Molar Ratio

Molar Ratios

Balanced equations can be used to show how the **quantities of reactants** relate to the **quantities of products** via **molar ratios**.

Example 1—Balanced Equation Neutralisation reaction

$2HCl(aq)$	+	$Na_2CO_3(aq)$	→	$2NaCl(aq)$	+	$CO_2(g)$	+	$H_2O(\ell)$
2 mole		1 mole		2 mole		1 mole		1 mole

Example 2—Balanced Equation Neutralisation reaction

$H_2SO_4(aq)$	+	$Na_2CO_3(aq)$	→	$Na_2SO_4(aq)$	+	$CO_2(g)$	+	$H_2O(\ell)$
1 mole		1 mole		1 mole		1 mole		1 mole

N5 Calculation Revision

When given the **mass of one reactant/product**, a second reactant/product's mass can be calculated by working out the G.F.M of both substances & using the mole ratio.

Example 1

If **4g** of **hydrogen** is used in a reaction, calculate the **mass of water** produced.

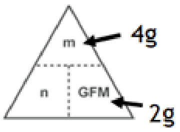
$$\boxed{2H_2} \quad + \quad O_2 \longrightarrow \boxed{2H_2O}$$

Worked Example—3 steps

Step 1: Establish the mole: mole ratio between the reactant H_2 and the product H_2O

H_2		**H_2O**
2 mole	⟶	**2 mole**

Step 2: Work out how many moles of reactant & product present



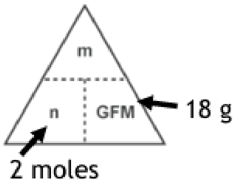
mass → mole

moles H_2 = mass/GFM
= 4/ 2
= 2 moles ⟶ 2 moles

Step 3: Convert moles of water to mass using the G.F.M

mole → mass

mass = moles x GFM
= 2 x 18
= <u>36g</u>



Molar Volume

Molar volume (l mol^{-1})

1 mole of any gas occupies the SAME volume (24L) at standard temperature and pressure.

Molar volume = **<u>24 l mol^{-1}</u>** (standard temperature & pressure)

The molar volume value is NOT 24 l mol $^{-1}$ at different temperatures/ pressures

Molar Volume Balanced Equation Calculations

Given a volume of a gaseous substance in litres & using the molar ratio from the balanced equation, the volume of a second gaseous substance from the equation can be calculated using molar volume.

Example

When **400 cm^3** of nitrogen reacts with excess hydrogen, calculate the mass of ammonia that will be produced under standard temperature & pressure.

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

Remember to
change to litres
(400/1000 = **0.4L**)

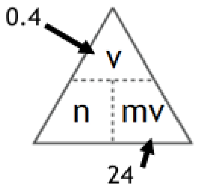
Worked Example—3 steps

Step 1: Establish the mole: mole ratio between the reactant N_2 and the product NH_3

$$\begin{array}{ccc} \mathbf{N_2} & & \mathbf{NH_3} \\ \textbf{1 mole} & \longrightarrow & \textbf{2 mole} \end{array}$$

Step 2: Work out how many moles of reactant & product present

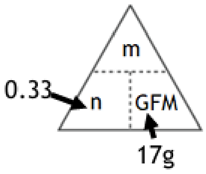
volume → moles



Mole = v/mv
= 0.4/24
= 0.0166666667 moles → 2 x 0.01666666667 = **0.0333333 moles**

Step 3: Convert moles of ammonia to mass using GFM of NH_3

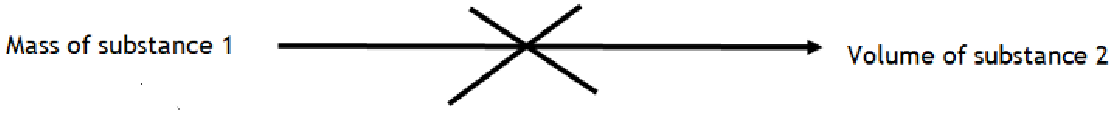
Volume = moles x mV
= 0.033333 x 17
= 5.7g of NH_3



