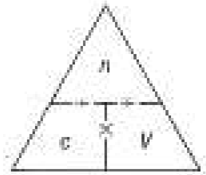
Making a solution

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:

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

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.

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

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

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.

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

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

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

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.

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

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}.

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

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

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.

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

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.

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

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

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}.

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

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

$$m = n \times M$$

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

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:

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

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.

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

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.

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

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 $= \frac{(22.50 + 22.60)}{2} = 22.55\text{ cm}^3$

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 dm³ (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.

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

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.

$$m = n \times M = 0.01 \times 95.3 = 0.953 \text{ g}$$

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.

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

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 dm³, of carbon dioxide made.
V (in dm³) = $n \times 24.0 = 0.025 \times 24 = 0.60$ dm³

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 dm³, 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 dm³, 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 dm³, 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 dm³, 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.

Empirical formula 1

The empirical formula

The empirical formula of a compound is the simplest whole number ratio of atoms of each of the elements present in the compound.

It can be calculated from the mass or the percentage composition by mass of each of the elements present in the compound. The amount of substance in mol (n), the mass of the substance (m) and the molar mass (M) are linked together using $n = \frac{m}{M}$

EMPIRICAL FORMULA FROM GIVEN MASSES

A sample of titanium oxide was analysed and found to contain 0.958 g of titanium and 0.640 g of oxygen. Calculate the empirical formula of the compound. (A_r: Ti, 47.9; O, 16.0)

	Ti	O
mass of each element	0.958	0.640
molar ratio of elements	$\frac{0.958}{47.9} = 0.02$	$\frac{0.640}{16} = 0.04$
next divide by the smallest number	$\frac{0.02}{0.02} = 1$	$\frac{0.04}{0.02} = 2$

The empirical formula is TiO_2.

PRACTICE QUESTIONS

1 A sample of phosphorus chloride was analysed and found to contain 0.62 g of phosphorus and 3.55 g of chlorine. Calculate the empirical formula of the compound.
(A_r: P, 31.0; Cl, 35.5)

2 A sample of a compound was analysed and found to contain 0.070 g of nitrogen and 0.015 g of hydrogen. Calculate the empirical formula of the compound.
(A_r: N, 14.0; H, 1.0)

ANALYSING A SALT

A sample of a salt was analysed and found to contain 0.115 g of sodium, 0.070 g of nitrogen and 0.240 g of oxygen. Calculate the empirical formula of the compound.
(A_r: Na, 23.0; N, 14.0; O, 16.0)

	Na	N	O
mass of each element	0.115	0.070	0.240
molar ratio of elements	$\frac{0.115}{23.0} = 0.005$	$\frac{0.070}{14.0} = 0.005$	$\frac{0.240}{16} = 0.015$
next divide by the smallest number	$\frac{0.005}{0.005} = 1$	$\frac{0.005}{0.005} = 1$	$\frac{0.015}{0.005} = 3$

The empirical formula is $NaNO_3$.

PRACTICE QUESTIONS

3 A sample of a metal carbonate was analysed and found to contain 0.162 g of magnesium, 0.080 g of carbon and 0.320 g of oxygen. Calculate the empirical formula of the compound.
(A_r: Mg, 24.3; C, 12.0; O, 16.0)

4 A sample of a compound was analysed and found to contain 0.254 g of copper, 0.128 g of sulfur and 0.256 g of oxygen. Calculate the empirical formula of the compound.
(A_r: Cu, 63.5; S, 32.1; O, 16.0)

Percentage composition by mass and empirical formulae

Sometimes the analysis of a compound is given as a percentage composition by mass. The empirical formula is calculated using the same method. Simply use the percentage mass by composition instead of the mass in the calculations.

ANALYSING A COMPOUND

A chemist analysed a sample of a compound to find the percentage composition by mass of each element. They found that it contained 20.2% magnesium, 26.7% sulfur and 53.1% oxygen.

Calculate the empirical formula of the compound.
(A_r: Mg, 24.3; S, 32.1; O, 16.0)

	Mg	S	O
percentage composition by mass	20.2	26.7	53.1
molar ratio of elements	$\frac{20.2}{24.3} = 0.831$	$\frac{26.7}{32.1} = 0.832$	$\frac{53.1}{16.0} = 3.319$
next divide by the smallest number	$\frac{0.831}{0.831} = 1$	$\frac{0.832}{0.831} \approx 1$	$\frac{3.319}{0.831} \approx 4$

The empirical formula is $MgSO_4$.

STRETCH YOURSELF

Information in names

The name of a compound reveals the elements it is made from. For example, a metal oxide will contain metal and oxygen only. If the mass of the sample is known and the mass of the metal is known:

mass of oxygen = mass of compound − mass of metal

This mass can then be used in calculations.

A 0.286 g sample of a metal oxide was analysed and found to contain 0.254 g of copper. Calculate the empirical formula of the compound. (A_r: Cu, 63.5; O, 16.0).

The mass of oxygen in this sample = 0.286 − 0.254 = 0.032 g

	Cu	O
mass of each element	0.254	0.032
molar ratio of elements	$\frac{0.254}{63.5} = 0.004$	$\frac{0.032}{16.0} = 0.002$
next divide by the smallest number	$\frac{0.004}{0.002} = 2$	$\frac{0.002}{0.002} = 1$

The empirical formula is Cu_2O.

PRACTICE QUESTIONS

5 A 0.798 g sample of a metal oxide was analysed and found to contain 0.558 g of iron. Calculate the empirical formula of the compound. (A_r: Fe, 55.8; O, 16.0)

6 A 1.268 g sample of a metal chloride was analysed and found to contain 0.558 g of iron. Calculate the empirical formula of the compound. (A_r: Fe, 55.8; Cl, 35.5)

Empirical formula 2

Empirical formula and molecular formula

The empirical formula of a compound is the simplest whole number ratio of atoms of each of the elements present in the compound.

The molecular formula is the actual number of atoms of each element in one molecule.

The molecular formula can be calculated from the empirical formula and the molar mass. If the empirical formula in grams is equal to the molar mass, then the molecular formula and empirical formula are the same.

Empirical formulae can also be calculated by analysing the products of the combustion of compounds. Complete combustion of carbon produces carbon dioxide, while combustion of hydrogen produces water vapour.

