

Bleaches are Examples of Oxidising Agents (e.g. NaOCl) or Reducing Agents (e.g. SO₂)

Both require moisture to work

Sodium Hypochlorite acts as a Bleach by Oxidising

- Permanent
- Strong and used to bleach cotton – can damage delicate fibres
- Forms nascent oxygen (oxygen atoms which are very reactive)

Sulphur Dioxide acts by Reduction

- Temporary
- Mild – so used to bleach silk

Nomenclature of Transition Metal Compounds.

Oxidation numbers of elements may change from compound to compound, especially those containing Transition Elements

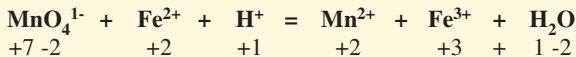
Transition elements have their oxidation state included in their names

e.g. KMnO₄ is potassium manganate (VII) MnO₂ is manganese (IV) oxide
CuCl is copper (I) chloride Fe₂O₃ is iron (III) oxide

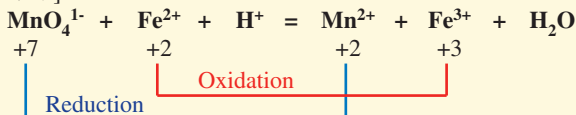
Balancing Complex Equations using Oxidation Numbers

Write unbalanced equation then

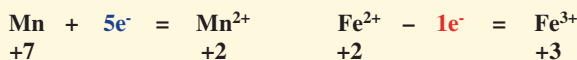
Assign Oxidation Numbers to all atoms



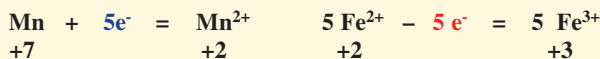
Select those which change and assign oxidation [increase in Ox. N^o.] and reduction [decrease in Ox. N^o.]



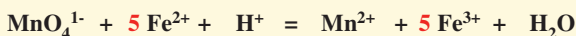
Write down electron gain and loss for each



Balance the electron transfer by multiplying the Fe equation by five

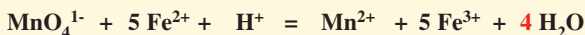


Now put these values into the equation

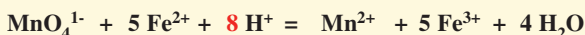


Finally balance the equation by inspection

4 O in MnO₄¹⁻ so need 4 H₂O to balance



There are 8 H in 4 H₂O so we need 8 H⁺ to balance these



Questions on this Section from Past Exams Year by Year

2009	2008	2007	2006	2005	2004	2003	2002
11 b	4 g	4 c	4 i,	4 a,b,	4 b,c,g,	5 b, c	4 e
	5	5 b	5 b,	5 c	5a		10 a
	10 b	10 c	10 b				

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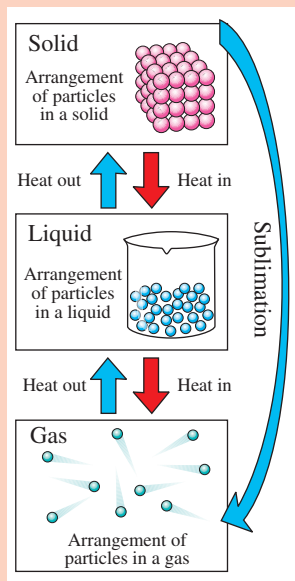
a 3.1 States of Matter

Motion of Particles in Solids, Liquids and Gases

Solid: Particles very closely packed and tightly bound so they have fixed shape and fixed volume. Vibrate constantly in all directions and as this increases with temperature it causes expansion. At melting point they break free from each other causing a greater increase in volume as they are less tightly packed.

Liquid: Particles are close together but can move over one another freely thus a liquid can flow. The volume is fixed at any given temperature. As temperature increases they spread out more and at boiling point they break free completely from the other particles and there is a large increase in volume

Gas: Particles are relatively far apart and free to move in all directions so it has no fixed volume and no fixed shape.



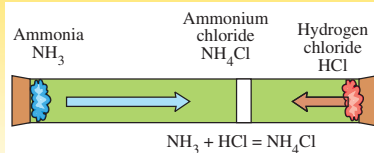
b Diffusion

Spontaneous spreading out of a substance due to the natural movement of its particles.

EXPERIMENT:

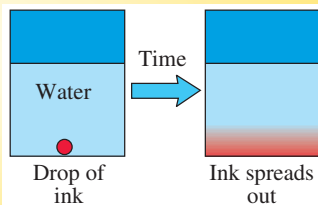
Demonstration NH_3 and Hydrogen Chloride

Ammonia and HCl diffuse along the tube. Ammonia diffuses faster because it is lighter. Where they meet is shown by a layer of ammonium chloride which precipitates.



Demonstration Ink and Water

If a pipette is used to carefully place a drop of ink in the bottom of a beaker of water and it is left for a week or two then the ink will gradually diffuse throughout the water. Slower than diffusion in gas



Smoke and Air

If someone is smoking in the room next door you can often smell it because the smoke diffuses through gaps in the wall or around the door.

c 3.2 Gas Laws

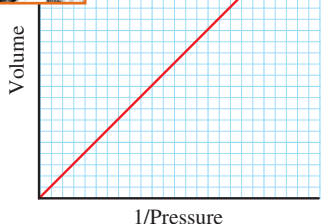
Boyle's Law

Boyle

At a constant temperature, the volume of a given mass of any gas is inversely proportional to the pressure on the gas.

Boyle and his assistant Robert Hooke developed an improved air pump which helped him carry out his experiments into pressure-volume relationships.

$$P \times V = \text{constant}$$



Volume inversely proportional to pressure

d Charles's Law

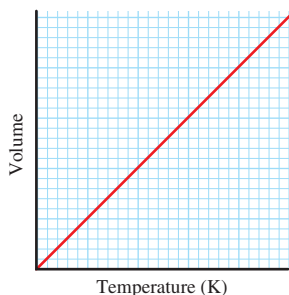
At a constant pressure, the volume of a given mass of any gas is directly proportional to the Kelvin temperature.

$$\frac{V}{T} = \text{constant}$$



Charles

Volume directly proportional to temperature



Gay-Lussac's Law of Combining Volumes

When gases react, the volumes consumed in the reaction bear a simple whole number ratio to each other, and to the volumes of any gaseous products of the reaction, all volumes being measured under the same conditions of temperature and pressure



Gay Lussac

Avogadro's Law

Equal volumes of gases, under the same conditions of temperature and pressure contain equal numbers of molecules



Avogadro

Combined Gas Law

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} = \text{constant.}$$

Correction of Gas Volumes to s.t.p. (units: Pa, cm³, K).

A definite mass of gas occupies 500 cm³ at 77°C and a pressure of 100 000 Pa. What would its volume be at s.t.p.?

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\begin{aligned} V_1 &= \frac{P_2 V_2}{P_1 T_2} \\ &= \frac{100\,000 \times 500 \times 273}{101\,325 \times 350} \\ &= 384.90 \text{ cm}^3 \end{aligned}$$

s.t.p

$$P_1 = 101\,325 \text{ Pa}$$

$$V_1 = ?$$

$$T_1 = 273 \text{ K}$$

Experimental

$$P_2 = 100\,000 \text{ Pa}$$

$$V_2 = 500 \text{ cm}^3$$

$$T_2 = 77 + 273 = 350 \text{ K}$$

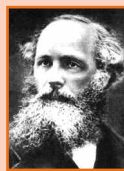
e 3.3 The Kinetic Theory of Gases:

Developed by James Clerk Maxwell and Ludwig Boltzmann

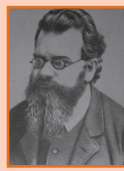
Ideal Gases

Obey all the assumptions of the Kinetic Theory under all conditions of temperature and pressure.

Gases are closest to ideal gas behaviour at low pressures and high temperatures



Clerk Maxwell



Boltzmann

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Assumptions of the Kinetic Theory

- Gases are made of particles whose **diameters are negligible compared to the distances between them**
- There are no **attractive or repulsive forces** between the particles
- The particles are in **constant, rapid, random motion** colliding with each other and the walls of the container
- All **collisions are perfectly elastic**
- The **average kinetic energy** of the particles is **proportional to the Kelvin temperature**.

Equation of State for an Ideal Gas:

$$PV = nRT \quad (\text{units: Pa, m}^3, \text{K}).$$

Calculate the volume occupied by 8 g of oxygen gas at 27°C and a pressure of 200 kPa.

Answer

$$P = 200 \text{ kPa} = 200 \times 10^3 \text{ Pa}$$

$$V = ?$$

$$n = \frac{\text{mass}}{\text{Relative Molecular Mass}} = \frac{8}{32} = 0.25 \text{ mole}$$

$$R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1}$$

$$T = 27 + 273 = 300 \text{ K}$$

$$PV = nRT$$

$$V = \frac{nRT}{P} = \frac{0.25 \times 8.31 \times 300}{200 \times 10^3} = 0.003116 \text{ m}^3$$

$$= 3.12 \text{ litres}$$

N.B. Converting units

$$\text{cm}^3 \rightarrow \text{m}^3 \text{ divide by } 1,000,000$$

$$\text{litres} \rightarrow \text{m}^3 \text{ divide by } 1,000$$

4.2 g of a gas occupies a volume of 1.5 litres at 77°C and 100,000 Pa.