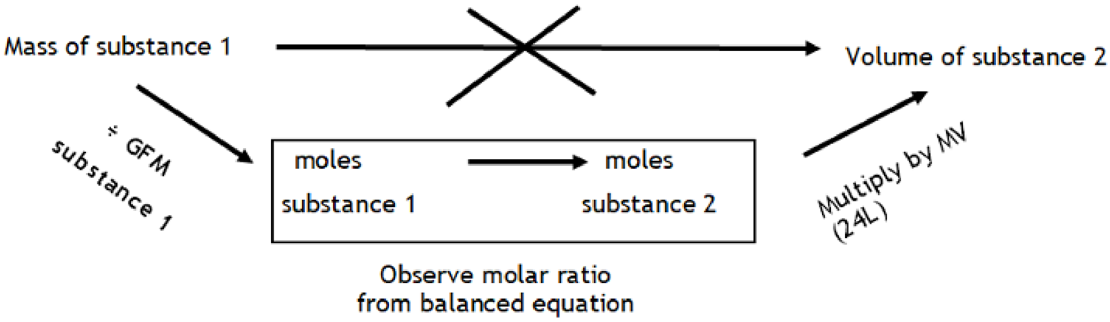
Volume to Mass/Mass to volume from equations

Calculating volume of product from mass of reactants

It is not possible to directly link mass of a reactant to the volume of a product as they are not proportional to one another. NO magic triangle.



However changing the **mass of one substance into moles using the G.F.M** and using the mole ratio from a balanced equation allows the volume of the second substance to be calculated using the molar volume.



Example

Calculate the volume of carbon dioxide released when 0·5 g of calcium carbonate reacts with excess hydrochloric acid.

$$\mathbf{CaCO_3(s)} + 2HCl(aq) \qquad CaCl_2(aq) + H_2O(l) + \mathbf{CO_2(g)}$$

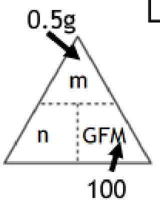
Step 1: Establish the mole: mole ratio between the limiting reactant and the desired product

$$\mathbf{CaCO_3(s)} \longrightarrow \mathbf{CO_2(g)}$$

1 mole **1 mole**

Step 2: Work out how many moles of limiting reactant present

mass → moles

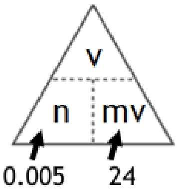


moles = mass/GFM

= 0.5 / 100

= <u>**0.005 moles $CaCO_3$**</u> ⟶ **0.005 moles CO_2**

Step 3: Convert moles of gas produced to volume using the molar volume



mole → volume

volume = mole x mv

= 0.005 moles x 24

= **0.12l CO_2**

Calculating volume of products from mass of reactants

It is not possible to directly link mass of a reactant to the volume of a product as they are not proportional to one another substance.

Method

1. Compare the **mole: mole ratio in the balanced equation of the 2 substances**
2. Change the **mass into moles** of substance 1 given in the question by **dividing** by its GFM
3. Convert **moles into volume** of substance 2 by **multiplying** by the molar volume (24 l)

Example

Calculate the volume of carbon dioxide released when 0·5 g of calcium carbonate reacts with excess hydrochloric acid.

$$\mathbf{CaCO_3(s)} + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + \mathbf{CO_2(g)}$$

Step 1: Establish the mole: mole ratio between the limiting reactant and the desired product

$$\mathbf{CaCO_3(s)} \longrightarrow \mathbf{CO_2(g)}$$

1 mole **1 mole**

Step 2: Work out how many moles of limiting reactant present

Mass —> moles
—> 0.5 / 100
—> 0.005 moles ⟶ 0.005 moles CO_2

Step 3: Convert moles of gas produced to volume using the molar volume

0.005 moles x 24 = **0.12 litres CO_2**

Excess Calculations

Limiting Reactant

Is completely used up in the reaction and will determine the quantity of the product produced.