WORKED EXAMPLES

1) A sample was analysed and found to contain 0.048 g of carbon and 0.016 g of hydrogen.

a) Calculate the empirical formula of the compound. (A_r: H, 1.0; C, 12.0)

	C	H
mass of each element	0.048	0.016
molar ratio of elements	$\frac{0.048}{12.0}$ = 0.004	$\frac{0.016}{1.0}$ = 0.016
next divide by the smallest number	$\frac{0.004}{0.004}$ = 1	$\frac{0.016}{0.004}$ = 4

The empirical formula is CH_4.

b) The compound has a molar mass of 16.0 g mol^{-1}. Find the molecular formula of the compound.
The empirical formula mass = molecular mass
So the molecular formula is the same as the empirical formula, CH_4.

2) A second sample was analysed and found to contain 0.240 g of carbon and 0.040 g of hydrogen.

a) Calculate the empirical formula of the compound. (A_r: H, 1.0; C, 12.0)

	C	H
mass of each element	0.24	0.040
molar ratio of elements	$\frac{0.024}{12.0}$ = 0.020	$\frac{0.040}{1.0}$ = 0.040
next divide by the smallest number	$\frac{0.020}{0.020}$ = 1	$\frac{0.040}{0.020}$ = 2

The empirical formula is CH_2

b) The compound has a molar mass of 28.0 g mol^{-1}. Find the molecular formula of the compound.
The empirical formula mass × 2 = molecular mass
So, the molecular formula is C_2H_4.

PRACTICE QUESTIONS

1 A sample was analysed and found to contain 0.300 g of carbon and 0.050 g of hydrogen.
  a Calculate the empirical formula of the compound. (A_r: H, 1.0; C, 12.0)
  b The compound has a molar mass of 84.0 g mol^{-1}. Find the molecular formula of the compound.

2 A sample was analysed and found to contain 0.28 g of nitrogen and 0.04 g of hydrogen.
  a Calculate the empirical formula of the compound. (A_r: H, 1.0; N, 14.0)
  b The compound has a molar mass of 32.0 g mol^{-1}. Find the molecular formula of the compound.

3 A sample was analysed and found to contain 0.18 g of carbon and 0.03 g of hydrogen.
  a Calculate the empirical formula of the compound. (A_r: H, 1.0; C, 12.0)
  b The compound has a molar mass of 98.0 g mol^{-1}. Find the molecular formula of the compound.

4 A sample was analysed and found to contain 0.16 g of oxygen and 0.01 g of hydrogen.
  a Calculate the empirical formula of the compound. (A_r: H, 1.0; O, 16.0)
  b The compound has a molar mass of 34.0 g mol^{-1}. Find the molecular formula of the compound.

STRETCH YOURSELF

Combustion analysis

An organic compound A was analysed and found to contain hydrogen and carbon only. A sample of compound A was burnt in excess oxygen and the products of combustion were analysed and found to contain 0.220 g of carbon dioxide and 0.180 g of water vapour. What is the empirical formula of compound A?

The molar mass of carbon dioxide = 44.0 g mol^{-1}

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 carbon dioxide, in mol = $\frac{0.220\,g}{44.0\,g\,mol^{-1}}$ = 0.005 mol

The molar mass of water = 18.0 g mol^{-1}

The amount of water, in mol = $\frac{0.180\,g}{18.0\,g\,mol^{-1}}$ = 0.010 mol

Each molecule of carbon dioxide, CO_2, contains one atom of carbon so compound A must contain 0.005 moles of carbon.

Each molecule of water vapour, H_2O, contains two atom of hydrogen so compound A must contain 0.010 × 2 = 0.020 moles of hydrogen.

	C	H
molar ratio of elements	0.005	0.020
next divide by the smallest number	$\frac{0.005}{0.005}$ = 1	$\frac{0.020}{0.005}$ = 4

The empirical formula of compound A is CH_4.

PRACTICE QUESTION

5 A compound X was analysed and found to contain hydrogen and carbon only. A sample of compound X was burnt in excess oxygen and the products of combustion were analysed and found to contain 0.176 g of carbon dioxide and 0.054 g of water vapour. What is the empirical formula of compound X?

Water of crystallisation

Hydrated salts

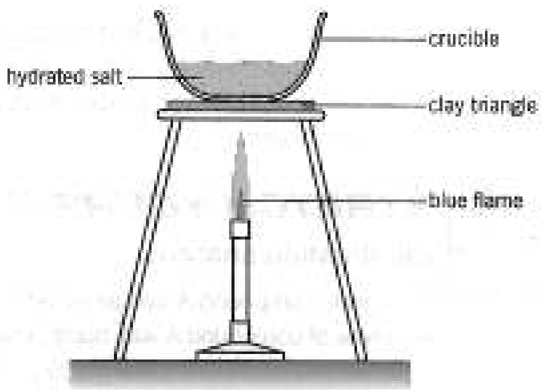
Hydrated salts are crystalline compounds that contain water molecules, for example, hydrated copper(II) sulfate. Hydrated copper(II) sulfate has the chemical formula $CuSO_4.5H_2O$. The last part of the name refers to the water of crystallisation. This is the water molecules that are found within the 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 $Na_2SH_{20}O_{14}$.

Group 7

Good oxidising agents

Group 7 elements are called halogens. They are found in the p-block of the periodic table. Atoms of Group 7 elements have seven electrons in their outer shell. During chemical reactions atoms of Group 7 elements typically gain one electron to form halide ions with an oxidation number of −1. This means that Group 7 elements are good oxidising agents.

Reactions with metals

Group 7 elements react with metals to form metal halides.

WORKED EXAMPLE

Potassium and chlorine

Potassium reacts with chlorine to form potassium chloride.

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

a) Write separate half equations for this reaction.

$2K \rightarrow 2K^+ + 2e^-$

$Cl_2 + 2e^- \rightarrow 2Cl^-$

b) Use oxidation numbers to identify which species are oxidised and reduced in this reaction.

The oxidation number of potassium increases from 0 to +1 so the potassium is oxidised.

The oxidation number of chlorine decreases from 0 to −1 so the chlorine is reduced.

c) Identify the oxidising agent and the reducing agent in this reaction.

The potassium reduces the chlorine, so the potassium is the reducing agent. The chlorine oxidises the potassium, so the chlorine is the oxidising agent.

