1 STOICHIOMETRIC RELATIONSHIPS

Particulate nature of matter

ELEMENTS

All substances are made up of one or more elements. An element cannot be broken down by any chemical process into simpler substances. There are just over 100 known elements. The smallest part of an element is called an atom.

Names of the first 20 elements

Atomic Number	Name	Symbol	Relative atomic mass
1	hydrogen	H	1.01
2	helium	He	4.00
3	lithium	Li	6.94
4	beryllium	Be	9.01
5	boron	B	10.81
6	carbon	C	12.01
7	nitrogen	N	14.01
8	oxygen	O	16.00
9	fluorine	F	19.00
10	neon	Ne	20.18
11	sodium	Na	22.99
12	magnesium	Mg	24.31
13	aluminium	Al	26.98
14	silicon	Si	28.09
15	phosphorus	P	30.97
16	sulfur	S	32.07
17	chlorine	Cl	35.45
18	argon	Ar	39.95
19	potassium	K	39.10
20	calcium	Ca	40.08

COMPOUNDS

Some substances are made up of a single element, although there may be more than one atom of the element in a particle of the substance. For example, oxygen is diatomic, that is, a molecule of oxygen contains two oxygen atoms and has the formula O_2. A compound contains more than one element combined chemically in a fixed ratio. For example, a molecule of water contains two hydrogen atoms and one oxygen atom. It has the formula H_2O. Water is a compound not an element because it can be broken down chemically into its constituent elements: hydrogen and oxygen. Compounds have different chemical and physical properties from their component elements.

MIXTURES

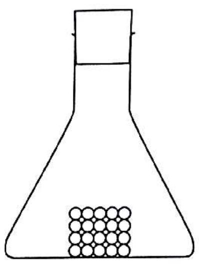
The components of a mixture may be elements or compounds. These components are not chemically bonded together. Because they are not chemically combined, the components of a mixture retain their individual properties. All the components of a mixture may be in the same phase, in which case the mixture is said to be **homogeneous**. Air is an example of a gaseous homogenous mixture. If the components of a mixture are in different phases the mixture is said to be **heterogeneous**. There is a physical boundary between two phases. A solid and a liquid is an example of a two-phase system. It is possible to have a single state but two phases. For example, two immiscible liquids such as oil and water form a heterogeneous mixture.

STATES OF MATTER

sublimation

SOLID STATE

- Fixed shape
- Fixed volume
- Particles held together by intermolecular forces in a fixed position
- Particles can vibrate about a fixed point but do not have translational velocity
- As heat is supplied at a certain temperature the vibration is sufficient to overcome the attractive forces holding the solid together and the solid melts

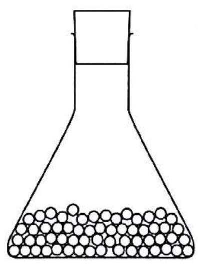


melting

freezing

LIQUID STATE

- Fixed volume
- Takes up shape of container
- Particles held closely together by intermolecular forces
- Particles have translational velocity so diffusion can occur
- As heat is supplied the liquid particles move faster. Some particles move faster than others and escape from the surface of the liquid to form a vapour. Once the pressure of the vapour is equal to the pressure above the liquid the liquid boils.

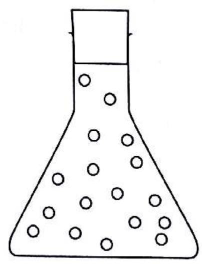


boiling/
vaporization/
evaporation

condensation

GASEOUS STATE

- Widely spaced particles that completely fill container
- Pressure of the gas due to gaseous particles colliding with the walls of the container
- Intermolecular forces between particles negligible
- Volume occupied by molecules themselves negligible compared with total volume of gas
- Particles moving with rapid, random motion so diffusion can occur



deposition

The mole concept and chemical formulas

MOLE CONCEPT AND AVOGADRO'S CONSTANT

A single atom of an element has an extremely small mass. For example, an atom of carbon-12 has a mass of 1.993×10^{-23} g. This is far too small to weigh. A more convenient amount to weigh is 12.00 g. 12.00 g of carbon-12 contains 6.02×10^{23} atoms of carbon-12. This number is known as Avogadro's constant (N_A or L).

Chemists measure amounts of substances in moles. A mole is the amount of substance that contains L particles of that substance. The mass of one mole of **any** substance is known as the **molar mass** and has the symbol M. For example, hydrogen atoms have $\frac{1}{12}$ of the mass of carbon-12 atoms so a mole of hydrogen atoms contains 6.02×10^{23} hydrogen atoms and has a mass of 1.01 g. In reality elements are made up of a mixture of isotopes.