Excess Reactant

Is not used up completely and does not have any impact on the quantity of the product produced.

In order to ensure that a **costly reactant** is converted into the desired product, an excess of the less expensive reactant(s) can be used.

Using a balanced equation, the limiting reactant and the reactant(s) in excess can be identified by calculation.

Method

1. Calculate the number of moles of each reactant present
2. Use the balanced equation to determine the mole to mole ratio of the reactants.
3. Choose 1 reactant, assume it all reacts determine whether you have enough of the other

Yes you have more than enough = excess reactant

No not enough = limiting reactant

Example

9·72 g of magnesium was reacted with 200 cm^3 of 2 mol l^{-1} hydrochloric acid.

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

By considering which reactant was the limiting reactant, determine the mass of magnesium chloride produced.

Step 1:Calculate the number of moles of each reactant present.

Mg 9.72/24.3 = 0.4 moles HCl - 0.2 x 2 - 0.4 moles

Step 2: Use the equation to determine the mole: mole ratio between the reactants

Mg : HCl

1 : 2

Step 3: Choose one of the reactants, assuming that it all reacts, determine whether or not you have enough of the other reactant to allow this to happen.

Mg : HCl

1 : 2

0.4 : 0.8 moles of HCl required to react all the Mg. There are only 0.4 moles of HCl available therefore not enough to react all the Mg.

Theoretical Yield

Theoretical Yield

The quantity of desired product obtained, assuming **full conversion of the limiting reagent** as calculated from the balanced equation.

Actual Yield

The quantity of the desired product formed under the prevailing reaction conditions.

Percentage Yield

For a particular set of reaction conditions, the percentage yield provides a measure of the degree to which the limiting reagent is converted into the desired product.

Using a balanced equation, calculations involving percentage yield can be performed from data provided.

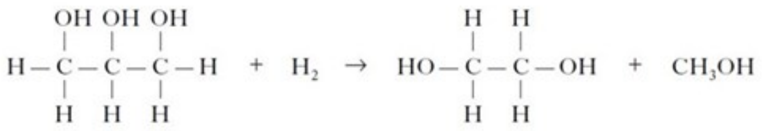
Actual yield % yield = x 100
Theoretical yield

Method

1. Determine from the data, the actual yield limiting factor
2. Calculate the theoretical yield
3. Using percentage yield formula to calculate percentage yield.

Example

Ethane-1,2-diol is produced in industry by reacting glycerol with hydrogen.



$$\mathrm{H-\overset{OH}{\underset{H}{C}}-\overset{OH}{\underset{H}{C}}-\overset{OH}{\underset{H}{C}}-H + H_2 \rightarrow HO-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-OH + CH_3OH}$$

glycerol ethane-1,2–diol

Excess hydrogen reacts with 27.6kg of glycerol to produce 13.4kg of ethane-1,2-diol.

Calculate the **percentage yield** of ethane-1,2-diol.

Step 1: Determine from the data the actual yield, limiting factor.

Actual yield: 13.4 kg ethan-1,2-diol Limiting factor: 27.6 kg glycerol

Step 2: Calculate the theoretical yield

1 mole glycerol —> 1 mole ethan-1,2-diol

Mass —> moles

27600 —> 27600/ 92

—> 300 moles —> 300 moles ethan-1,2-diol —> mass

—> 300 x 62= 18600g —> 18.6 kg

Step 3: Calculate % yield

$$\%\ \text{yield} = \frac{13.4}{18.6} \times 100$$

$$= 72\ \%$$

Atom Economy

Atom Economy

The proportion of the **total mass of all starting materials** converted into the desired product in the balanced equation.