PRACTICE QUESTIONS

1 A chemist reacts sodium with fluorine to form sodium fluoride.

$2Na(s) + F_2(g) \rightarrow 2NaF(s)$

a Use oxidation numbers to identify which species have been oxidised and reduced in this reaction.

b Identify the oxidising agent and the reducing agent in this reaction.

2 A chemist reacts potassium with fluorine to form potassium fluoride.

$2K(s) + F_2(g) \rightarrow 2KF(s)$

a Use oxidation numbers to identify which species have been oxidised and reduced in this reaction.

b Identify the oxidising agent and the reducing agent in this reaction.

Redox reactions

Group 7 elements become less reactive going down the group. A more reactive halogen will oxidise a less reactive halogen and displace it from a solution of its salt.

WORKED EXAMPLE

Chlorine reacts with an aqueous solution of potassium bromide to form potassium chloride and bromine.

$Cl_2(aq) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(aq)$

a) Write separate half equations for this reaction.

$2Br^- \rightarrow Br_2 + 2e^-$

$Cl_2 + 2e^- \rightarrow 2Cl^-$

b) Use oxidation numbers to identify which species are oxidised and reduced in this reaction.

The oxidation number of bromine increases from −1 to 0 so the bromine is oxidised.

The oxidation number of chlorine decreases from 0 to −1 so the chlorine is reduced.

c) Identify the oxidising agent and reducing agent in this reaction.

The bromine reduces the chlorine, so the potassium bromide is the reducing agent. The chlorine oxidises the bromine, so the chlorine is the oxidising agent.

> **REMEMBER:** The reducing agent is the name of the whole reagent, i.e. potassium bromide.

PRACTICE QUESTIONS

3 Bromine reacts with an aqueous solution of potassium iodide to form potassium bromide and iodine.

$Br_2(aq) + 2KI(aq) \rightarrow 2KBr(aq) + I_2(aq)$

a Use oxidation numbers to identify which species are oxidised and reduced in the reaction.

b Identify the oxidising agent and the reducing agent.

4 Chlorine reacts with an aqueous solution of potassium iodide to form potassium chloride and iodine.

$Cl_2(aq) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$

a Use oxidation numbers to identify which species are oxidised and reduced in the reaction.

b Identify the oxidising agent and the reducing agent.

STRETCH YOURSELF

Disproportionation reactions

In disproportionation reactions a species is simultaneously oxidised and reduced.

Example

Chlorine reacts with water to form chloric(I) acid and hydrochloric acid.

$Cl_2(aq) + H_2O(l) \rightarrow HClO(aq) + HCl(aq)$

Use oxidation numbers to prove that this is a disproportionation reaction.

The oxidation number of chlorine increases from 0 to +1 (in chloric(I) acid), while the oxidation number of chlorine decreases from 0 to −1 (in hydrochloric acid). So chlorine is simultaneously oxidised and reduced.

PRACTICE QUESTION

5 Copper(I) oxide reacts to form copper(II) oxide and copper.

$Cu_2O \rightarrow CuO + Cu$

Use oxidation numbers to show that this is a disproportionation reaction.

Enthalpy changes 1

Enthalpy changes, ΔH

Enthalpy, H, is the heat content that is stored in a chemical system. Enthalpy change, ΔH, can be calculated from experimental data.

In exothermic reactions, heat is lost by the chemical system and gained by the surroundings so a temperature rise is observed. The enthalpy change for an exothermic reaction has a negative sign.

In endothermic reactions, heat is gained by the chemical system from the surroundings so a temperature fall is observed. The enthalpy change for an endothermic reaction has a positive sign.

Heat exchanged

The heat gained or lost by the surroundings can be calculated using the relationship:

$$Q = mc\Delta T$$

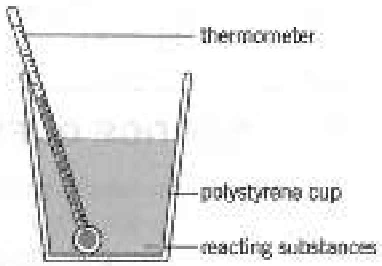
where Q is the heat exchange measured in J,

m is the mass of the surroundings involved in the heat exchange measured in g,

c is the specific heat capacity, the amount of energy required to raise the temperature of 1 g of the substance by 1 °C. It is measured in $J\,g^{-1}\,K^{-1}$.

ΔT is the change in temperature:

the final temperature − the initial temperature (measured in °C)



WORKED EXAMPLE

An excess of magnesium is added to 50 cm³ of aqueous copper sulfate solution.
The temperature of the solution increases from 21 °C to 32 °C.

Calculate the heat exchanged in this reaction.

Specific heat capacity of the solution, c, is $4.18\,J\,g^{-1}\,K^{-1}$.

The solution has a density of $1.00\,g\,cm^{-3}$.

$$Q = mc\Delta T = 50 \times 4.18 \times 11 = 2299\,J$$

PRACTICE QUESTIONS

1 An excess of zinc is added to 100 cm³ of dilute sulfuric acid. The temperature of the solution increases from 19 °C to 24 °C.
Calculate the heat exchanged in this reaction.
Specific heat capacity of the solution, c, is $4.18\,J\,g^{-1}\,K^{-1}$.
The solution has a density of $1.00\,g\,cm^{-3}$.

2 An excess of iron is added to 150 cm³ of dilute copper sulfate solution. The temperature of the solution increases from 18 °C to 32 °C.
Calculate the heat exchanged in this reaction.
Specific heat capacity of the solution, c, is $4.18\,J\,g^{-1}\,K^{-1}$.
The solution has a density of $1.00\,g\,cm^{-3}$.

Determining the enthalpy change of reaction

The enthalpy change of a reaction can be calculated from the heat exchange and the amount of substance, in mol, that has reacted.

CALCULATING AN ENTHALPY CHANGE

An excess of iron is added to 50 cm³ of $2.0\,mol\,dm^{-3}$ aqueous copper(II) sulfate solution.

$$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$$

The temperature of the solution increases from 22 °C to 35 °C.

Specific heat capacity of the solution, c, is $4.18\,J\,g^{-1}\,K^{-1}$.