- How many moles of the gas does the container contain?
- Calculate the relative molecular mass of the gas

Answer

$$(i) \quad PV = nRT$$

$$P = 100,000 \text{ Pa}$$

$$V = 1.5 \text{ litres} = 1.5 \times 10^{-3} \text{ m}^3$$

$$n = ?$$

$$R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1}$$

$$T = 273 + 77 = 350 \text{ K}$$

$$n = \frac{PV}{RT} = \frac{100,000 \times 1.5 \times 10^{-3}}{8.31 \times 350} = 0.05175$$

$$(ii) \quad 0.05175 \text{ moles} = 4.2 \text{ g}$$

$$1 \text{ mole} = \frac{4.2}{0.05175} = 81.44 \text{ g}$$

Reasons why Gases Deviate from Ideal Gas Behaviour

At low temperatures and high pressures

- There are **attractive or repulsive forces** between the particles
- The **diameters of particles are not negligible compared to the distances between them**

f 3.4 The Mole

$$\text{Avogadro constant} = 6 \times 10^{23} \text{ mol}^{-1}$$

The Mole

The SI unit for amount of substance containing the Avogadro number of particles.

Standard Temperature & Pressure (s.t.p.)

$$\text{Standard Temperature} = 273 \text{ K } (0^\circ\text{C})$$

$$\text{Standard Pressure } 101,325 \text{ Nm}^{-2} = 101,325 \text{ Pa} = 101 \text{ kPa}$$

Molar Volume at s.t.p

$$\text{Molar volume at STP} = 22.4 \text{ l} = 22,400 \text{ cm}^3 = 2.24 \times 10^{-2} \text{ m}^3$$

Relative Molecular Mass (M_r)

The mass of a molecule compared to $\frac{1}{12}$ of the mass of a carbon 12 isotope

Calculation of Relative Molecular Mass from Relative Atomic Masses

Add the relative atomic masses of its constituent atoms.

$$\text{H}_2\text{O} = (1 \times 2) + 16 = 18$$

$$(\text{NH}_4)_2\text{SO}_4 = [14 + (1 \times 4)] \times 2 + 32 + (16 \times 4) = 132$$

$$\text{MgCO}_3 \cdot 10\text{H}_2\text{O} = 24 + 12 + (16 \times 3) + 10 \times [(1 \times 2) + 16] = 264$$

Molar Mass

The mass in grams of a mole of that substance

3.5 Converting Moles to Grams, Litres & Number of Particles

Grams – multiply number of moles by Relative Molecular Mass

What is the mass of 0.2 moles of CaCO_3 ?

$$\text{CaCO}_3 = 40 + 12 + (16 \times 3) = 100$$

$$0.2 \text{ moles} = 0.2 \times 100 = 20\text{g}$$

Litres – multiply by 22.4 if STP or 24 if RTP

What is the volume of 2 moles of H_2 at STP?

$$2 \times 22.4 = 44.8 \text{ litres}$$

Number of Particles – multiply by 6×10^{23}

How many molecules are there in 5 moles of $\text{Ba}(\text{NO}_3)_2$?

$$5 \times 6 \times 10^{23} = 30 \times 10^{23} = 3.0 \times 10^{24}$$

3.6 Converting Grams and Litres to Moles and Number of Particles

Grams to Moles – divide by Relative Molecular Mass (M_r)

How many moles are there in 88g of CO_2 ?

$$\text{Relative molecular mass } \text{CO}_2 = 12 + (16 \times 2) = 44$$

$$88 / 44 = 2 \text{ moles}$$

Grams to Number of Particles –

divide by Relative Molecular Mass then multiply by 6×10^{23}

How many molecules are there in 9 grams of water?

$$\text{Relative molecular mass } \text{H}_2\text{O} = (1 \times 2) + 16 = 18$$

$$9 / 18 \times 6 \times 10^{23} = 3 \times 10^{23}$$

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Litres to moles – divide by 22.4 (STP) or 24 (RTP) Room Temp. & Pressure

How many moles of SO_2 are there in 3 litres of the gas @ STP?

$$\frac{3}{22.4} = 0.13 \text{ moles}$$

cm^3 to moles – divide by 22,400 if STP or 24,000 if RTP

How many moles of NO_2 are there in 175 cm^3 of the gas @ STP?

$$175 / 22,400 = 0.0078125 \text{ moles}$$

Litres to Number of Particles – divide by 22.4 then multiply by 6×10^{23}

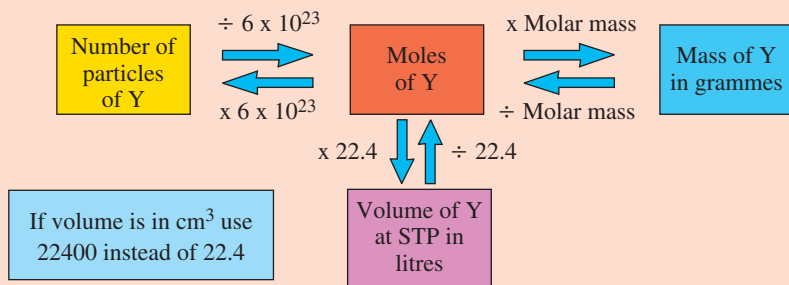
How many molecules are there in 4 litres of NH_3 ?

$$4 / 22.4 \times 6 \times 10^{23} = 1.07 \times 10^{23}$$

3·7 Converting Moles to Number of Atoms of a Molecular Species

How many atoms are there in 0.1 mole of ammonia [NH_3]?

$$0.1 \times 6 \times 10^{23} \times 4 = 2.4 \times 10^{23}$$



Mass Spectrometer to find Relative Molecular Mass of Compounds

The relative molecular mass of a compound can be found using a mass spectrometer. The molecules are ionised and broken up into positively charged pieces with different masses. These are separated and the amounts of each piece recorded. This is called the mass spectrum of the substance. The mass spectrum usually also contains the whole molecule as well as fragments

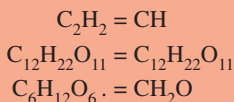
3·8 Chemical Formulas

Empirical Formula

Simplest ratio of one element to another in a compound

Ethyne has a molecular formula of C_2H_2 , sucrose $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ and glucose $\text{C}_6\text{H}_{12}\text{O}_6$.

What are their empirical formulas?



Calculation of Empirical Formulas, given the Percentage Composition by Mass

On analysis a compound was found to contain
C = 54.54%, H = 9.10% and O = 36.36%.

Calculate its empirical formula.

Imagine we have 100 g of the compound. This will contain
54.54 g carbon = 54.54/12 moles of carbon atoms = 4.55
9.10 g hydrogen = 9.10/1 moles of hydrogen atoms = 9.10
36.36 g oxygen = 36.36/16 moles of oxygen atoms = 2.27

4.55 : 9.10 : 2.27*

2 : 4 : 1



Alternative method

$$\frac{\%C}{A_r C} : \frac{\%H}{A_r H} : \frac{\%O}{A_r O}$$

$$\frac{54.54}{12} : \frac{9.10}{1} : \frac{36.36}{16}$$

$$4.55 : 9.1 : 2.27^*$$

$$2 : 4 : 1$$



* Divide each by the smallest number

Calculation of Empirical Formulas, given the Masses of Reactants & Product

1.44 g of magnesium was completely burned in oxygen and resulted in the formation of 2.40 g of magnesium oxide.
Calculate the empirical formula of magnesium oxide.

$$\text{Moles of Magnesium} = \frac{1.44}{24} = 0.06$$

$$\text{Mass of oxygen} = 2.40 - 1.44 = 0.96$$

$$\text{Moles of oxygen} = \frac{0.96}{16} = 0.06$$

$$\text{Mg} : \text{O} = 0.06 : 0.06 = 1 : 1$$

$$\text{Empirical formula} = \text{MgO}$$

Alternative method

$$\text{Mass of magnesium} = 1.44$$

$$\text{Mass of magnesium oxide} = 2.40$$

$$\text{Mass of Oxygen} = 2.40 - 1.44 = 0.96$$

$$\frac{\text{Mass Mg}}{A_r \text{ Mg}} : \frac{\text{Mass of O}}{A_r \text{ O}}$$

$$\frac{1.44}{24} : \frac{0.96}{16}$$

$$0.06 : 0.06$$

$$1 : 1$$



$$A_r = \text{Relative Atomic Mass}$$

k 3.9 Percentage Composition by Mass

Percentage by mass of each element in a molecule of a compound

What is the percentage of oxygen in H_2SO_4 ?

$$\text{Molecular mass} = (1 \times 2) + 32 + (16 \times 4) = 98$$

$$\frac{\text{Mass of Oxygen}}{\text{Molecular mass}} \times 100 = \frac{(16 \times 4)}{98} \times 100 = 65.3\%$$

Molecular Formula

Gives the actual numbers of atoms of each element in a molecule

I Structural Formulas

Shows the actual number and arrangement of atoms within a molecule of the substance

Compound	Empirical Formula	Molecular Formula	Structural Formula
Ethane	CH_3	C_2H_6	$\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array}$
Ethene	CH_2	C_2H_4	$\begin{array}{c} \text{H} \quad \text{H} \\ \diagdown \quad \diagup \\ \text{C}=\text{C} \\ \diagup \quad \diagdown \\ \text{H} \quad \text{H} \end{array}$
Ethyne	CH	C_2H_2	$\text{H}-\text{C}\equiv\text{C}-\text{H}$

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m Calculation of Molecular Formula given the Empirical Formulas and the Relative Molecular Mass

- Calculate the mass of an empirical formula unit
- Divide the relative molecular mass by this to find the number of units there are in a molecule.
- Multiply the empirical formula by this number.