The **relative atomic mass** of an element A_r is the weighted mean of all the naturally occurring isotopes of the element relative to carbon-12. This explains why the relative atomic masses given for the elements on page 1 are not whole numbers. The units of molar mass are g mol^{-1} but relative molar masses M_r have no units. For molecules **relative molecular mass** is used. For example, the M_r of glucose, $C_6H_{12}O_6 = (6 \times 12.01) + (12 \times 1.01) + (6 \times 16.00) = 180.18$. For ionic compounds the term **relative formula mass** is used.

Be careful to distinguish between the words **mole** and **molecule**. A molecule of hydrogen gas contains two atoms of hydrogen and has the formula H_2. A mole of hydrogen gas contains 6.02×10^{23} hydrogen molecules made up of two moles (1.20×10^{24}) of hydrogen atoms.

FORMULAS OF COMPOUNDS

Compounds can be described by different chemical formulas.

Empirical formula (literally the formula obtained by experiment)

This shows the simplest whole number ratio of atoms of each element in a particle of the substance. It can be obtained by either knowing the mass of each element in the compound or from the percentage composition by mass of the compound. The percentage composition can be converted directly into mass by assuming 100 g of the compound are taken.

Example: A compound contains 40.00% carbon, 6.73% hydrogen and 53.27% oxygen by mass, determine the empirical formula.

	Amount / mol	Ratio	
C	40.00/12.01 = 3.33	1	Empirical formula $= CH_2O$
H	6.73/1.01 = 6.66	2	
O	53.27/16.00 = 3.33	1	

Molecular formula

For molecules this is much more useful as it shows the actual number of atoms of each element in a molecule of the substance. It can be obtained from the empirical formula if the molar mass of the compound is also known.

Methanal CH_2O ($M_r = 30$), ethanoic acid $C_2H_4O_2$ ($M_r = 60$) and glucose $C_6H_{12}O_6$ ($M_r = 180$) are different substances with different molecular formulas but all with the same empirical formula CH_2O. Note that subscripts are used to show the number of atoms of each element in the compound.

Structural formula

This shows the arrangement of atoms and bonds within a molecule and is particularly useful in organic chemistry.

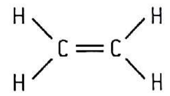
The three different formulas can be illustrated using ethene:

CH_2

empirical formula

C_2H_4

molecular formula



(can also be written $H_2C{=}CH_2$)

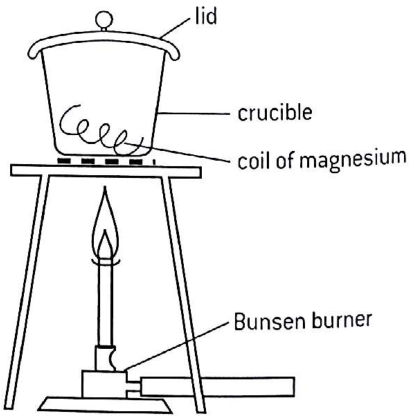
structural formula

EXPERIMENTAL DETERMINATION OF AN EMPIRICAL FORMULA

The empirical formula of magnesium oxide can be determined simply in the laboratory. A coil of magnesium ribbon about 10 cm long is placed in a pre-weighed crucible and its mass recorded. The crucible is placed on a clay triangle and heated strongly. When the magnesium ribbon starts to burn the lid is lifted slightly to allow more air to enter and the heating is continued until all the magnesium has burned. After cooling the crucible, its lid and its contents are reweighed.

Table of typical raw quantitative data:

	Mass / g (± 0.001 g)
Mass of crucible + lid	30.911
Mass of crucible + lid + magnesium	31.037
Mass of crucible + lid + magnesium oxide	31.106



Mass of magnesium = 31.037 − 30.911 = 0.126 g

Mass of magnesium oxide = 31.106 − 30.911 = 0.195 g

Mass of oxygen combining with magnesium = 0.195 − 0.126 = 0.069 g

Amount of magnesium $= \frac{0.126}{24.31} = 5.2 \times 10^{-3}$ mol

Amount of oxygen $= \frac{0.069}{16.00} = 4.3 \times 10^{-3}$ mol

Ratio of Mg to O $= \frac{5.2 \times 10^{-3}}{4.3 \times 10^{-3}} = 1.2$ to 1

Convert to whole number ratio = 6 : 5

Empirical formula of magnesium oxide as determined by this experiment is Mg_6O_5.