The solution has a density of $1.00\,g\,cm^{-3}$.

a) Find the energy exchanged in this reaction.
$Q = mc\Delta T = 50 \times 4.18 \times 13 = 2717\,J$ or $2.717\,kJ$

b) Find the amount, in mol, of copper(II) sulfate solution used in the reaction.
$n = V$ (in dm³) × concentration $= 0.05 \times 2.0 = 0.10\,mol$

c) Calculate the enthalpy change in $kJ\,mol^{-1}$.
To scale up the enthalpy change to $kJ\,mol^{-1} = \dfrac{\text{energy exchange (kJ)}}{\text{amount of substance (mol)}} = \dfrac{2.717\,kJ}{0.10\,mol} = 27.17\,kJ\,mol^{-1}$

As the temperature of the surroundings increased, heat was lost by the chemical system and gained by the surroundings. This means the reaction was exothermic and the enthalpy change of the reaction has a negative sign.

The enthalpy change is $-27.17\,kJ\,mol^{-1}$.

PRACTICE QUESTIONS

3 An excess of magnesium is added to 100 cm³ of $1.00\,mol\,dm^{-3}$ hydrochloric acid.
The temperature increases from 19 °C to 27 °C.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

Specific heat capacity of the solution, c, is $4.18\,J\,g^{-1}\,K^{-1}$.
The solution has a density of $1.00\,g\,cm^{-3}$.

a Calculate the energy change for the reaction.
b Calculate the amount, in mol, of hydrochloric acid used.
c Calculate the enthalpy change for the reaction.

4 An excess of zinc is added to 150 cm³ of $1.00\,mol\,dm^{-3}$ nitric acid. The temperature increases from 22 °C to 29 °C.

$$Zn(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$$

Specific heat capacity of the solution, c, is $4.18\,J\,g^{-1}\,K^{-1}$.
The solution has a density of $1.00\,g\,cm^{-3}$.

a Calculate the energy change for the reaction.
b Calculate the amount, in mol, of hydrochloric acid used.
c Calculate the enthalpy change for the reaction.

STRETCH YOURSELF

Experimental and theoretical values

The enthalpy change calculated from experimental results is less exothermic or less endothermic than the theoretical value for the reaction. This is due to heat loss from the surroundings. This reduces the temperature change that is detected during the experiment.

PRACTICE QUESTION

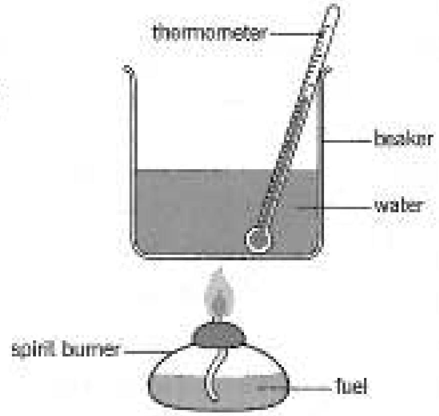
5 How could a student reduce heat loss during an experiment to determine the enthalpy change of reaction between iron and an aqueous solution of copper(II) sulfate?

Enthalpy changes 2

Standard enthalpy change of combustion, $\Delta H_c^{\ominus}$

The standard enthalpy change of combustion, $\Delta H_c^{\ominus}$, is the enthalpy change when one mole of a substance is completely burnt in oxygen, under standard conditions with all reactants and products being in their standard states.

Enthalpy changes of combustion are always exothermic so the sign for enthalpy change of combustion reactions is negative.



The chemical equation for the standard enthalpy change of combustion of methane is:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

The enthalpy change of combustion can be determined from experimental data.

A known mass of fuel is burnt and heats up a known volume of water. The temperature change of the water is found.

Heat exchanged

The heat lost by the system and gained by the surroundings can be calculated using the relationship:

$$Q = mc\Delta T$$

where Q is the heat exchange measured in J,

m is the mass of the water involved in the heat exchange measured in g.

c is the specific heat capacity of water, the amount of energy required to raise the temperature of 1 g of the substance by 1 °C. It has a value of 4.18 J g^{-1} K^{-1}.

ΔT is the change in temperature of the water:

the final temperature − the initial temperature (measured in °C)

CALCULATING HEAT EXCHANGE

Burning ethanol

During the combustion of a sample of ethanol, 100 cm^3 of water was heated from 21 °C to 35 °C

Calculate the heat exchanged in this reaction.

Specific heat capacity of water, c is 4.18 J g^{-1} K^{-1}.

Water has a density of 1.00 g mol^{-1}.

$\Delta T = 35 - 21 = 14$ K

$Q = mc\Delta T = 100 \times 4.18 \times 14 = 5852$ J or 5.852 kJ

PRACTICE QUESTIONS

1 A student burns a sample of propan-1-ol. During the experiment the temperature of 250 cm^3 of water increases from 18 °C to 24 °C.
Calculate the heat exchanged in this reaction.
Specific heat capacity of water, c, is 4.18 J g^{-1} K^{-1}.
Water has a density of 1.00 g mol^{-1}.

2 During combustion of a sample of ethanol the temperature of 150 cm^3 of water increased from 24 °C to 39 °C.
Calculate the heat exchanged in this reaction.
Specific heat capacity of water, c, is 4.18 J g^{-1} K^{-1}.
Water has a density of 1.00 g mol^{-1}.

3 During combustion of a sample of butan-2-ol the temperature of 200 cm^3 of water increased from 15 °C to 27 °C.
Calculate the heat exchanged in this reaction.
Specific heat capacity of water, c, is 4.18 J g^{-1} K^{-1}.
Water has a density of 1.00 g mol^{-1}.

Determining the enthalpy change of combustion

The enthalpy change of a combustion reaction can be calculated from the heat gained by the surroundings and the amount of substance, in mol, that has been burnt.

CALCULATING THE ENTHALPY CHANGE OF COMBUSTION

During combustion, 0.46 g of ethanol, C_2H_5OH, heated 100 cm^3 of water. The temperature of the water increased from 22 °C to 48 °C.

Specific heat capacity of water, c, is 4.18 J g^{-1} K^{-1}.