Glucose has an empirical formula of CH_2O and a relative molecular mass of 180. Calculate its molecular formula.

Formula mass of $\text{CH}_2\text{O} = 12 + (1 \times 2) + 16 = 30$

Number of CH_2O units = $180 / 30 = 6$

Molecular formula = $6 \times \text{CH}_2\text{O} = \text{C}_6\text{H}_{12}\text{O}_6$

Urea has an empirical formula of CON_2H_4 and a relative molecular mass of 60. Find its molecular formula.

Formula mass = $12 + 16 + (14 \times 2) + (1 \times 4) = 60$

Number of CON_2H_4 units = $60/60 = 1$

Molecular formula = CON_2H_4

n 3.10 Chemical Equations

- Must contain the **correct molecular formula** of all the reactants and products
- Must be **balanced**

Balancing Chemical Equations

When balancing, the formulae of the compounds cannot be changed in any way, but can only be multiplied by appropriate numbers.

Carbon and oxygen react to form carbon monoxide. Write a balanced equation for this reaction.

$\text{C} + \text{O}_2 = \text{CO}$

One C on each side so C is OK

2 O's on left and only one on right so

CO must be multiplied by 2

$\text{C} + \text{O}_2 = 2 \text{CO}$

Now 1 C on left and 2 on right so C

must be multiplied by 2

$2 \text{C} + \text{O}_2 = 2 \text{CO}$

Now balanced – recheck to be sure

Methane burns in oxygen to form carbon dioxide and water. Write a balanced equation for the reaction.

$\text{CH}_4 + \text{O}_2 = \text{CO}_2 + \text{H}_2\text{O}$

One C on each side so C is OK

4 H's on left and only 2 on right so H_2O

must be multiplied by 2

$\text{CH}_4 + \text{O}_2 = \text{CO}_2 + 2 \text{H}_2\text{O}$

C and H now OK

2 O's on left and 4 on right so O_2 must

be multiplied by 2

$\text{CH}_4 + 2 \text{O}_2 = \text{CO}_2 + 2 \text{H}_2\text{O}$

Now balanced – recheck to be sure

Balancing Redox Equations

You will only be asked to balance ionic equations and should **ignore spectator ions**

Balance the following ionic equation

$\text{Fe}^{2+} + \text{Cl}_2 = \text{Fe}^{3+} + \text{Cl}^-$

One Fe on each side so Fe is OK

2 Cl on left side but only one on right so Cl^- must be multiplied by 2

$\text{Fe}^{2+} + \text{Cl}_2 = \text{Fe}^{3+} + 2 \text{Cl}^-$

This equation appears to be balanced but it is NOT as electron movements are not balanced – it needs to be balanced using oxidation numbers – see across

Balance the following ionic equation

$\text{Fe}^{2+} + \text{Cl}_2 = \text{Fe}^{3+} + \text{Cl}^-$

+2 0 +3 -1

Fe^{2+} loses an electron to become Fe^{3+}

$\text{Fe}^{2+} - 1\text{e}^- = \text{Fe}^{3+}$

Each Chlorine gains an electron to become Cl^-

But there are 2 Cl on the left so we need 2 on the right

$\text{Cl}_2 + 2\text{e}^- = 2 \text{Cl}^-$

We now need to multiply the Fe^{2+} to get a loss of 2e^-

$2 \text{Fe}^{2+} - 2 \text{e}^- = 2 \text{Fe}^{3+}$

Now we put these values into the equation

$2 \text{Fe}^{2+} + \text{Cl}_2 = 2 \text{Fe}^{3+} + 2 \text{Cl}^-$

Calculations Based on Balanced Equations using the Mole Concept

Balanced equations should be given for this type of calculation and should be in g and kg rather than tonnes

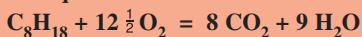
0 Calculations Involving Masses and Volumes

The first thing you need to do is change the quantities you are given into moles

Divide mass by the relative molecular mass

Divide volumes by 22.4 if given in litres or 22400 if given in cm^3

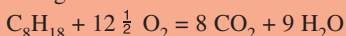
Octane burns in oxygen according to the equation



What volume of oxygen (measured at s.t.p.) is required for complete combustion of 11.4 g of octane?

$$\text{C}_8\text{H}_{18} = (8 \times 12) + (1 \times 18) = 114$$

$$11.4 \text{ g of octane} = 11.4 / 114 = 0.1 \text{ mole}$$



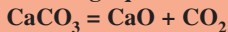
$$1 \text{ mole} \quad 12.5 \text{ mole}$$

$$0.1 \text{ mole} \quad 1.25 \text{ mole}$$

$$1 \text{ mole of O}_2 \text{ at s.t.p.} = 22.4 \text{ l}$$

$$1.25 \text{ moles} = 22.4 \times 1.25 = 28 \text{ litres}$$

Calcium carbonate decomposes when heated strongly according to the following equation



What mass of calcium oxide would be obtained if 50 g of calcium carbonate were decomposed?

$$\text{CaCO}_3 = 40 + 12 + (16 \times 3) = 100$$

$$50 \text{ g CaCO}_3 = 50 / 100 \text{ mole} = 0.5 \text{ mole}$$



$$1 \text{ mole} \quad 1 \text{ mole}$$

$$0.5 \text{ mole} \quad 0.5 \text{ mole}$$

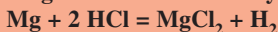
$$1 \text{ mole of CaO} = 40 + 16 = 56 \text{ g}$$

$$0.5 \text{ moles} = 56 \times 0.5 = 28 \text{ g}$$

Calculations Involving Excess of One Reactant

In most chemical reactions one of the reactants is present in excess

Magnesium reacts with hydrochloric acid according to the equation

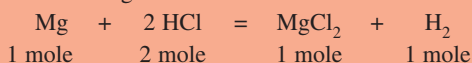


If 550 cm^3 aqueous HCl containing 3.65 g of HCl is added to 2 g of Mg

Show that the magnesium is present in excess.

$$\text{Moles of HCl} = 3.65 / 36.5 = 0.10 \text{ mole}$$

$$\text{Moles of Mg at start} \quad 2 / 24 = 0.08 \text{ mole}$$



$$0.05 \text{ mole} \quad 0.1 \text{ mole}$$

0.1 mole of HCl needs 0.05 mole Mg but we have 0.08 therefore we have excess of 0.03 mole of magnesium

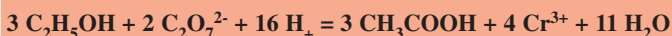
Find out how much Mg the 0.10 mole of HCl will react with since if magnesium is in excess then HCl is the limiting reagent.

Calculation of Percentage Yields

In many reactions the amount of product obtained is less than that predicted by the equation. Sometimes the reaction does not go to completion and at other times some product is lost while purifying it.

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

Ethanoic acid is prepared in the laboratory by reacting ethanol with excess acidified dichromate solution.



A sample of ethanoic acid was prepared in the laboratory by reacting 6 cm^3 of ethanol [Density = 0.8 g cm^{-3}] with excess acidified sodium dichromate.

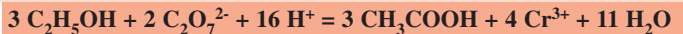
After purification it was found that 2.15g of ethanoic acid was collected. Calculate the percentage yield of ethanoic acid.

$$\text{Mass of ethanol} = 6 \times 0.8 \text{ g} = 4.80 \text{ g}$$

$$\text{Moles of ethanol} = 4.8 / 46 = 0.1 \text{ moles}$$

Mass = Density x Volume
is often used in exams so be aware of it. If you are given the density of a liquid it is a sure sign that this is required.

STOICHIOMETRY



3 moles 2 moles 16 moles **3 moles** 4 moles 11 moles

1 mole 1 mole

0.1 mole 0.1 mole

Mass of ethanoic acid in 0.1 mole = $0.1 \times 60 = 6 \text{ g}$

$$\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \frac{2.15}{6} \times 100 = \mathbf{35.83\%}$$

EXPERIMENT: To Measure the Relative Molecular Mass of a Volatile Liquid

-
- A diagram illustrating the setup for distilling propanone using a steam bath. A Bunsen burner is placed under a tripod stand, which supports a beaker of water. A flask containing propanone is suspended in the water by a clamp. The flask is covered with a foil top that has a hole in the center. Labels indicate the following components: Foil top with a hole in the centre, Clamp, Water, and Propanone.

WORKED EXAMPLE

In an experiment to measure the relative molecular mass of a volatile liquid 0.275 g of the liquid was vaporised at 97°C. The volume occupied was found to be 95 cm³.

The pressure was 1×10^5 Pa.