Chemical reactions and equations

PROPERTIES OF CHEMICAL REACTIONS

In all chemical reactions:

- new substances are formed
- bonds in the reactants are broken and bonds in the products are formed resulting in an energy change between the reacting system and its surroundings
- there is a fixed relationship between the number of particles of reactants and products resulting in no overall change in mass – this is known as the stoichiometry of the reaction.

CHEMICAL EQUATIONS

Chemical reactions can be represented by chemical equations. Reactants are written on the left-hand side and products on the right-hand side. The number of moles of each element must be the same on both sides in a balanced chemical equation, e.g. the reaction of nitric acid (one of the acids present in acid rain) with calcium carbonate (the main constituent of marble statues).

$$\underbrace{\underset{\text{calcium carbonate}}{CaCO_3(s)} + \underset{\text{nitric acid}}{2HNO_3(aq)}}_{\text{REACTANTS}} \rightarrow \underbrace{\underset{\text{calcium nitrate}}{Ca(NO_3)_2(aq)} + \underset{\text{carbon dioxide}}{CO_2(g)} + \underset{\text{water}}{H_2O(l)}}_{\text{PRODUCTS}}$$

STATE SYMBOLS

Because the physical state that the reactants and products are in can affect both the rate of the reaction and the overall energy change it is good practice to include the state symbols in the equation.

(s) – solid (l) – liquid (g) – gas (aq) – in aqueous solution

→ OR ⇌

A single arrow → is used if the reaction goes to completion. Sometimes the reaction conditions are written on the arrow:

e.g. $C_2H_4(g) + H_2(g) \xrightarrow{\text{Ni catalyst, 180 °C}} C_2H_6(g)$

Reversible arrows are used for reactions where both the reactants and products are present in the equilibrium mixture:

e.g. $3H_2(g) + N_2(g) \underset{\text{250 Atm}}{\overset{\text{Fe(s), 550 °C}}{\rightleftharpoons}} 2NH_3(g)$

COEFFICIENTS AND MOLAR RATIO

The coefficient refers to the number in front of each reactant and product in the equation. The coefficients give information on the molar ratio. In the first example above, two moles of nitric acid react with one mole of calcium carbonate to produce one mole of calcium nitrate, one mole of carbon dioxide and one mole of water. In the reaction between hydrogen and nitrogen above, three moles of hydrogen gas react with one mole of nitrogen gas to produce two moles of ammonia gas.

IONIC EQUATIONS

Because ionic compounds are completely dissociated in solution it is sometimes better to use ionic equations. For example, when silver nitrate solution is added to sodium chloride solution a precipitate of silver chloride is formed.

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$$

$Na^+(aq)$ and $NO_3^-(aq)$ are spectator ions and do not take part in the reaction. So the ionic equation becomes:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

From this we can deduce that any soluble silver salt will react with any soluble chloride to form a precipitate of silver chloride.

Mass and gaseous volume relationships

SOLIDS

Normally measured by weighing to obtain the mass.

1.000 kg = 1000 g

When weighing a substance the mass should be recorded to show the accuracy of the balance. For example, exactly 16 g of a substance would be recorded as 16.00 g on a balance weighing to + or − 0.01 g but as 16.000 g on a balance weighing to + or − 0.001 g.

MEASUREMENT OF MOLAR QUANTITIES

In the laboratory moles can conveniently be measured using either mass or volume depending on the substances involved.

SOLUTIONS

Volume is usually used for solutions.

1.000 litre = 1.000 dm^3 = 1000 cm^3

Concentration is the amount of **solute** (dissolved substance) in a known volume of **solution** (solute plus **solvent**). It is expressed either in g dm^{-3}, or, more usually in mol dm^{-3}. For very dilute solutions it is also sometimes expressed in parts per million, ppm. A solution of known concentration is known as a **standard solution**.

To prepare a 1.00 mol dm^{-3} solution of sodium hydroxide dissolve 40.00 g of solid sodium hydroxide in distilled water and then make the total volume up to 1.00 dm^3.

Concentration is often represented by square brackets, e.g.

$[NaOH(aq)] = 1.00$ mol dm^{-3}

A 25.0 cm^3 sample of this solution contains

$1.00 \times \frac{25.0}{1000} = 2.50 \times 10^{-2}$ mol of NaOH

LIQUIDS

Pure liquids may be weighed or the volume recorded.

The density of the liquid $= \frac{\text{mass}}{\text{volume}}$ and is usually expressed in g cm^{-3}.