Water has a density of 1.00 g cm^{-3}.

a) Find the energy exchanged in this reaction.
$Q = mc\Delta T = 100 \times 4.18 \times 26 = 10\,868$ J or 10.868 kJ

b) Find the amount, in mol, of ethanol that was used in this reaction.
Molar mass of ethanol, C_2H_5OH = 46.0 g mol^{-1}
$n = \frac{m}{M} = \frac{0.46}{46.0} = 0.01$ mol

c) Calculate the enthalpy change in kJ mol^{-1}.
To scale up the enthalpy change to kJ mol^{-1}:
$$\frac{\text{energy exchange (kJ)}}{\text{amount of substance (mol)}} = \frac{10.868\text{ kJ}}{0.01\text{ mol}} = 1086.8\text{ kJ mol}^{-1}$$
The enthalpy change is −1086.8 kJ mol^{-1}.

d) What is the significance of the sign for the enthalpy change?
The negative sign shows the reaction is exothermic.

PRACTICE QUESTIONS

4 During combustion, 1.80 g of propan-1-ol, C_3H_7OH, heated 200 cm^3 of water. The temperature of the water increased from 20 °C to 51 °C.
Specific heat capacity of water, c is 4.18 J g^{-1} K^{-1}.
Water has a density of 1.00 g cm^{-3}.
Give your answers to three significant figures.
a Find the heat energy change in this reaction.
b Find the amount, in mol, of propan-1-ol used.
c Find the enthalpy change for this reaction in kJ mol^{-1}.
d What is the significance of the sign of the enthalpy change?

5 During combustion, 2.30 g of ethanol, C_2H_5OH, heated 100 cm^3 of water. The temperature of the water increased from 20 °C to 32 °C.
Specific heat capacity of water, c, is 4.18 J g^{-1} K^{-1}.
Water has a density of 1.00 g cm^{-3}.
Give your answers to three significant figures.
a Find the heat energy change in this reaction.
b Find the amount, in mol, of ethanol used.
c Find the enthalpy change for this reaction in kJ mol^{-1}.

Bond enthalpy 1

Exothermic and endothermic reactions

Energy is released when bonds are made and required when bonds are broken. Bond enthalpy is the enthalpy change when one mole of a given bond in a gaseous molecule is broken by homolytic fission.

Sometimes the same bond, e.g. C—H, can be found in many different molecules. The exact value of the bond enthalpy varies slightly from one molecule to another. The average bond enthalpy is the average enthalpy change for breaking one mole of the bond in a variety of gaseous molecules.

The enthalpy change for all exothermic reactions has a negative sign. More energy is released when new, stronger bonds are formed than was required to break the old, weaker bonds.

In contrast, the enthalpy change for all endothermic reactions has a positive sign. More energy is required to break the old, stronger bonds than is released when the new, weaker bonds are formed.

CALCULATING THE ENTHALPY CHANGE FOR REACTIONS

Formation of hydrogen chloride

Determine the enthalpy change for the reaction between hydrogen and chlorine to produce hydrogen chloride using the bond enthalpies shown in the table.

$H_2 + Cl_2 \rightarrow 2HCl$

Bond	Bond enthalpy/kJ mol^{-1}
H—H	+436
H—Cl	+432
Cl—Cl	+243

H—H + Cl—Cl → H—Cl
H—Cl

The bonds broken in this reaction = 1 × H—H = 1 × 436 = 436 kJ mol^{-1} and
1 × Cl—Cl = 1 × 243 = 243 kJ mol^{-1}

The total amount of energy required = 436 + 243 = 679 kJ mol^{-1}

The bonds made in this reaction = 2 × H—Cl = 2 × 432 = 864 kJ mol^{-1}

The total amount of energy released = 864 kJ mol^{-1}

So the enthalpy change for the reaction = total energy required − energy released
= 679 − 864 = −185 kJ mol^{-1}

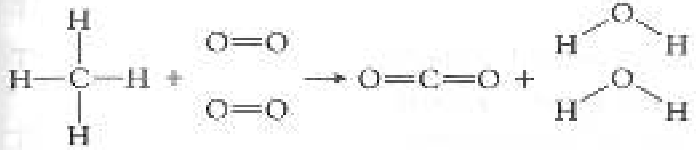
Combustion of methane

Determine the enthalpy change for the complete combustion of methane using the bond enthalpies in this table.

Bond	Bond enthalpy/kJ mol^{-1}
O=O	+497
C=O	+805
O—H	+463
C—H	+413

During combustion, methane reacts with oxygen to form carbon dioxide and water.

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$



The bonds broken in this reaction = 4 × C—H = 4 × 413 = 1652 kJ mol^{-1} and
2 × O=O = 2 × 497 = 994 kJ mol^{-1}

The total amount of energy required = 1652 + 994 = 2646 kJ mol^{-1}

The bonds made in this reaction = 2 × C=O = 2 × 805 = 1610 kJ mol^{-1} and
4 × O—H = 4 × 463 = 1852 kJ mol^{-1}

The total amount of energy released = 1610 + 1852 = 3462 kJ mol^{-1}

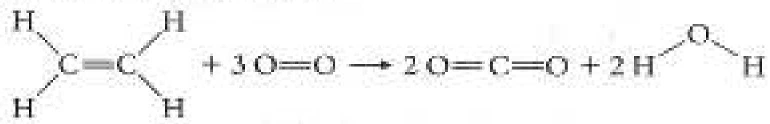
So the enthalpy change for the reaction = total energy required − energy released
= 2646 − 3462 = −816 kJ mol^{-1}

PRACTICE QUESTION

1 Use the bond enthalpy values in the table below to determine the enthalpy change for the following reactions:

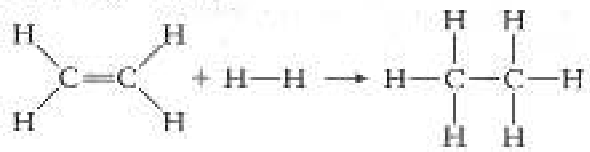
a The complete combustion of ethene

$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$



b The reaction between ethene and hydrogen to produce ethane

$C_2H_4 + H_2 \rightarrow C_2H_6$



c The reaction between oxygen and hydrogen to produce water

$O_2 + 2H_2 \rightarrow 2H_2O$

Bond	Bond enthalpy/kJ mol^{-1}
C—H	+413
C—C	+347
O=O	+497
C=O	+805
O—H	+463
H—H	+436
C=C	+612

REMEMBER: In these calculations the sign of the enthalpy change of reaction is significant. Exothermic reactions have a negative sign, while endothermic reactions have a positive sign. Always include the sign in your answers.