- (a) Calculate the number of moles of the volatile liquid vaporised.
(b) Calculate the relative molecular mass of the volatile liquid.

$$T = 97 + 273 = 370 \text{ K}$$

- (a) Let P_1 , V_1 and T_1 be STP and P_2 , V_2 and T_2 be experiment conditions

$$V_1 = \frac{P_2 V_2 T_1}{P_1 T_2} = \frac{1 \times 10^5 \text{ Pa} \times 95 \text{ cm}^3 \times 273 \text{ K}}{1.013 \times 10^5 \text{ Pa} \times 370 \text{ K}} = 69.19 \text{ cm}^3$$

$$\text{moles} = \frac{\text{volume}}{22,400} = \frac{69.19}{22,400} = \mathbf{0.00309 \text{ moles}}$$

- (b) $0.00309 \text{ moles} = 0.275 \text{ g}$

$$1 \text{ mole} = \frac{0.275}{0.00309} = 88.99 = \mathbf{89}$$

2009	2008	2007	2006	2005	2004	2003	2002
10 c	11 b	4 e, h	4 d	4 h	4 d, h	3	1 d
		10 b	11 a	10 b	10 a, c	4 b, g	4 i
				11 b			11 a

a 4.1 Concentration of Solutions

Solutions are intimate mixtures of a **solute** (e.g. salt) and a **solvent** (e.g. water)
A **concentrated** solution contains a large amount of solute per litre of solution
A **dilute** solution contains a small amount of solute per litre of solution

Concentration

Is the amount of solute in a specified amount of solution

Ways of Expressing Concentration

Moles per Litre mol l^{-1} (molarity)

This is the most useful way of expressing concentration for a chemist

If a solution contains **1 mole of solute in a litre of solution is said to be 1 molar [1M]**

If a solution contains 2 moles of solute in a litre of solution is said to be 2 molar [2M]

If a solution contains 0.5 mole of solute in a litre of solution is said to be 0.5 molar [0.5M]

Grams per litre g l^{-1}

This is the number of **grams of solute per litre of solution**

Parts per million p.p.m.

This is the number of **milligrams per litre – mg l^{-1}**

Percentage weight per volume % (w/v)

This is the number of grams of solute per 100 cm^3 of solution

A 5% sodium chloride solution contains **5g NaCl in 100 cm^3 of solution**

Percentage volume per volume % (v/v)

This is the number of grams of solute per 100 cm^3 of solution

A 5% vinegar solution contains **5 cm^3 of ethanoic acid per 100 cm^3 of vinegar.**

A bottle of wine labelled as 13% contains **13 cm^3 of ethanol per 100 cm^3 of wine**

Percentage weight per weight % (w/w)

A 2% Arnica ointment contains **2g of arnica per 100g of ointment**

b Calculation of Molarity from Concentration in Grams per Litre

$$\text{molarity} = \frac{\text{grams per litre}}{\text{molar mass}}$$

What is the molarity of a NaOH solution containing 4 g of NaOH per litre?

$$\text{Molarity} = \frac{\text{grams per litre}}{\text{molar mass}} = \frac{4}{40} = 0.1$$

Calculation of Concentration in Grams per Litre from Molarity

concentration in grams per litre = molar mass \times molarity

What is the concentration in grams per litre of a 0.1 M H_2SO_4 solution?

$$\begin{aligned} \text{Concentration in grams per litre} &= \text{molar mass} \times \text{molarity} \\ &= 98 \times 0.1 = 9.8 \text{ g l}^{-1} \end{aligned}$$

VOLUMETRIC ANALYSIS

Calculation of Number of Moles from Molarity and Volume

$$\text{Number of moles} = \frac{\text{volume}}{1000} \times \text{molarity}$$

How many moles are there in 250 cm³ of 0.1 M HCl?

$$\text{number of moles} = \frac{\text{volume}}{1000} \times \text{molarity} = \frac{250}{1000} \times 0.1 = 0.025$$

Calculations involving Percentage Concentrations

A 500 cm³ bottle of vinegar contains 25 cm³ of pure ethanoic acid.
What is its concentration as % (v/v)?

$$25 \text{ cm}^3 \text{ in } 500 \text{ cm}^3 \rightarrow 5 \text{ cm}^3 \text{ in } 100 \text{ cm}^3 \rightarrow 5\% \text{ (V/V)}$$

A 70 cl bottle of wine is labelled 13% (v/v). What volume of alcohol does it contain?

$$13\% \text{ (v/v)} \rightarrow 100 \text{ cm}^3 \text{ contains } 13 \text{ cm}^3$$

$$700 \text{ cm}^3 \text{ contains } 7 \times 13 \text{ cm}^3 = 91 \text{ cm}^3 \quad [70 \text{ cl} = 700 \text{ cm}^3]$$

250 cm³ of NaOH solution contains 20 g of NaOH.

What is its concentration as % (W/V)?

250 cm³ contains 20 g.

$$100 \text{ cm}^3 \text{ contains } \frac{20}{2.5} = 8\% \text{ (w/v)}$$

Colour Intensity as a Function of Concentration (*simple treatment only*)

If a solution is coloured then the intensity of the colour is proportional to the concentration.

Calculation of the Effect of Dilution on Concentration

$$V_{\text{dil}} \times M_{\text{dil}} = V_{\text{conc.}} \times M_{\text{conc.}}$$

What volume of 1 M NaOH solution is needed to make 300 cm³ of 0.05 M solution?

$$V_{\text{dil}} \times M_{\text{dil}} = V_{\text{conc.}} \times M_{\text{conc.}}$$

$$V_{\text{conc.}} = \frac{V_{\text{dil}} \times M_{\text{dil}}}{M_{\text{conc}}} = \frac{300 \times 0.05}{1} = 15 \text{ cm}^3$$

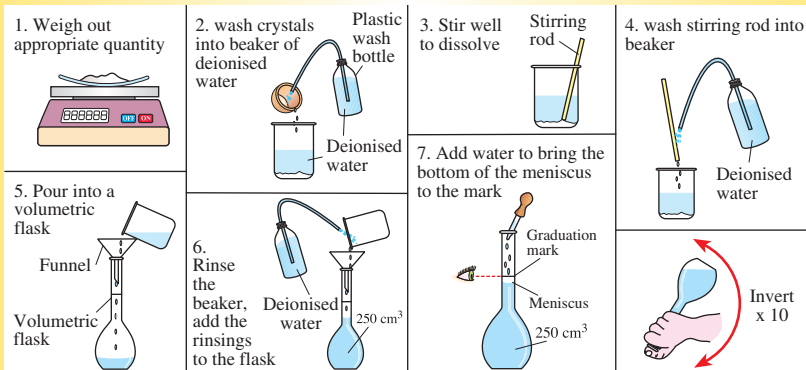
20 cm³ of 0.2 M HCl is diluted to a volume of 250 cm³ with deionised water.
What is the concentration of the diluted solution?

$$V_{\text{dil}} \times M_{\text{dil}} = V_{\text{conc.}} \times M_{\text{conc.}}$$

$$M_{\text{dil}} = \frac{V_{\text{conc.}} \times M_{\text{conc.}}}{V_{\text{dil}}} = \frac{20 \times 0.2}{250} = 0.016 \text{ M}$$

EXPERIMENT: Preparation of a Standard Solution of Sodium Carbonate [250 cm^3 of $0.1\text{M Na}_2\text{CO}_3$]

- Weigh out 7.15g of $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ on clock glass
[$250/1000 \times 0.1 \times 286 \text{ (RMM } \text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}) = 7.15$]
volume/1000 \times molarity \times RMM including water of crystallisation
- Dissolve in deionised water in beaker with washings using deionised water from clock glass
- Pour through funnel into 250 cm^3 volumetric flask with washings from the beaker using deionised water
- Make up to mark with deionised water
- Read from the bottom of meniscus and with eye level with mark
- Stopper and invert 10 times to make solution homogeneous [long thin neck makes this necessary]



4.2 Acids and Bases

Acids

Turn blue litmus red, React with metals forming a salt plus hydrogen,
Monobasic acids have one hydrogen that can be replaced by a metal. E.g. HCl ,
 HNO_3 and CH_3COOH

Dibasic acids have two hydrogens that can be replaced by a metal. E.g. H_2SO_4

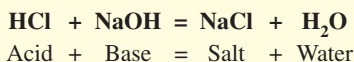
Bases

Turn red litmus blue, Neutralise acids. **Alkalis** are bases that are soluble in water

Salt

The **result of the neutralisation of an acid with a base.**

E.g. the salt sodium chloride is formed when HCl reacts with NaOH



Sulphates are the salts of sulphuric acid
Nitrates are the salts of nitric acid
Chlorides are the salts of hydrochloric acid

Neutralisation

Formation of a salt from an acid and a base.

Household acids

Vinegar (ethanoic acid), Lemon juice (citric acid)

Household bases

Sodium hydroxide as oven cleaner, Ammonia is used in household cleaners

Everyday examples of neutralisation

- Lime (CaO) is spread on fields in agriculture to neutralise acid soils
- Stomach powders contain a base such as magnesium hydroxide to neutralise excess stomach acid (HCl) which can cause indigestion.