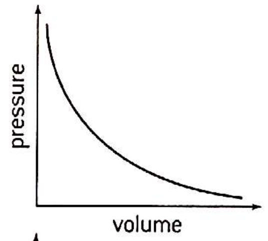
GASES

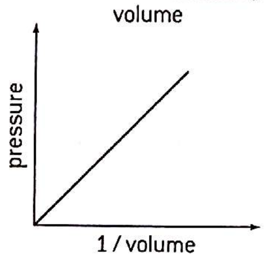
Mass or volume may be used for gases.

CHANGING THE VARIABLES FOR A FIXED MASS OF GAS

$P \propto \frac{1}{V}$ (or PV = constant)

At constant temperature: as the volume decreases the concentration of the particles increases, resulting in more collisions with the container walls. This increase in pressure is inversely proportional to the volume, i.e. doubling the pressure halves the volume.





$P \propto T$ (or $\frac{P}{T}$ = constant)

At constant volume: increasing the temperature increases the average kinetic energy so the force with which the particles collide with the container walls increases. Hence pressure increases and is directly proportional to the absolute temperature, i.e. doubling the absolute temperature doubles the pressure.

IDEAL GAS EQUATION

The different variables for a gas are all related by the ideal gas equation.

$PV = nRT$

P = pressure in Pa (N m^{-2})
(1 atm = 1.013×10^5 Pa)

T = absolute temperature in K

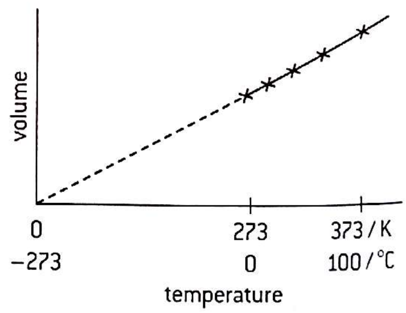
V = volume in m^3
(1 cm^3 = 1×10^{-6} m^3)

n = number of moles

R = gas constant = 8.314 J K^{-1} mol^{-1}

$V \propto T$ (or $\frac{V}{T}$ = constant)

At constant pressure: at higher temperatures the particles have a greater average velocity so individual particles will collide with the container walls with greater force. To keep the pressure constant there must be fewer collisions per unit area so the volume of the gas must increase. The increase in volume is directly proportional to the absolute temperature, i.e. doubling the absolute temperature doubles the volume.



Extrapolating the graph to zero volume gives the value for absolute zero.

UNITS

The gas constant can be expressed in different units but it is easier to use SI units.

$$R = \frac{PV}{nT} = \frac{\text{N m}^{-2} \times \text{m}^3}{\text{mol} \times \text{K}} = \text{N m mol}^{-1}\text{ K}^{-1}$$
$$= \text{J K}^{-1}\text{ mol}^{-1}$$

REAL GASES

An ideal gas exactly obeys the gas laws. Real gases do have some attractive forces between the particles and the particles themselves do occupy some space so they do not exactly obey the laws. If they did they could never condense into liquids. A gas behaves most like an ideal gas at high temperatures and low pressures.

Molar volume of a gas and calculations

MOLAR VOLUME OF A GAS

The ideal gas equation depends on the amount of gas (number of moles of gas) but not on the nature of the gas. Avogadro's Law states that equal volumes of different gases at the same temperature and pressure contain the same number of moles. From this it follows that one mole of any gas will occupy the same volume at the same temperature and pressure. This is known as the molar volume of a gas. At 273 K and 1.00×10^5 Pa pressure this volume is 22.7×10^{-2} m³ (22.7 dm³ or 22 700 cm³).

When the mass of a particular gas is fixed (nR is constant) a useful expression to convert the pressure, temperature and volume under one set of conditions (1) to another set of conditions (2) is:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

In this expression there is no need to convert to SI units as long as the same units for pressure and volume are used on both sides of the equation. However do not forget that T refers to the absolute temperature and must be in kelvin.

CALCULATIONS FROM EQUATIONS

Work methodically.

Step 1. Write down the correct formulas for all the reactants and products.

Step 2. Balance the equation to obtain the correct stoichiometry of the reaction.

Step 3. If the amounts of all reactants are known work out which are in **excess** and which one is the limiting reagent. By knowing the **limiting reagent** the maximum **yield** of any of the products can be determined.

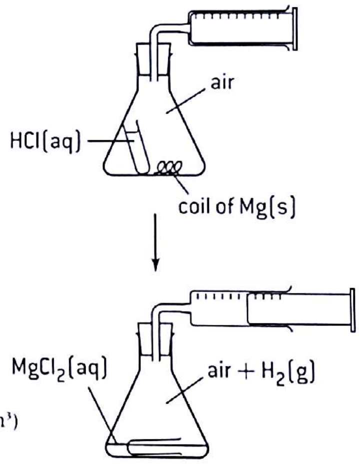
Step 4. Work out the amount (in mol) of the substance required.