The percentage atom economy can be calculated using the equation

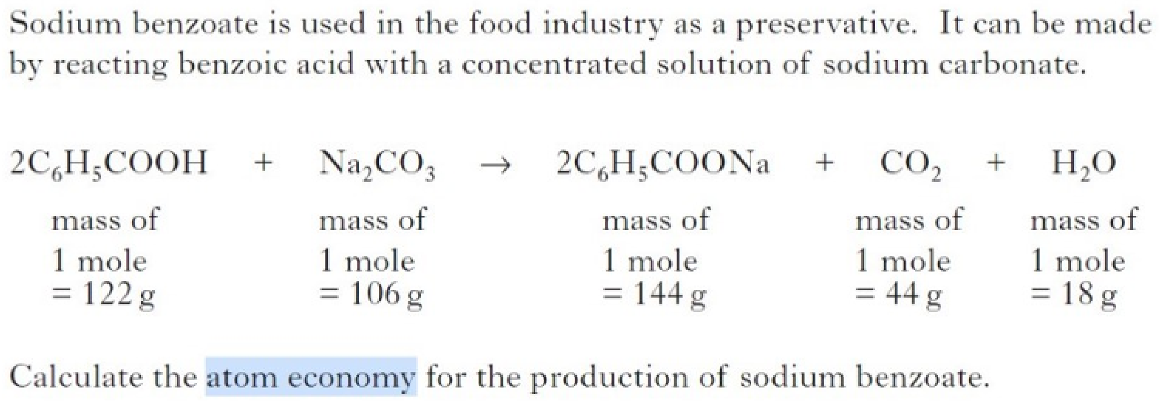
$$\%\ \text{atom economy} = \frac{\text{Mass of desired product}}{\text{Total mass of reactants}}$$

Reactions with a high percentage yield may have a low atom economy value if large quantities of by-products are formed.

Method

1. Determine from the data, the actual yield limiting factor
2. Calculate the theoretical yield
3. Using percentage yield formula to calculate percentage yield.

Example



Sodium benzoate is used in the food industry as a preservative. It can be made by reacting benzoic acid with a concentrated solution of sodium carbonate.

$$2C_6H_5COOH + Na_2CO_3 \rightarrow 2C_6H_5COONa + CO_2 + H_2O$$

$2C_6H_5COOH$	Na_2CO_3	$2C_6H_5COONa$	CO_2	H_2O
mass of 1 mole = 122 g	mass of 1 mole = 106 g	mass of 1 mole = 144 g	mass of 1 mole = 44 g	mass of 1 mole = 18 g

Calculate the atom economy for the production of sodium benzoate.

$$\text{Atom economy} = \frac{2 \times 144}{(2 \times 122) + 106} \times 100$$

$$= \mathbf{82.26\%}$$

Note!!! You must take account of the moles present in the balanced chemical equation and include this in the mass calculations for each species.

Rates of Reaction

Rates of Reactions

The rate of all chemical reactions is determined by the collision theory:

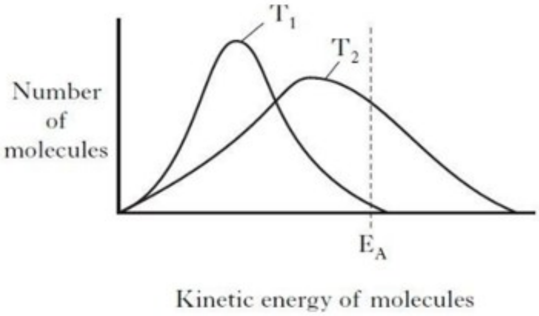
1. Particles must collide with energy **greater** than the **activation energy**
2. Particles must also collide at the **correct orientation.**

Activation energy is defined as the **minimum energy** required for successful collisions to be possible.