VOLUMETRIC ANALYSIS

f 4.3 Arrhenius Theory of Acids and Bases

An Acid

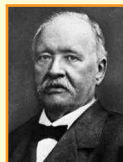
is a neutral molecule which dissociates in water to form a hydrogen ion(s) and an anion $HX = H^+ + X^-$

Acidic properties due to H^+

A Base

dissociates in water to form a hydroxide ion and a cation
 $MOH = M^+ + OH^-$

Basic properties due to OH^-



Arrhenius

Strong Acids and Bases Dissociate Fully in aqueous solution

Weak Acids and Bases only Dissociate Slightly in aqueous solution

Neutralisation of Strong Acid with a Strong Base e.g. HCl with NaOH

Acid dissociates to form H^+ and Cl^- while base dissociates to form Na^+ and OH^-

H^+ ions react with OH^- ions to form water

Na^+ and Cl^- do not react and thus can be left out of the equation

Half equation or ionic equation is



Conductivity of Strong and Weak Acids

Weak acids conduct poorly when dissolved in water because there are **few ions** present

Strong acids conduct well when dissolved in water as there are **numerous ions** present

Strong acids only conduct when dissolved in water because no ions are present in pure acid.

Limitations of Arrhenius Theory

Arrhenius Theory is **limited to dilute aqueous solutions** and has been superseded

g 4.4 Brønsted-Lowry Theory of Acids and Bases

Acid is a proton donor
Base proton acceptor

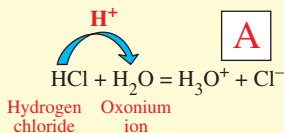
An acid-base reaction involves the transfer of a proton (H^+) from the acid to the base

Amphoteric Nature of Water

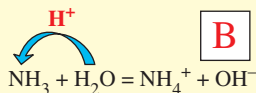
In equation A the water is acting as a base because it accepts a proton

In equation B the water is acting as an acid as it donates a proton

Water can thus **act as an acid or a base** depending on circumstances. Water is **amphoteric**



A



B

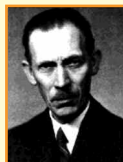
Strength of Acids and Bases

Stronger acid more readily donates a proton

Weaker acid less readily donates a proton

Stronger base more readily accepts a proton

Weaker base less readily accepts a proton



Bronsted



Lowry

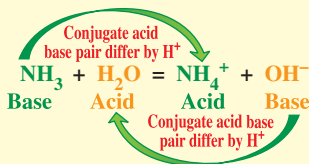
Conjugate Acid-Base Pairs

Species that differ from each other by H^+

NH_3 accepts a proton to become NH_4^+ and so is a base

NH_4^+ donates a proton to become NH_3 and so is an acid

NH_4^+ and NH_3 differ by H^+ and thus are a conjugate acid-base pair

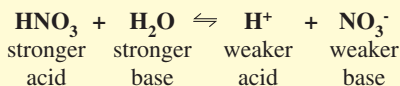


H_2O donates a proton to become OH^- and so is an acid

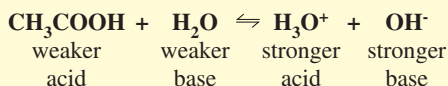
OH^- accepts a proton to become H_2O and so is a base

H_2O and OH^- differ by H^+ and thus are a conjugate acid-base pair

Reactions Lead Preferentially to the Production of the Weaker Acid and Base



the above reaction goes mainly to the right



the above reaction goes mainly to the left

The stronger an acid the weaker its conjugate base

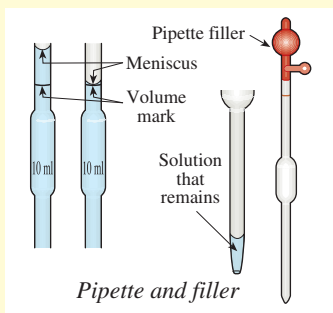
The stronger a base the weaker its conjugate acid

4.5 Volumetric Analysis

Correct Titrimetric Procedure

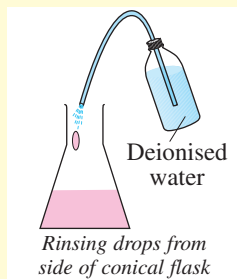
Pipette

- **Rinse with deionised water** - to wash out any impurities
- Then with the **solution it is going to contain** - to wash out the deionised water.
- **Fill using a pipette filler** - the solution may be poisonous or caustic.
- **Read from the bottom of meniscus**
 - It should be **level with the ring on the stem**
 - And your eye should also be level with this ring.
- **Empty into the conical flask and touch tip against the side of the flask.**
- **DON'T BLOW** it is calibrated to allow for the drop in the tip.



Conical flask

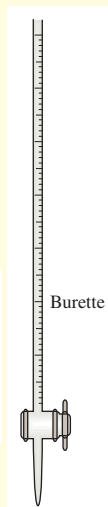
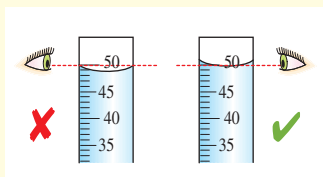
- Rinse out with deionised water only.
- Place on white tile - to see colour change more easily.
- Mix continuously during the titration.
- Add only a few drops of indicator (They are weak acids or bases and may upset the results)
- Wash down any drops on the side of the flask with deionised water. (This won't affect amount of reactant in flask or change the result.)



VOLUMETRIC ANALYSIS

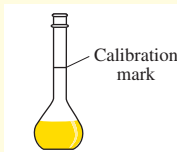
j Burette

- Rinse with deionised water
- Then with the **solution it is going to contain**
- Fill using a funnel then
- Remove funnel (as drops may fall from it or it may dip into the liquid giving a false level).
- Remove the air bubble from the tip by opening the tap quickly
- Read from the bottom of meniscus - with eye level with this point. (KMnO_4 - read from the top of the meniscus)
- Don't put NaOH in burette it may react with glass of burette or block tap [not really valid now]



Volumetric Flask

- Long thin neck makes it accurate.
- Make up to mark with deionised water – use a dropper near end point
- Read from bottom of meniscus at eye level.
- Get bottom of meniscus level with calibration mark.
- Make sure it is at room temperature - it is calibrated at 20°C .
- Mix by inverting 10 times to ensure solution is homogeneous - (long thin neck makes this necessary).



Titration

- Use the **correct indicator** [SAWBMO see bottom of page]
- Only **3 - 4 drops of indicator** [indicators are weak acids or bases so may affect result if too much added]
- Mix well
- Add from burette drop by drop near the end point.
- End Point - Point at which reaction is complete - (shown by colour change)
- Do one rough and 2 accurate titres
- Two accurate should be within 0.1 cm^3
- Average the 2 accurate titres
- Identify the standard solution [one given the concentration of] - for calculations to follow.

k Indicator

Substance that changes colour with pH or at the end-point of a titration

Indicator	Initial Colour	Final Colour
Phenolphthalein	Colourless in acid	Pink in alkali
Methyl orange	Red / pink in acid	Yellow in alkali
Litmus	Red in acid	Blue in alkali
Starch	Blue when iodine $[\text{I}_2]$ present	Colourless when iodine absent
MnO_4^{1-}	Pink when present	Colourless when absent [all turned to Mn^{2+}]
Eriochrome Black	Wine red when Ca^{2+} present	Blue when Ca^{2+} gone

Indicator Choice

Remember which indicator using these “Eskimo” words. The second and third ones are the most important

“Eskimo” word	Acid/base combination	Indicator to use
SASBANY	Strong Acid Strong Base	ANY
SAWBMO	Strong Acid Weak Base	Methyl Orange
WASBPH	Weak Acid Strong Base	Phenolphthalein
WAWBNONE	Weak Acid Weak Base	NONE

Strong / Weak Acids and Bases

Class	Definition	Examples
Strong acids	Dissociate [splits up] fully in aqueous solution	HCl, HNO ₃ , H ₂ SO ₄
Weak acids	Only dissociate partly in aqueous solution	CH ₃ COOH, HCN
Strong bases	Dissociate [splits up] fully in aqueous solution	NaOH, KOH
Weak bases	Only dissociate partly in aqueous solution	NH ₄ OH

Solutions

Standard Solution is one whose **concentration is known accurately**

Standardise means to find the concentration of a solution using titration

Primary Standard is a **pure compound** with a **high molecular mass** from which **solutions of known concentration can be made**. A **Primary Standard Solution** is **Pure**, **100% Soluble** and **Stable** once made up.

The following can't be used as Primary Standards

- MnO₄⁻ and sodium thiosulphate because *they can't be got pure*.
- Iodine [I₂] because *it sublimes*.
- KOH or NaOH because *they absorb CO₂ and moisture from the atmosphere*

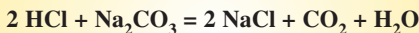
Secondary Standard

Make up a solution and then standardise this solution using a primary standard.