Bond enthalpy 2

Energy and bonds

Energy is released when bonds are made and required when bonds are broken.

An unknown bond enthalpy can be calculated if the enthalpy change for the reaction and the other bond enthalpies involved in the reaction are known.

FINDING AN UNKNOWN BOND ENTHALPY VALUE

Hydrogen bromide

A chemist adds hydrogen to bromine to produce hydrogen bromide.

$H_2 + Br_2 \rightarrow 2HBr$

$$\text{H—H} + \text{Br—Br} \rightarrow \begin{array}{c}\text{H—Br}\\ \text{H—Br}\end{array}$$

The enthalpy change for the reaction is $-99\,\text{kJ mol}^{-1}$.

Determine the bond enthalpy value for the Br—Br bond using the bond enthalpy values shown in the table below.

Bond	Bond enthalpy/kJ mol^{-1}
H—H	+436
H—Br	+364

The bonds broken in this reaction = 1 × H—H = 1 × 436 = 436 kJ mol^{-1} and
1 × Br—Br = ?

The total amount of energy required = 436 + Br—Br kJ mol^{-1}

The bonds made in this reaction = 2 × H—Br = 2 × 364 = 728 kJ mol^{-1}

The total amount of energy released = 728 kJ mol^{-1}

So, the enthalpy change for the reaction = total energy required − energy released

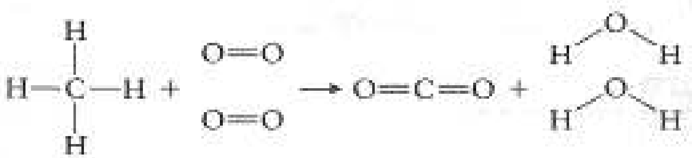
−99 = (436 + Br—Br) − 728

So, Br—Br = 193 kJ mol^{-1}

Combustion of methane

A student burnt methane to produce carbon dioxide and water.

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$



The enthalpy change for the reaction is $-810\,\text{kJ mol}^{-1}$.

Determine the bond enthalpy value for the C—H bond using the bond enthalpy values shown in the table below.

Bond	Bond enthalpy/kJ mol^{-1}
O=O	+500
C=O	+805
O—H	+463

The bonds broken in this reaction = 4 × C—H
2 × O=O = 2 × 500 = 1000 kJ mol^{-1}

The total amount of energy required = 4 × C—H + 1000 kJ mol^{-1}

The bonds made in this reaction = 2 × C=O = 2 × 805 = 1610 kJ mol^{-1}
4 × O—H = 4 × 463 = 1852 kJ mol^{-1}

The total amount of energy released = 1610 + 1852 = 3462 kJ mol^{-1}

So, the enthalpy change for the reaction = total energy required − energy released

−810 = 4 × C—H + 1000 − 3462

4 × C—H = 1652 kJ mol^{-1}

So C—H = $\frac{1652}{4}$ = 413 kJ mol^{-1}

REMEMBER: Experimental values for enthalpy changes may be slightly different to the theoretical values for enthalpy changes found using average bond enthalpy. The calculations use average bond enthalpies, which are average values for a particular bond in a variety of different gaseous molecules. The actual value for the bond in a given molecule may be slightly different.

PRACTICE QUESTIONS

1 A chemist adds hydrogen to chlorine to make hydrogen chloride.
The reaction is shown below.

$H_2 + Cl_2 \rightarrow 2HCl$

The enthalpy change for the reaction is found to be $-183\,\text{kJ mol}^{-1}$.
Use the bond enthalpy values shown in the table below to determine the bond enthalpy value for the Cl—Cl bond.

Bond	Bond enthalpy/kJ mol^{-1}
H—H	+436
H—Cl	+431

2 A chemist adds hydrogen to ethene to make ethane.
The reaction is shown below.

$H_2 + C_2H_4 \rightarrow C_2H_6$

The enthalpy change for the reaction is found to be $-119\,\text{kJ mol}^{-1}$.
Use the bond enthalpy values shown in the table below to determine the bond enthalpy value for the C—C bond.

Bond	Bond enthalpy/kJ mol^{-1}
C=C	+612
C—H	+410
H—H	+436

3 A chemist adds hydrogen to iodine to make hydrogen iodide.
The reaction is shown below.

$H_2 + I_2 \rightarrow 2HI$

The enthalpy change for the reaction is found to be $-7\,\text{kJ mol}^{-1}$.
Use the bond enthalpy values shown in the table below to determine the bond enthalpy value for the I—I bond.

Bond	Bond enthalpy/kJ mol^{-1}
H—H	+436
H—I	+297

Percentage yields

Actual and theoretical yields

Chemists can predict the amount of product made in a reaction.

The percentage yield links the actual amount of product made, in moles, and the theoretical yield, in moles.

$$\text{Percentage yield} = \frac{\text{actual amount (in moles) of product}}{\text{theoretical amount (in moles) of product}} \times 100\%$$

CALCULATING THE THEORETICAL YIELD

When calcium carbonate is heated it decomposes to form calcium oxide and carbon dioxide.

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

A chemist used 75.1 g of calcium carbonate, $CaCO_3$, in this reaction.

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

$$n = \frac{m}{M} = \frac{75.1}{100.1} = 0.750\ \text{mol}$$

b) Calculate the amount, in mol, of calcium oxide that was made.
From the balanced equation:
the amount of calcium carbonate, in mol = amount of calcium oxide, in mol = 0.750 mol

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

$$m = n \times M = 0.750 \times 56.1 = 42.1\ \text{g}$$

The actual yield

However, in practice chemists often find that the experiment actually makes a smaller amount of product than they had anticipated.

This could be because:

- other reactions take place which produce an alternative product
- the reaction reaches an equilibrium and so does not go to completion
- some of the product is lost during transfer or purification, for example, during filtering or distillation.

CALCULATING THE PERCENTAGE YIELD

A student added ethanol to propanoic acid to make the ester, ethyl propanoate, and water.

$$C_2H_5OH + C_2H_5COOH \rightarrow C_2H_5COOC_2H_5 + H_2O$$

The experiment has a theoretical yield of 5.00 g.

The actual yield is 4.50 g.