The following factors will have an impact on Successful Collisions:

1. Temperature

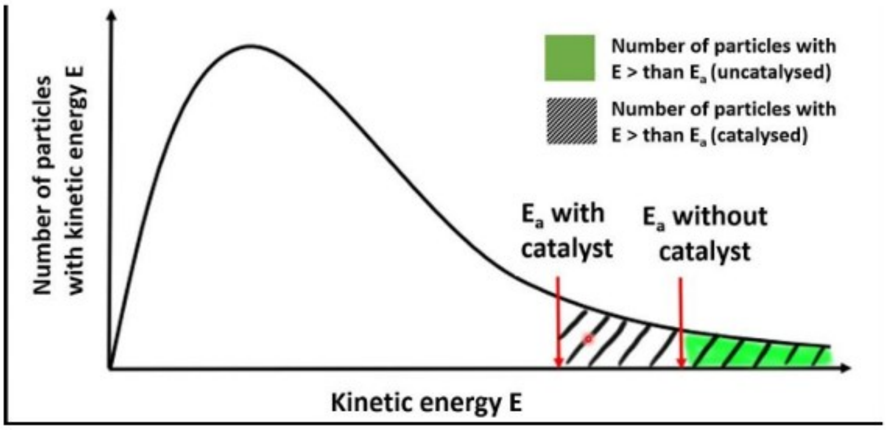
Temperature is a measure of the average kinetic energy of particles.
A small rise in temperature is enough to double the rate of a reaction.



At T_1 the area under the curve beyond the activation energy is small. Increasing the Temperature T_2 increases the area under the curve beyond the activation energy. This means that there is an increased number of particles with sufficient energy to successfully collide.

2. Catalyst

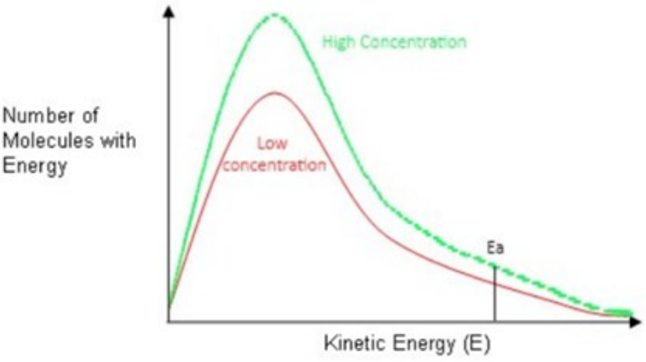
A catalyst speeds up the rate of a reaction by lowering the activation energy. This means that more particles will have energy beyond the new activation energy so the likelihood of successful collisions increases.



Rates of Reaction

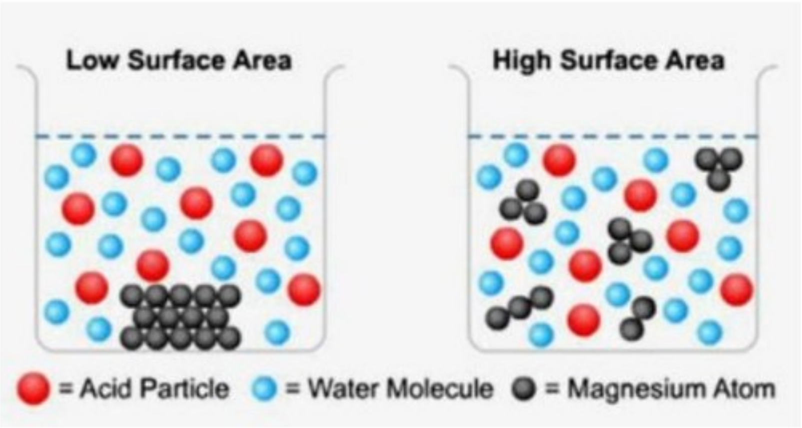
3. Concentration

The concentration of particles can affect the rate of a reaction by increasing the number of particles across the range of kinetic energies, therefore there will be a greater likelihood of successful collisions as more particles have energies beyond the activation energy.