This secondary standard can then be used to standardise other solutions e.g. KMnO₄

EXPERIMENT: Hydrochloric Acid Standardisation using 0.1 M Sodium Carbonate

- Pipette 25 cm³ dil. HCl into a 250cm³ volumetric flask
- Make up to the mark with deionised water (dilutes acid to a factor of 10)
- Mix well to make homogeneous
- Pipette 20 cm³ of diluted HCl into a conical flask
- Add a few drops of **methly orange** indicator [SAWBMO] it turns **pink**
- Add standard Na₂CO₃ solution [0.1 M] from burette
- Mix constantly
- Wash any drops from the side of flask with deionised water [doesn't affect amount of acid in flask]
- Note volume on burette as indicator changes colour (**red/pink** to **orange/yellow**)
- Do one rough and two accurate titres and take the average of two accurate



- Let **a** = values for acid and **b** the values for the base
- Assume average titre = 19.5 cm³ for calculation in this example.

$$\text{Use } \frac{V_a M_a}{n_a} = \frac{V_b M_b}{n_b}$$

$$V_a = 20 \text{ cm}^3 \text{ [pipette volume]}$$

$$M_a = ? \text{ [Molarity of acid]}$$

$$n_a = 2 \text{ [moles of acid in balanced equation]}$$

$$V_b = 19.5 \text{ cm}^3 \text{ [average titre]}$$

$$M_b = 0.1 \text{ [Molarity of base i.e. standard solution]}$$

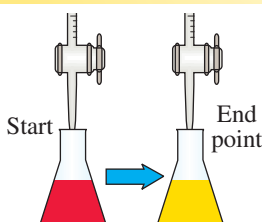
$$n_b = 1 \text{ [moles of base in balanced equation]}$$

$$M_a = \frac{V_b M_b n_a}{V_a n_b}$$

$$M_a = \frac{19.5 \times 0.1 \times 2}{20 \times 1} = 0.195$$

Multiply by 10 for dilution

Answer = **1.95 M**



VOLUMETRIC ANALYSIS

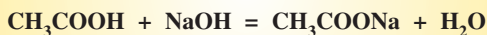
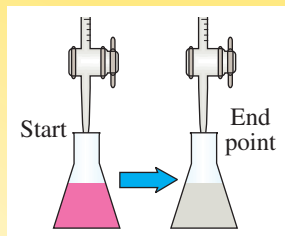
EXPERIMENT: Making a Salt (Sodium Chloride)

- Pipette 25 cm³ of 1 M NaOH into a conical flask
- Add a few drops of **phenolphthalein indicator** [SASBANY] – it should turn **pink**
- Add 1 M HCl from a burette and note the volume at which the indicator goes **colourless** i.e. when the NaOH has been completely neutralised.
- Repeat this till two values agree to within 0.1 cm³ and note this volume.
- This is the volume of HCl needed to completely neutralise the NaOH.
- Pipette 25 cm³ of fresh 1 M NaOH into a clean beaker
- Do not add any indicator**
- Add the amount of HCl need to completely neutralise the NaOH
- Gently heat the solution until all the water has evaporated.
- There will be a sample of sodium chloride in the bottom of the beaker

n

EXPERIMENT: Find the Concentration of Ethanoic acid in Vinegar as % (w/v)

- Pipette 20 cm³ of 0.1 M NaOH into conical flask,
- Add a few drops of **phenolphthalein indicator** [WASBPH], turns pink.
- Dilute vinegar by factor of 10 [25 cm³ into a 250 cm³ volumetric flask]
- Rinse the burette with deionised water and then with ethanoic acid
- Put diluted ethanoic acid (vinegar) in burette
- Remove the air bubble by opening the tap quickly.
- Add ethanoic acid from burette to the NaOH in flask – mixing constantly
- Wash any drops from side with deionised water [*doesn't affect amount acid in flask*]
- When **pink turns colourless** note volume (adding drop by drop near end).
- Do one rough and two accurate titres and take the average of the two accurate
- Let a = values for the acid [ethanoic acid] and b for base [sodium hydroxide]
- Assume average titre = 24 cm³ for calculation in this example.



$$V_a = 24 \text{ cm}^3 \text{ [average titre]}$$

$$M_a = ? \text{ [molarity of acid]}$$

$$n_a = 1 \text{ [moles of base in balanced equation]}$$

$$V_b = 20 \text{ cm}^3 \text{ [pipette volume]}$$

$$M_b = 0.1 \text{ [Molarity of base – standard solution]}$$

$$n_b = 1 \text{ [moles of base in balanced equation]}$$

$$\frac{V_a M_a}{n_a} = \frac{V_b M_b}{n_b}$$



$$M_a = \frac{V_b M_b n_a}{V_a n_b} = \frac{20 \times 0.1 \times 1}{24 \times 1} = 0.083$$

Multiply by 10 for dilution

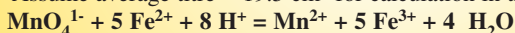
Answer = **0.83 M**

Find the % (W/V) concentration of the vinegar.

$$0.83 \times 60 = 49.99 \text{ g per litre} \\ = 4.99 \text{ g per 100 ml} = \mathbf{4.99\% \text{ (W/V)}}$$

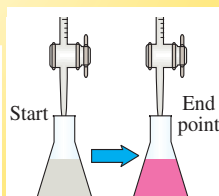
EXPERIMENT: Standardise Potassium Manganate (VII) using Ammonium Iron(II) Sulphate

- Use 0.1 M Ammonium iron (II) sulphate as **standard solution**
- Acidified using H_2SO_4 to prevent air oxidising Fe^{2+} to Fe^{3+}**
- $KMnO_4$ in burette – **read from top of meniscus** – dark colour
- Pipette 20 cm^3 ammonium iron (II) sulphate into conical flask
- Add $10\text{ ml cm}^3 H_2SO_4$ to flask to make sure that Mn (VII) is fully reduced to Mn^{2+} . Brown precipitate of MnO_2 formed if insufficient acid used.**
- Pink colour is slow to disappear for first few drops then goes quickly because Mn^{2+} is an autocatalyst i.e. a product of reaction that catalyses the reaction.
- MnO_4^- acts as its own indicator. End point when **permanent pink colour**
- MnO_4^- is oxidising agent and Fe^{2+} is reducing agent
- Assume average titre = 19.5 cm^3 for calculation in this example.



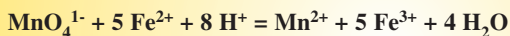
Calculation – use: $\frac{V_o M_o}{n_o} = \frac{V_r M_r}{n_r} \Rightarrow M_o = \frac{V_r \times M_r \times n_o}{V_o \times n_r}$

$$M_o = \frac{20 \times 0.12 \times 1}{19.5 \times 5} = 0.02\text{ M}$$

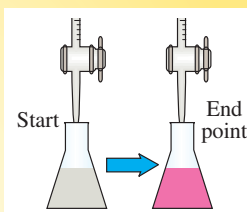


EXPERIMENT: Find the amount of Iron in an Iron Tablet and hence the Percentage of Iron in the Tablet.

- Find the mass of the five iron tablets– [to give average]. Let us say 1.8 g
- Crush tablets with pestle and mortar in dilute H_2SO_4
- Easier to dissolve and **stops atmospheric oxygen oxidising Fe^{2+} to Fe^{3+}**
- Transfer the paste with washings to beaker and stir to dissolve the paste.
- Using a funnel transfer solution into 250 cm^3 volumetric flask with washings
- Make up to mark with deionised water. Eye and meniscus both level with mark.
- Invert 10 times to make solution homogeneous.
- Pipette 20 cm^3 of tablet solution into conical flask.
- Add 20 cm^3 of dil H_2SO_4 acid** to ensure full reaction
- Titrate with $0.02\text{ M } KMnO_4$ from burette until a **permanent pink colour**
- MnO_4^{1-} is oxidising agent and Fe^{2+} is reducing agent



- Mn^{2+} is an **autocatalyst** (product of reaction that catalyses reaction)
- Do one rough and two accurate titres. Average the 2 accurate
- Assume average titre = 17.5 cm^3 for calculation in this example



Calculate the Molarity of Iron Sulphate solution

use: $\frac{V_o M_o}{n_o} = \frac{V_r M_r}{n_r} \Rightarrow M_r = \frac{V_o \times M_o \times n_r}{V_r \times n_o}$

$$M_r = \frac{17.5 \times 0.02 \times 5}{20.0 \times 1} = 0.0875\text{ M}$$

Calculate Moles of Fe in 250 cm^3 of Solution

$$\frac{0.0875 \times 250}{1000} = 0.021875\text{ moles}$$

Calculate Mass of Fe in 250 cm^3 of Solution

$$\begin{aligned} \text{Mass} &= \text{molarity} \times \text{molar mass} \\ &= 0.021875 \times 56 = 1.225\text{ g} \end{aligned}$$

Calculate Mass of Fe in each Tablet

$$1.225 / 5 = 0.245\text{ g}$$

Calculate the Mass of each Tablet

$$\text{Mass of 5 tablets} = 1.8\text{ g}$$

$$\text{Mass of 1 tablet} = 1.8 / 5 = 0.36\text{ g}$$

Calculate % Fe in each Tablet

$$\% \text{ iron} = \frac{\text{mass of Fe}}{\text{Mass of 1 tablet}} \times 100$$

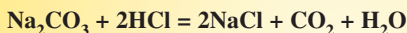
$$= \frac{0.245}{0.36} \times 100 = 68.06\%$$

VOLUMETRIC ANALYSIS

q

EXPERIMENT: To find the Molar Mass of Hydrated Sodium Carbonate $[\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}]$, the Number of Molecules of Water of Crystallisation and the Percentage of Water.