The molar mass of $C_2H_5COOC_2H_5 = 102.0\ \text{g mol}^{-1}$

Calculate the percentage yield of the reaction.

$$\text{Actual amount of ethyl propanoate, } n = \frac{m}{M} = \frac{4.5}{102} = 0.0441\ \text{mol}$$

$$\text{Theoretical amount of ethyl propanoate, } n = \frac{m}{M} = \frac{5.0}{102} = 0.0490\ \text{mol}$$

$$\text{Percentage yield} = \frac{4.5}{5.0} \times 100\% = 90\%$$

PRACTICE QUESTIONS

1. Calculate the percentage yield of a reaction that has a theoretical yield of 4.75 moles of product and an actual yield of 3.19 moles of product. Give your answer to three significant figures.
2. Calculate the percentage yield of a reaction that has a theoretical yield of 3.00 moles of product and an actual yield of 2.75 moles of product. Give your answer to three significant figures.
3. Calculate the percentage yield of a reaction that has a theoretical yield of 12.00 moles of product and an actual yield of 6.25 moles of product. Give your answer to three significant figures.

Limiting reagents

In some reactions, one of the reactants is stated as being in excess. The other reactant is said to be the limiting reagent. The limiting reagent will be used up first.

WORKED EXAMPLE

An excess of magnesium is added to 50 cm^3 of 1.0 mol dm^{-3} copper(II) sulfate solution.

$$Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$$

2.54 g of copper is produced.

Calculate the percentage yield of this reaction.

$$\text{The amount, in mol, of } CuSO_4 \text{ used, } n = \frac{V\ (\text{in cm}^3)}{1000} \times c = \frac{50}{1000} \times 1.0 = 0.05\ \text{mol}$$

The theoretical amount, in mol, of copper produced in this reaction is also 0.05 mol.

$$\text{The actual amount, in mol, of copper made} = \frac{m}{M} = \frac{2.54}{63.5} = 0.04\ \text{mol}$$

$$\text{Percentage yield} = \frac{0.04}{0.05} \times 100\% = 80\%$$

PRACTICE QUESTIONS

4. An excess of zinc is added to 25.0 cm^3 of 1.0 mol dm^{-3} iron(II) sulfate solution.
   $$Zn(s) + FeSO_4(aq) \rightarrow ZnSO_4(aq) + Fe(s)$$
   1.116 g of iron is produced.
   Calculate the percentage yield of this reaction.
5. An excess of magnesium is added to 50.0 cm^3 of 1.0 mol dm^{-3} aqueous hydrochloric acid solution.
   $$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$
   0.953 g of magnesium chloride is produced.
   Calculate the percentage yield of this reaction.
6. An excess of magnesium is added to 100.0 cm^3 of 0.5 mol dm^{-3} iron(II) sulfate solution.
   $$Mg(s) + FeSO_4(aq) \rightarrow MgSO_4(aq) + Fe(s)$$
   0.558 g of iron is produced.
   Calculate the percentage yield of this reaction.

Finding the reagent that is in excess

In some reactions the amount, in moles, of both reactants must be worked out. The balanced equation is then used to work out which of the reagents is in excess and which is the limiting reagent.

Atom economy

Wanted and unwanted products

Chemical reactions often produce more than one product. The product we want to produce is the desired product. The other (unwanted) products made in the reaction are called by-products.

Atom economy is a way of measuring the proportion of desired product compared with the total amount of product made in a reaction.

$$\%\ \text{atom economy} = \frac{\text{molar mass of the desired product}}{\text{molar mass of all the products}} \times 100\%$$

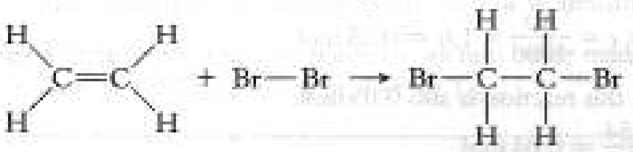
Some by-products are of no use and have to be disposed of. This can be expensive and adds to the overall costs of the process. In addition there are only a limited number of landfill sites available. Chemists try to find uses for by-products. If the by-products can be sold on, it not only has economic benefits but it also means that natural resources are conserved for future use and the amount of waste for landfill sites is reduced.

WORKING OUT THE % ATOM ECONOMY

Addition reactions

In addition reactions, reactants join together to form just one product.

Bromine is added to ethene to produce 1,2-dibromoethane.



REMEMBER: All addition reactions have a % atom economy of 100%; all the reactant atoms are found in the desired product.

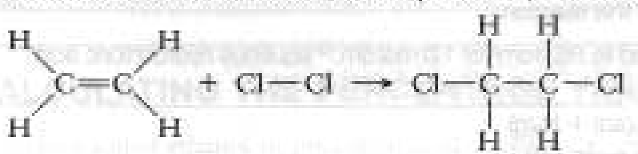
Determine the % atom economy of this reaction.

The molar mass of 1,2-dibromoethane = 187.8 g mol^{-1}

$$\%\ \text{atom economy} = \frac{\text{molar mass of the desired product}}{\text{molar mass of all the products}} \times 100\% = \frac{187.8}{187.8} \times 100 = 100\%$$

PRACTICE QUESTIONS

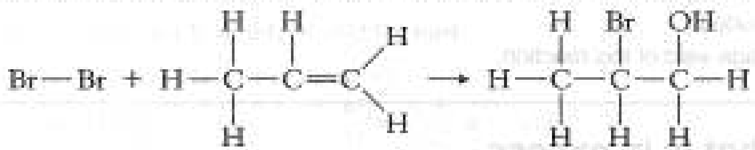
1 A chemist added chlorine to ethene to produce 1,2-dichloroethane.



Determine the % atom economy of this reaction.

The molar mass of 1,2-dichloroethane = 99.0 g mol^{-1}

2 A chemist adds bromine to propene to produce 1,2-dibromopropane.



Determine the % atom economy of this reaction.

The molar mass of 1,2-dibromopropane = 201.8 g mol^{-1}

FURTHER EXAMPLES

1) A student heated 2.5 g of calcium carbonate. It decomposed to form 1.2 g of the desired product, calcium oxide.

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

Determine the % atom economy of this reaction.

Give your answer to three significant figures.