4. Particle Size

Decreasing the particle size will provide a larger surface area thus increasing the opportunities for successful collisions



Reaction rates must be controlled in industrial processes. If the rate is too low then the process will not be economically viable; if it is too high there will be a risk of explosion. The rate of reactions can be determines experimentally to produce data to calculate the relative rate.

$$\text{Relative rate} = \frac{1}{\text{time}} \qquad \text{Units : } s^{-1}$$

$$\text{Time} = \frac{1}{\text{Rate}}$$

Enthalpy Change: Potential Energy Diagram

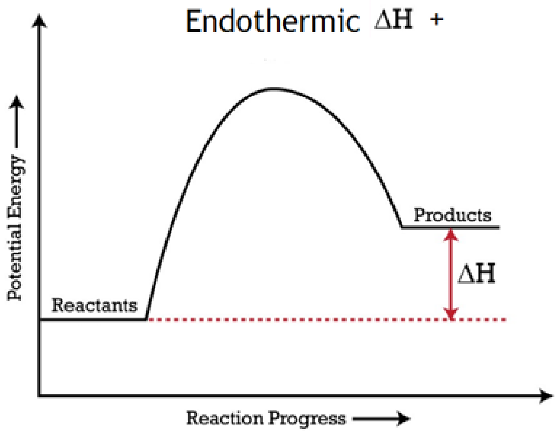
Enthalpy change (ΔH)

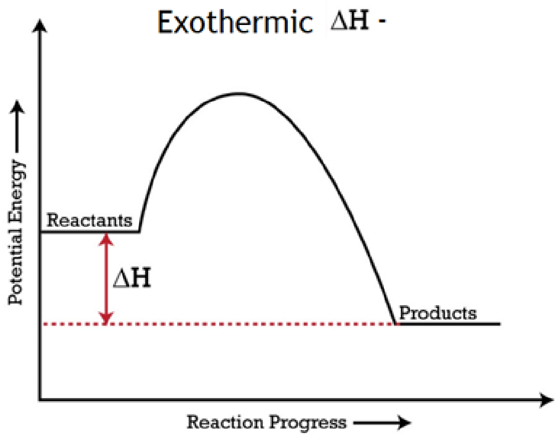
The energy difference between the enthalpy of the products and the reactants.

$$\Delta H = H_P - H_R$$

Potential Energy Diagrams

A potential energy diagram can be used to show the energy pathway for a reaction.





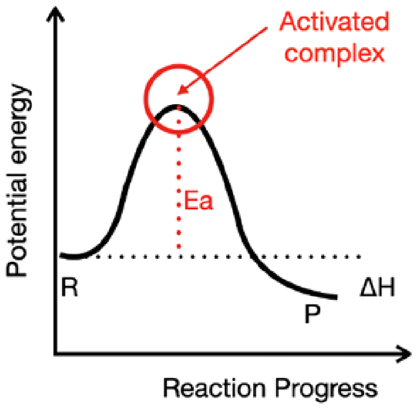
Activated complex

An unstable arrangement of atoms formed at the maximum of the potential energy barrier, during a reaction.

Activation energy (E_a)

Minimum energy required by colliding particles to form an activated complex.

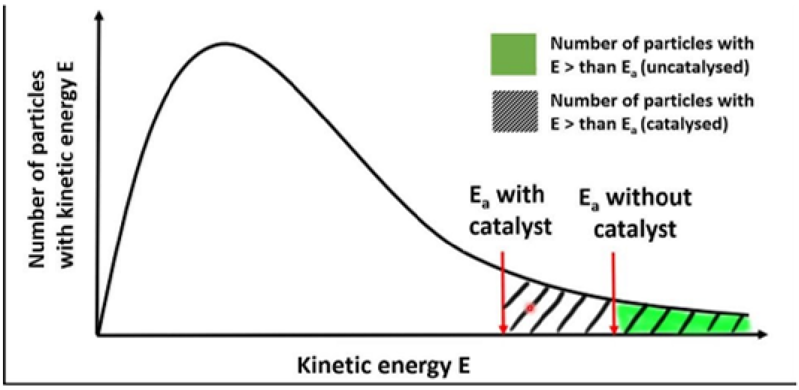
Lower activation energies require a smaller input of energy and are more likely to take place at room temperature.