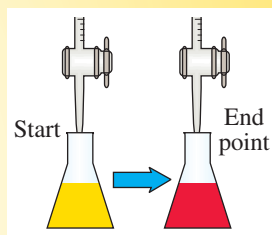
- weigh out 3.15g washing soda crystals $[\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}]$
- dissolve in 100cm^3 deionised water, stir
- transfer to 250cm^3 volumetric flask with washings,
- make solution up to mark – bottom of meniscus and eye level with the mark
- invert 20 times to make homogeneous
- pipette 25cm^3 of Na_2CO_3 solution into conical flask
- add 3 drops of **methyl orange** [SAWBMO] indicator – it goes **yellow**
- fill burette with standard 0.15 M HCl solution
- titrate: colour change from **yellow to red/pink**
- do one rough and 2 accurate – average the 2 accurate titres
- Assume average titre = 15.0 cm^3 for calculation in this example.



Calculate the Molarity of the Sodium Carbonate

use: $\frac{V_a M_a}{n_a} = \frac{V_b M_b}{n_b} \Rightarrow M_b = \frac{V_a \times M_a \times n_b}{V_b \times n_a}$

$$M_b = \frac{15 \times 0.15 \times 1}{25 \times 2} = 0.045 \text{ M}$$



Find the Molar Mass of $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$

Concentration of sodium carbonate in mol l^{-1} = 0.045

Concentration of sodium carbonate in $\text{g l}^{-1} = 3.15 \times 4 = 12.6 \text{ g l}^{-1}$ [because we had 250 cm^3]

Mass of 0.045 moles of $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$ = 12.6 g

Molar mass of $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O} = \frac{\text{mass}}{\text{No. of moles}} = \frac{12.6}{0.045} = 280$

Calculate the value of x in the formula

Molar mass of $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O} = 46 + 12 + 48 + 18x = 106 + 18x$ [where x is the mass of water]

$106 + 18x = 280$ [molar mass]

$\Rightarrow 18x = 280 - 106 = 174$ [total mass of water in crystal in mass units]

$\Rightarrow x = \frac{174}{18} = 9.7 \Rightarrow \text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$

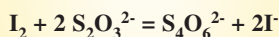
Calculate the percentage of water in the crystals

Molar mass of hydrated sodium carbonate = 280

Percentage of water of crystallisation in the compound = $\frac{174}{280} \times 100 = 62.14\%$

EXPERIMENT: Calculate the concentration of a Thiosulphate Solution Using 0.06M Iodine Standard Solution

- Neither sodium thiosulphate nor iodine are suitable as primary standards
 - For each titration make up **0.06M iodine (M_o)** solution as follows
 - Pipette **25cm^3 (V_o)** of **0.02 M potassium iodate** standard solution into the conical flask
 - Add **20cm^3 of dilute H_2SO_4** and **10cm^3 of 0.5M KI** solution [excess] (*KI solution keeps the iodine in solution*)
 - Note the **reddish/brown** colour of liberated **iodine**.
 - Do this for each titration
- Weigh 6.25g of sodium thiosulphate crystals onto a clock glass
- Transfer crystals to beaker containing 100cm^3 of de-ionised water
- Stir and when dissolved transfer to 250cm^3 volumetric flask with washings
- Make up to mark with de-ionised water. Invert 20 times.
- Add the sodium thiosulphate from the burette**
- When solution becomes **straw coloured** add a few drops of **starch indicator**
 - if added too early it will complex with the iodine and ruin the experiment*
- Blue colour** seen.
- Continue adding the thiosulphate solution until the solution in flask becomes **colourless** – *when all the I_2 is gone*.
- Do one rough and two accurate titres. Average the two accurate titres.
- Assume average titre = 30 cm^3 for calculation in this example.
- The iodine reacts with the sodium thiosulphate according to the equation



- Iodine is the oxidising agent and thiosulphate is the reducing agent

Iodine

$$V_o = 25\text{cm}^3$$

$$M_o = 0.06\text{M}$$

$$n_o = 1$$

Thiosulphate

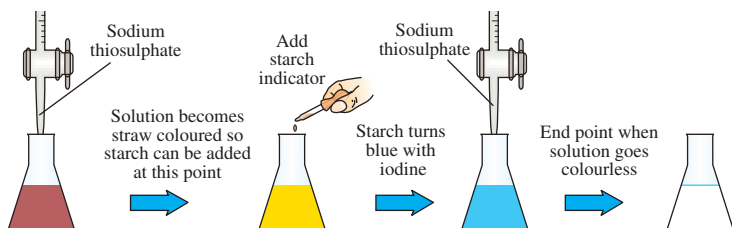
$$V_r = 30\text{cm}^3$$

$$M_r = ?$$

$$n_r = 2$$

$$\frac{V_o \times M_o}{n_o} = \frac{V_r \times M_r}{n_r} \Rightarrow M_r = \frac{V_o \times M_o \times n_r}{V_r \times n_o}$$

$$\Rightarrow M_r = \frac{25 \times 0.06 \times 2}{30 \times 1} = \frac{3}{30} = 0.1\text{ M}$$

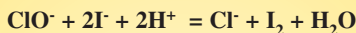


VOLUMETRIC ANALYSIS

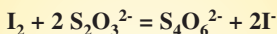
S

EXPERIMENT: Determine the % (w/v) of Sodium Hypochlorite in Bleach using 0.1 M Sodium Thiosulphate

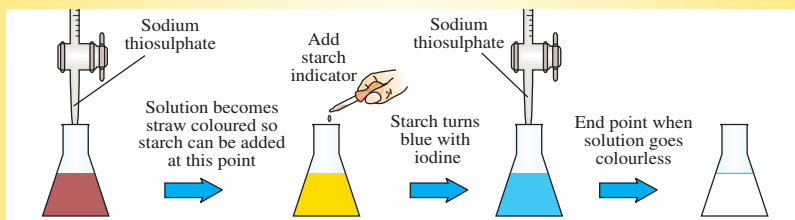
- Pipette **25 cm³ concentrated bleach into 250 cm³ volumetric flask**
- Fill up to the mark with deionised water, **invert it 20 times**.
- Diluted by factor 10** because conc. bleach is dangerous
- Pipette 20 cm³ diluted bleach into conical flask
- Then add 20 cm³ sulphuric acid and 10 cm³ 0.5M potassium iodide (KI) solution.
- Note **reddish/brown colour of iodine**



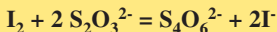
- Each hypochlorite molecule produces one iodine molecule
- Add 0.1 M sodium thiosulphate solution from the burette until the solution is a **pale yellow** (straw) colour
- Add a few drops of **starch indicator**, note **blue colour**
- Add thiosulphate drop by drop until colour changes from **blue to colourless**.
- Do one rough and two accurate titres. Take average of the two accurate titres
- Assume average titre = 14.8 cm³ for calculation in this example.



- Iodine is the reducing agent and thiosulphate is the oxidising agent



Calculate molarity of iodine and thus sodium hypochlorite (NaOCl)



$\text{ClO}^- = \text{I}_2$ [eqn 1] and $\text{I}_2 = 2\text{S}_2\text{O}_3^{2-}$ [eqn 2] therefore $1\text{ClO}^- = 2\text{S}_2\text{O}_3^{2-}$

[ClO^- is the oxidising agent and $\text{S}_2\text{O}_3^{2-}$ is the reducing agent]

$$\frac{V_o M_o}{n_o} = \frac{V_r M_r}{n_r} \Rightarrow M_o = \frac{V_r \times M_r \times n_o}{V_o \times n_r}$$

$$M_o = \frac{14.8 \times 0.1 \times 1}{20 \times 2} = \frac{1.48}{40} = 0.037 \text{ M}$$

Multiply by 10 for dilution

Concentration of hypochlorite in bleach = $0.037 \times 10 = 0.37 \text{ M}$

Calculate mass of hypochlorite in 1 litre

concentration = molarity \times molar mass of sodium hypochlorite

$$= 0.37 \times 74.5 = 27.57 \text{ g l}^{-1} [\text{NaOCl} = 23 + 16 + 35.5 = 74.5]$$

Convert to % (w/v)

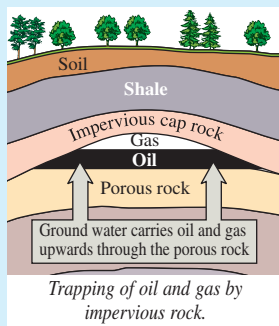
$$\text{Percentage (w/v) of hypochlorite in bleach} = \frac{27.57}{1000} \times 100 = 2.76\%$$

Questions on this section from Past Exams Year by Year

2009	2008	2007	2006	2005	2004	2003	2002
1	1	1	1	1	1	1	1
						10 b	

a 5.1 Sources of Hydrocarbons

- Coal, natural gas and petroleum are our main sources of hydrocarbons.
- Coal is the fossilised remains of swamp plants while oil is the fossilised remains of marine plankton.
- Natural gas [methane – CH₄]** is formed with both of these – it is called fire damp in coal mines
- The decomposition of animal waste and vegetation produces methane in slurry pits, coal mines and refuse dumps. This source can be tapped and used as a fuel. Otherwise it can collect and pose the potential **hazard of an explosion or fires or suffocation** – it is not poisonous.
- Methane** produced by the above and also by ruminants such as cows and sheep contributes significantly to the enhanced greenhouse effect.