Step 5. Convert the amount (in mol) into the mass or volume.

Step 6. Express the answer to the correct number of significant figures and include the appropriate units.

WORKED EXAMPLES

(a) Calculate the volume of hydrogen gas evolved at 273 K and 1.00×10^5 Pa when 0.623 g of magnesium reacts with 27.3 cm³ of 1.25 mol dm⁻³ hydrochloric acid.



Equation:

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

A_r for Mg = 24.31. Amount of Mg present $= \frac{0.623}{24.31} = 2.56 \times 10^{-2}$ mol

Amount of HCl present $= 1.25 \times \frac{27.3}{1000} = 3.41 \times 10^{-2}$ mol

From the equation $2 \times 2.56 \times 10^{-2} = 5.12 \times 10^{-2}$ mol of HCl would be required to react with all of the magnesium.

Therefore the magnesium is in excess and the limiting reagent is the hydrochloric acid.

The maximum amount of hydrogen produced $= \frac{3.41 \times 10^{-2}}{2} = 1.705 \times 10^{-2}$ mol

Volume of hydrogen at 273 K, 1.00×10^5 Pa $= 1.705 \times 10^{-2} \times 22.7 = 0.387$ dm³ (or 387 cm³)

(b) Calculate the volume occupied by the hydrogen evolved. In the example above if it had been collected at 22 °C and at a pressure of 1.12×10^5 Pa

Step 1. Express the temperature as an absolute temperature

22 °C = 295 K

Step 2. Apply the ideal gas equation $pV = nRT$

$1.12 \times 10^5 \times V = 1.705 \times 10^{-2} \times 8.314 \times 295$

$$V = \frac{1.705 \times 10^{-2} \times 8.314 \times 295}{1.12 \times 10^5} = 3.73 \times 10^{-4}\ \text{m}^3\ (373\ \text{cm}^3)$$

This could also be solved using $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$$V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1} = 0.387 \times \frac{1.00 \times 10^{-5}}{1.12 \times 10^{-5}} \times \frac{295}{273}$$

$= 0.373$ cm³ (373 cm³)

(c) The actual volume of hydrogen collected under the conditions stated in (a) was 342 cm³. Determine the percentage yield.

Step 1. Use the mole ratio from the equation and the amounts of reactants to determine the limiting reagent and hence the theoretical maximum yield. From part (a) theoretical yield = 387 cm³

Step 2. Apply the relationship:

$$\text{Percentage yield} = \frac{\text{Experimental yield}}{\text{Theoretical yield}} \times 100$$

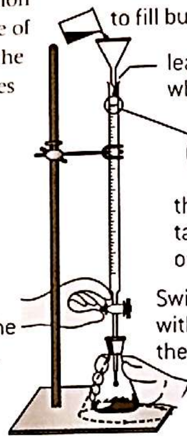
$$\text{Percentage yield} = \frac{342}{387} \times 100 = 88.4\%$$

Titration and atom economy

DETERMINING AN UNKNOWN CONCENTRATION BY TITRATION

Titration is a useful technique to find the concentration of a solution of unknown concentration by reacting it with a stoichiometric amount of a standard solution. A known accurate volume of one of the solutions is placed in a conical flask using a pipette. A burette is then used to add the other solution dropwise until the reaction is complete. This can be seen when one drop causes the solution to just change colour. For acid–base titrations, it is usual to add an indicator but this is not always necessary for some other types of titration, e.g. redox titrations using acidified potassium permanganate, as the reactant itself causes the colour change.

It is usual to obtain at least two accurate readings, which should be within 0.15 cm^3 of each other.



Worked examples

1. 25.00 cm^3 of a solution of sodium hydroxide of unknown concentration required 23.65 cm^3 of 0.100 mol dm^{-3} hydrochloric acid solution for complete neutralization. Calculate the concentration of the sodium hydroxide solution.