Molar mass of CaO = 56.1 g mol^{-1}

Molar mass of CO_2 = 44.0 g mol^{-1}

$$\%\ \text{atom economy} = \frac{56.1}{100.1} \times 100 = 56.0\%$$

2) A chemist decides to prepare potassium nitrite, KNO_2, by decomposing potassium nitrate, KNO_3. The equation for the reaction is:

$2KNO_3 \rightarrow 2KNO_2 + O_2$

Determine the % atom economy of this reaction.

Give your answer to three significant figures.

The molar mass of KNO_2 = 85.1 g mol^{-1}

The molar mass of O_2 = 32.0 g mol^{-1}

Total molar mass of KNO_2 in the reaction = 170.2

Molar mass of all the products = 202.2

$$\%\ \text{atom economy} = \frac{170.2}{202.2} \times 100 = 84.2\%$$

PRACTICE QUESTIONS

3 A student prepares a sample of the ester ethyl ethanoate by reacting ethanol and ethanoic acid.

$C_2H_5OH + CH_3COOH \rightarrow CH_3COOC_2H_5 + H_2O$

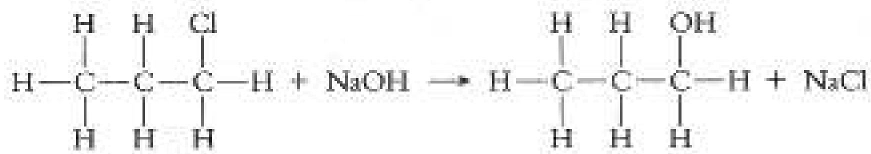
The molar mass of ethyl ethanoate = 88.0 g mol^{-1}

The molar mass of water = 18.0 g mol^{-1}

Determine the % atom economy of this reaction.

Give your answer to three significant figures.

4 A chemist prepares a sample of propan-1-ol by reacting 1-chloropropane with sodium hydroxide solution.



Determine the % atom economy of this reaction.

Give your answer to three significant figures.

5 A chemist prepares a sample of ethanol by reacting sodium hydroxide solution with bromoethane.

$NaOH + CH_3CH_2Br \rightarrow CH_3CH_2OH + NaBr$

Determine the % atom economy of this reaction.

Give your answer to three significant figures.

Percentage error in apparatus

Calculating percentage error

The percentage error of a measurement is calculated from the maximum error for the piece of apparatus being used and the value measured.

$$\text{Percentage error} = \frac{\text{maximum error}}{\text{measured value}} \times 100\%$$

PERCENTAGE ERROR

An excess of zinc powder was added to 50 cm^3 of copper(II) sulfate solution to produce zinc sulfate and copper.

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

The initial and final temperature of the solution was recorded and used to work out the temperature change. The experiment was completed by four students who recorded their results in the table below.

Student	Initial temperature/°C	Final temperature/°C	Temperature change/°C
A	21.0	24.9	3.9
B	21.2	25.0	3.8
C	21.0	26.5	5.5
D	21.1	25.1	4.0

a) What is the advantage of comparing the results of four sets of experiments?

So you can spot any anomalous results.

b) What is the average temperature change for this reaction?

Result C appears to be anomalous results so should not be included in the average.

$$\text{The average} = \frac{(3.9 + 3.8 + 4.0)}{3} = 3.9\,°C$$

c) The measuring cylinder used to measure the copper(II) sulfate solution has a maximum error of ± 2 cm^3.

Calculate the percentage error.

$$\text{Percentage error} = \frac{2}{50} \times 100\% = 4\%$$

d) A thermometer has a maximum error of ± 0.05 °C.

i) Calculate the percentage error when the thermometer is used to record a temperature of 25.0 °C.

$$\text{Percentage error} = \frac{0.05}{25.0} \times 100\% = 0.2\%$$

ii) Calculate the percentage error when the thermometer is used to record a temperature rise of 3.9 °C Give your answer to three significant figures.

$$\text{Percentage error} = \frac{(2 \times 0.05)}{3.9} \times 100\% = 2.56\%$$

Notice that two temperatures are required to calculate the temperature change so the maximum error is doubled.

PRACTICE QUESTIONS

1 A measuring cylinder has a maximum error of ± 1 cm^3. Calculate the maximum error when recording these values. Give your answers to three significant figures.

a 25.0 cm^3 b 80.0 cm^3 c 38.0 cm^3

2 A thermometer has a maximum error of ± 0.5 °C. Calculate the maximum error when recording these values. Give your answers to three significant figures.

a 10.0 °C b 15.0 °C c 83.0 °C

3 A gas syringe has a maximum error of ± 0.5 cm^3. Calculate the maximum error when recording these values. Give your answers to three significant figures.

a 21.0 cm^3 b 26.0 cm^3 c 43.0 cm^3

4 A thermometer has a maximum error of ± 0.5 °C. Calculate the maximum error when recording these temperature rises. Give your answers to three significant figures.

a 12.0 °C b 21.0 °C c 37.6 °C

PERCENTAGE ERROR IN TITRATIONS

A chemist carries out a titration between sodium hydroxide solution and hydrochloric acid.

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

A pipette is used to transfer 25.0 cm^3 of sodium hydroxide solution to a flask. The pipette has a maximum error of ± 0.5 cm^3.

The results for the amount of hydrochloric acid required in the titration are shown below. The burette has a maximum error of ± 0.05 cm^3.

Experiment	Volume of hydrochloric acid/cm^3
1	22.20
2	21.80
3	21.70

a) What is the average volume of hydrochloric acid used?

Result 1 appears to be anomalous results so should not be included in the average.

$$\text{The average} = \frac{(21.80 + 21.70)}{2} = 21.75\ cm^3$$

b) Calculate the maximum percentage error in the measurement of the volume of hydrochloric acid. Give your answer to three significant figures.

Two volume readings are required to calculate the volume of hydrochloric acid added from the burette. Calculating the percentage error in the volume of hydrochloric acid used in the first experiment:

$$\text{Percentage error} = \frac{(2 \times 0.05)}{22.2} \times 100 = 0.450\%$$

PRACTICE QUESTION

5 The experiment above was repeated by another group and the new results are shown opposite.

a What is the average volume of hydrochloric acid used?

b Calculate the maximum error in the measurement of the volume of hydrochloric acid in Experiment 1. Give your answer to three significant figures.

Experiment	Volume of hydrochloric acid/cm^3
1	21.60
2	21.70
3	21.30