b 5.2 Aliphatic Hydrocarbons

Aliphatic

Consists of Chains of Carbon Atoms

Hydrocarbon

Made of Carbon and Hydrogen Atoms Only

Homologous Series – Families of organic chemicals are called Homologous Series

They have the following properties

- Same Functional Group**
- Same General Formula**
- Each differs from its neighbours by CH₂**
- Same method of preparation**
- [similar physical and chemical properties]**

Homologous Series and their Properties

Name	Functional group	First member of Series	Solubility in water [polar]	Solubility in cyclohexane [non-polar]
Alkanes	C – C	Methane CH ₄ [C _n H _{2n+2}]	Insoluble	Soluble
Alkenes	C = C	Ethene C ₂ H ₄ [C _n H _{2n}]	Insoluble	Soluble
Alkynes	C ≡ C	Ethyne C ₂ H ₂ [C _n H _{2n-2}]	Insoluble	Soluble
Alcohols	R – OH	Methanol CH ₃ OH [C _n H _{2n+1} OH]	Short Soluble	Short Insoluble
Aldehyde	R – CHO	Methanal HCHO	Short Soluble	Short Insoluble
Ketones	R – CO – R'	Propanone CH ₃ COCH ₃	Short Soluble	Short Insoluble
Carboxylic Acids	R – COOH	Methanoic acid HCOOH	Short Soluble	Short Insoluble
Esters	R – COO – R'	Methylmethanoate HCOOCH ₃	Short Soluble	Short Insoluble

c 5.3 Alkanes

General Formula C_nH_{2n+2}

Saturated

Has only single carbon to carbon covalent bonds

You need to know the formulae and structures of the alkanes up to C₈ [octane] and the isomers up to C₅

You also need to know the formulae and structures of cyclohexane and 2,2,4-trimethylpentane

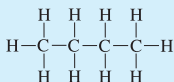
# Carbons	Name	Structure
1	methane	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$
2	ethane	$\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array}$
3	propane	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \end{array}$
4	butane	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$
5	pentane	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$
6	hexane	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$
7	heptane	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$
8	octane	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$

FUELS AND HEATS OF REACTION

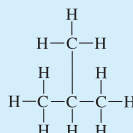
Isomers

Same Molecular Formula but Different Structural Formulae

Isomers of Butane

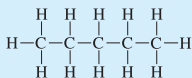


Butane

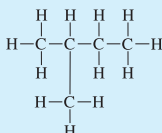


2-methylpropane

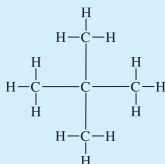
Isomers of Pentane



Pentane

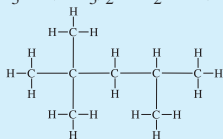


2-methylbutane

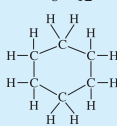


2,2-dimethylpropane

Other Structural Formulae



2,2,4-trimethylpentane



Cyclohexane

5.4 Alkenes

General formula C_nH_{2n}

You need to know the formulae and structures up to C_4

Unsaturated

They contain a Double carbon to carbon Bond

Planar around the double bond

Tests for Unsaturation

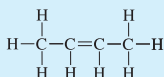
Turn acidified MnO_4^- from pink to colourless

Decolourise red/brown bromine (water)

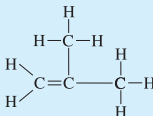
Isomers differ by

- Position of double bond
- Branching or Position of methyl group

You need to know the structures of but-2-ene and 2-methylpropene



But-2-ene

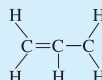


2-methylpropene

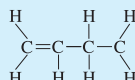
Ethene



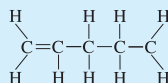
Propane



Butane



Pentene



State

Alkenes are gases up to butene [C_4] under normal conditions.

Longer chain alkenes [$\text{C}_5 - \text{C}_{15}$] are liquids and above this [C_{16+}] are solids.

Solubility

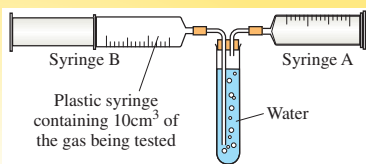
Insoluble in water because they are non-polar and water is polar.

Soluble in non-polar solvents such as benzene – like dissolves like

FUELS AND HEATS OF REACTION

EXPERIMENT: Solubility Properties of Methane, Ethene and Ethyne

- Fill a test tube about 75% with water and stopper as shown
- Empty one syringe and connect it to tube that is above water
- Fill the other tube with the gas of your choice and connect it to the tube that goes below the water surface
- Slowly pass the gas through the test tube to the other syringe and note the new volume
- Calculate the percentage of the gas that has dissolved



$$\% \text{ dissolved} = \frac{\text{Initial volume} \times 100}{\text{Final volume}}$$

- Repeat with the other gases
- Replace the water with cyclohexane and repeat the above

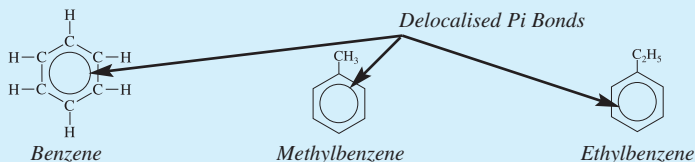
Solvent	Methane	Ethene	Ethyne
Water	insoluble	insoluble	insoluble
Cyclohexane	40%	35%	20%

5.5 Aromatic Hydrocarbons – Benzene

Aromatic

has a benzene ring in its structure

Benzene is a liquid under room conditions



Benzene is the basis of dyestuffs, detergents, herbicides, many pharmaceuticals and indicators such as methyl orange and phenolphthalein

Benzene and many of its derivatives are carcinogenic (can cause cancer) e.g. benzene in petrol – but not all are carcinogenic e.g. aspirin.

Benzene does not decolourise bromine showing that it does not have double bonds

Methylbenzene is used as an industrial **solvent** for non-polar compounds – it is **not** carcinogenic

- Insoluble in water** because it is **non-polar** and water is polar.
- Soluble in cyclohexane** because both are non-polar.
- Does not react like a normal saturated compound** [does not decolourise bromine] because of the **delocalised Pi bonds** – shown by ring in centre of molecule.

EXPERIMENT: THE Solubility Properties of Methylbenzene

- Put some methylbenzene in test tube
- Add small quantity of solute – stopper and shake
- Open carefully in case there is a pressure build-up
- Examine to see if solute has dissolved
- Non-polar solutes dissolve in methylbenzene but polar solutes do not

Solute	Solute Polarity	Result
Water	Polar	No solution
Iodine	Non-polar	dissolves
Sodium chloride	Polar	No solution
Cooking oil	Non-polar	dissolves

FUELS AND HEATS OF REACTION

h 5.6 Thermochemistry

Endothermic and Exothermic Reactions

Most chemical reaction result in a change in temperature.

Endothermic Reactions

Reactions which **use heat** and **feel cold** to the touch are called **endothermic**.

They have a **positive ΔH value** e.g. $\Delta H = +123 \text{ kJ}$ [ΔH means heat change]

Changes of state are also accompanied by heat changes.

Liquid turning to gas, and solid to liquid, both use heat and are **endothermic**.

sublimation (solid to gas directly) is also endothermic. **CO_2 and I_2 both sublime.**

When sherbet or ammonium chloride is dissolved in water the temperature drops so these are examples of endothermic reactions.

Exothermic Reactions

Reactions which **produce heat** and **feel hot** to the touch are called **exothermic**.

They have a **negative ΔH value** e.g. $\Delta H = -123 \text{ kJ}$

Gas to liquid, liquid to solid and desublimation release heat and are exothermic

The combustion of alkanes and other hydrocarbons is exothermic and produces heat, CO_2 and H_2O . e.g. natural gas in a bunsen burner.

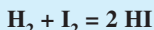
These **hydrocarbons are our main source of energy**



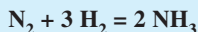
Combustion can be explosive especially in short chain members)

i Heat of Reaction [ΔH_r]

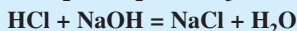
The heat change in kJ when the numbers of moles stated in the balance equation react



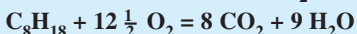
$$\Delta H = + 52.1 \text{ kJ}$$



$$\Delta H = - 92 \text{ kJ}$$



$$\Delta H = - 57 \text{ kJ}$$



$$\Delta H = - 5512 \text{ kJ}$$

j Bond Energy

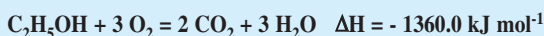
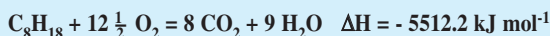
The average amount of energy in kJ needed to break one mole of bonds of the same type, all species being in the gaseous state.

Breaking the four bonds in a mole of methane to form carbon and hydrogen atoms takes 1648 kJ. What is the energy of C-H bond in methane?

There are 4 C-H bonds in methane and all are the same so it is $1648 / 4 = 412 \text{ kJ mol}^{-1}$

k Heat of Combustion [ΔH_c]

Heat change (in kJ mol^{-1}) when one mole of a substance is completely burned in excess oxygen



Bomb Calorimeter

Instrument used to **measure Heats of Combustion accurately**. Also used to find the calorific value of foods.

Different fuels have different Heats of Combustion.

$$\text{Methane} = - 880.4 \text{ kJ mol}^{-1}$$

$$\text{Ethanol} = - 1371.0 \text{ kJ mol}^{-1}$$