   Equation for the reaction: $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

   Amount of hydrochloric acid present in 23.65 cm^3 = $\left(\frac{23.65}{1000}\right) \times 0.100 = 2.365 \times 10^{-3}$ mol

   Since one mol of NaOH reacts with one mol of HCl

   Amount of sodium hydroxide present in 25.00 cm^3 = 2.365×10^{-3} mol

   Concentration of sodium hydroxide = $2.365 \times 10^{-3} \times \left(\frac{1000}{25.00}\right) = 0.0946$ mol dm^{-3}

2. 50.0 cm^3 of 1.00 mol dm^{-3} hydrochloric acid solution, HCl(aq) was added to some egg shell with a mass of 2.016 g. After all the egg shell had reacted the resulting solution was put into a 100 cm^3 volumetric flask and the volume made up to the mark with distilled water. 10.0 cm^3 of this solution required 11.40 cm^3 of 1.00×10^{-1} mol dm^{-3} sodium hydroxide solution, NaOH(aq) for complete neutralization. Calculate the percentage of calcium carbonate in the egg shell.

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

   Amount of sodium hydroxide present in 11.40 cm^3 = $\left(\frac{11.40}{1000}\right) \times 1.00 \times 10^{-1} = 1.140 \times 10^{-3}$ mol

   Since one mol of NaOH reacts with one mol of HCl

   Amount of diluted excess hydrochloric acid in 10.0 cm^3 = 1.140×10^{-3} mol

   Amount of excess hydrochloric acid in 100 cm^3 = $(1.140 \times 10^{-3}) \times 10 = 1.140 \times 10^{-2}$ mol

   Initial amount of hydrochloric acid added to egg shell = $\left(\frac{50.0}{1000}\right) \times 1.00 = 5.00 \times 10^{-2}$ mol

   Amount of hydrochloric acid reacting with egg shell = $(5.00 \times 10^{-2}) - (1.140 \times 10^{-2}) = 3.860 \times 10^{-2}$ mol

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

   Amount of $CaCO_3$ that reacted with the acid = ½ × $3.860 \times 10^{-2} = 1.930 \times 10^{-2}$ mol

   M_r $(CaCO_3) = 40.08 + 12.01 + (3 \times 16.00) = 100.09$

   Mass of $CaCO_3$ in egg shell = $100.09 \times 1.930 \times 10^{-2} = 1.932$ g

   Percentage of calcium carbonate in the egg shell = $\left(\frac{1.932}{2.016}\right) \times 100 = 95.8\%$

ATOM ECONOMY (AN EXAMPLE OF UTILIZATION)

As well as trying to achieve high yields in industrial processes, chemists try to increase the conversion efficiency of a chemical process. This is known as atom economy. Ideally in a chemical process no atom is wasted. The atom economy is a measure of the amount of starting materials that become useful products. A high atom economy means that fewer natural resources are used and less waste is created. The atom economy can be calculated by using the following steps:

1. Write the balanced equation for the reaction taking place.
2. Calculate the relative molecular mass of each product and then the total mass of each product formed assuming molar quantities. Note that this is the same as the total mass of the reactants.
3. Calculate the relative molecular mass of each desired product and then the total mass of each desired product formed assuming molar quantities.
4. Atom economy = $\frac{\text{total mass of desired product(s)}}{\text{total mass of all products}} \times 100$

For example, consider the production of iron by the reduction of iron(III) oxide using the thermite reaction.

$2Al(s) + Fe_2O_3(s) \rightarrow 2Fe(s) + Al_2O_3(s)$

The total mass of products formed = $2 \times 55.85 + [(2 \times 26.98) + (3 \times 16.00)] = 213.66$ g

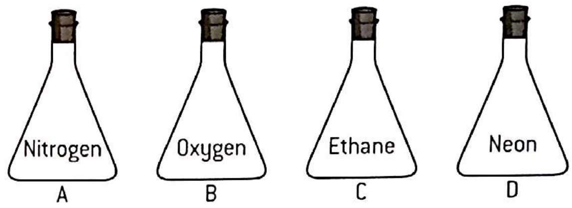
The total amount of iron (the desired product) formed = $2 \times 55.85 = 111.70$ g

The atom economy for this reaction is $\frac{111.70}{213.66} \times 100 = 52.3\%$

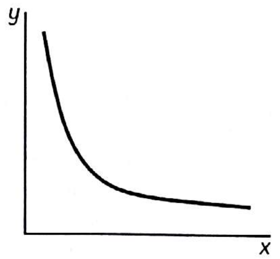
Obviously if a use can also be found for all the aluminium oxide produced then the atom economy for this reaction will increase to 100%.

MULTIPLE CHOICE QUESTIONS – STOICHIOMETRIC RELATIONSHIPS

1. How many oxygen **atoms** are in 0.100 mol of $CuSO_4.5H_2O$?
   A. 5.42×10^{22} C. 2.41×10^{23}
   B. 6.02×10^{22} D. 5.42×10^{23}
2. Which is not a true statement?
   A. One mole of methane contains four moles of hydrogen atoms
   B. One mole of ^{12}C has a mass of 12.00 g
   C. One mole of hydrogen gas contains 6.02×10^{23} atoms of hydrogen
   D. One mole of methane contains 75% of carbon by mass
3. A pure compound contains 24 g of carbon, 4 g of hydrogen and 32 g of oxygen. No other elements are present. What is the empirical formula of the compound?
   A. $C_2H_4O_2$ C. CH_4O
   B. CH_2O D. CHO
4. What is the mass in grams of one molecule of ethanoic acid CH_3COOH?
   A. 0.1 C. 1×10^{-22}
   B. 3.6×10^{25} D. 60
5. What is the relative molecular mass, M_r of carbon dioxide, CO_2?
   A. 44.01 g mol^{-1} C. 44.01 kg mol^{-1}
   B. 44.01 mol g^{-1} D. 44.01
6. Which of the following changes of state is an exothermic process?
   A. melting C. vaporizing
   B. condensing D. boiling
7. What is the empirical formula for the compound $C_6H_5(OH)_2$?
   A. C_6H_6O C. C_6H_7O
   B. $C_6H_5O_2H_2$ D. $C_6H_7O_2$
8. Phosphorus burns in oxygen to produce phosphorus pentoxide P_4O_{10}. What is the sum of the coefficients in the balanced equation?
   $\_P_4(s) + \_O_2(g) \rightarrow \_P_4O_{10}(s)$
   A. 3 C. 6
   B. 5 D. 7
9. Magnesium reacts with hydrochloric acid according to the following equation:
   $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
   What mass of hydrogen will be obtained if 100 cm^3 of 2.00 mol dm^{-3} HCl are added to 4.86 g of magnesium?
   A. 0.2 g C. 0.8 g
   B. 0.4 g D. 2.0 g
10. Butane burns in oxygen according to the equation below.
   $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$
   If 11.6 g of butane is burned in 11.6 g of oxygen which is the limiting reagent?
   A. Butane C. Neither
   B. Oxygen D. Oxygen and butane
11. Four identical containers under the same conditions are filled with gases as shown below. Which container and contents will have the highest mass?



12. What is the amount, in moles, of sulfate ions in 100 cm^3 of 0.020 mol dm^{-3} $FeSO_4(aq)$?
   A. 2.0×10^{-3} C. 2.0×10^{-1}
   B. 2.0×10^{-2} D. 2.0
13. 300 cm^3 of water is added to a solution of 200 cm^3 of 0.5 mol dm^{-3} sodium chloride. What is the concentration of sodium chloride in the new solution?
   A. 0.05 mol dm^{-3} C. 0.2 mol dm^{-3}
   B. 0.1 mol dm^{-3} D. 0.3 mol dm^{-3}
14. Separate samples of two gases, each containing a pure substance, are found to have the same density under the same conditions of temperature and pressure. Which statement about these two samples must be correct?
   A. They have the same volume
   B. They have the same relative molecular mass
   C. There are equal numbers of moles of gas in the two samples
   D. They condense at the same temperature
15. The graph below represents the relationship between two variables in a fixed amount of gas.



Which variables could be represented by each axis?

	x-axis	*y*-axis
A.	pressure	temperature
B.	volume	temperature
C.	pressure	volume
D.	temperature	volume

16. Sulfuric acid and sodium hydroxide react together according to the equation:
   $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
   What volume of 0.250 mol dm^{-3} NaOH is required to neutralize exactly 25.0 cm^3 of 0.125 mol dm^{-3} H_2SO_4?
   A. 25.0 cm^3 C. 50 cm^3
   B. 12.5 cm^3 D. 6.25 cm^3

SHORT ANSWER QUESTIONS – STOICHIOMETRIC RELATIONSHIPS

1. Aspirin, $C_9H_8O_4$, is made by reacting ethanoic anhydride, $C_4H_6O_3$ ($M_r = 102.1$), with 2-hydroxybenzoic acid ($M_r = 138.1$), according to the equation:

$$2C_7H_6O_3 + C_4H_6O_3 \rightarrow 2C_9H_8O_4 + H_2O$$

   a) If 15.0 g 2-hydroxybenzoic acid is reacted with 15.0 g ethanoic anhydride, determine the limiting reagent in this reaction. [3]

   b) Calculate the maximum mass of aspirin that could be obtained in this reaction. [2]

   c) If the mass obtained in this experiment was 13.7 g, calculate the percentage yield of aspirin. [1]

2. 14.48 g of a metal sulfate with the formula M_2SO_4 was dissolved in water. Excess barium nitrate solution was added in order to precipitate all the sulfate ions in the form of barium sulfate. 9.336 g of precipitate was obtained.

   a) Calculate the amount of barium sulfate $BaSO_4$ precipitated. [2]

   b) Calculate the amount of sulfate ions present in the 14.48 g of M_2SO_4. [1]

   c) Deduce the relative molar mass of M_2SO_4. [1]

   d) Calculate the relative atomic mass of M and hence identify the metal. [2]

3. A student added 7.40×10^{-2} g of magnesium ribbon to 15.0 cm^3 of 2.00 mol dm^{-3} hydrochloric acid. The hydrogen gas produced was collected using a gas syringe at 20.0 °C and 1.00×10^5 Pa.

   a) State the equation for the reaction between magnesium and hydrochloric acid. [1]

   b) Determine the limiting reactant. [3]

   c) Calculate the theoretical yield of hydrogen gas:

      (i) in mol [1]

      (ii) in cm^3, under the stated conditions of temperature and pressure. [2]

   d) The actual volume of hydrogen measured was lower than the calculated theoretical volume. Suggest two reasons why the volume of hydrogen gas obtained was less. [2]

4. In 1921 Thomas Midgley discovered that the addition of a lead compound could improve the combustion of hydrocarbons in automobile (car) engines. This was the beginning of the use of leaded gasoline (petrol).

   The percentage composition, by mass, of the lead compound used by Midgley is Pb: 64.052%, C: 29.703% and H: 6.245%.

   a) (i) Determine the empirical formula of the lead compound. [3]

      (ii) Leaded gasoline has been phased out because the lead(IV) oxide, PbO_2, produced as a side product in the combustion reaction may cause brain damage in children.
      0.01 mol of Midgley's lead compound produces 0.01 mol of lead(IV) oxide. Deduce the molecular formula of Midgley's compound. [1]

      (iii) Determine the equation for the complete combustion of Midgley's compound. [2]

   b) The combustion of unleaded gasoline still produces pollution with both local and global consequences. Identify one exhaust gas that causes local pollution and one exhaust gas that causes global pollution. [2]

5. An experiment was performed to determine the percentage of iron present in a sample of iron ore. 3.682×10^{-1} g of the sample was dissolved in acid and all of the iron was converted to Fe^{2+}.

   The resulting solution was titrated with a standard solution of potassium manganate(VII), $KMnO_4$ with a concentration of 2.152×10^{-2} mol dm^{-3}. The end point was indicated when one drop caused a slight pink colour to remain. It was found that 22.50 cm^3 of the potassium manganate(VII) solution was required to reach the end point.

   In acidic solution, MnO_4^- reacts with Fe^{2+} ions to form Mn^{2+} and Fe^{3+} ions according to the following equation:

$$MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$$

   a) Calculate the amount (in mol) of MnO_4^- used in the titration. [2]

   b) Calculate the amount (in mol) of Fe present in the 3.682×10^{-1} g sample of iron ore. [2]

   c) Determine the percentage by mass of Fe present in the 3.682×10^{-1} g sample of iron ore. [2]

6. Copper metal may be produced by the reaction of copper(I) oxide and copper(I) sulfide according to the equation.

$$2Cu_2O(s) + Cu_2S(s) \rightarrow 6Cu(s) + SO_2(g)$$

   A mixture of 10.0 kg of copper(I) oxide and 5.00 kg of copper(I) sulfide was heated until no further reaction occurred.

   a) Determine the limiting reagent in this reaction. [3]

   b) Calculate the maximum mass of copper that could be obtained from these masses of reactants. [2]

   c) Assuming the reaction to produce copper goes to completion according to the equation deduce the atom economy for this reaction. [3]

7. The empirical formula of magnesium oxide is MgO. Suggest four assumptions that were made in the experiment detailed on page 2 that may not be true and which might account for the wrong result being obtained. [4]

8. The percentage composition by mass of a hydrocarbon is C: 85.6% and H: 14.4%.

   a) Calculate the empirical formula of the hydrocarbon. [2]

   b) A 1.00 g sample of the hydrocarbon at a temperature of 273 K and a pressure of 1.00×10^5 Pa has a volume of 0.405 dm^3.

      (i) Calculate the molar mass of the hydrocarbon. [2]

      (ii) Deduce the molecular formula of the hydrocarbon. [2]

   c) Explain why the incomplete combustion of hydrocarbons is harmful to humans. [2]