Enthalpy Change: Potential Energy Diagram

Catalyst

A catalyst speeds up the rate of a reaction by lowering the activation energy. This means that more particles will have energy beyond the new activation energy so the likelihood of successful collisions increases.

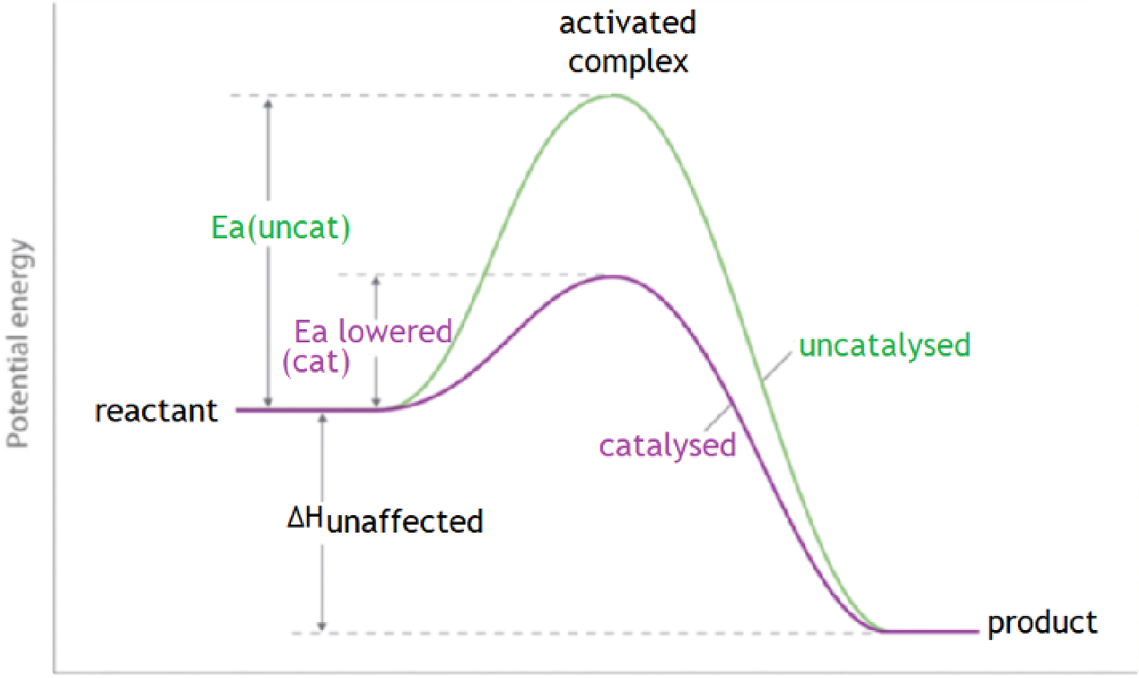


Potential Energy Diagram Catalysts

A catalyst will

1. **Lower E_a** (activation energy) by providing an alternative pathway for the reaction.

2. **No effect** on ΔH (enthalpy change).



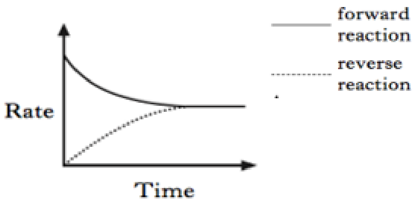
Equilibrium

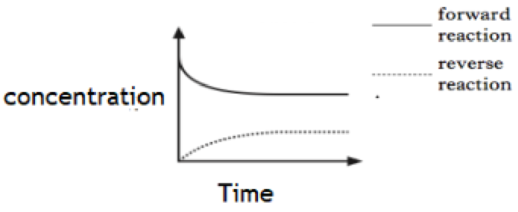
1. Rate

Rate of the **forward** reaction = **rate** of the **backward** reaction

2. Concentrations

The **concentrations** of reactants and products are **constant**





Le Chatelier's Principle:

> If a dynamic equilibrium is disturbed by changing the conditions, the position of equilibrium moves to counteract the change.

The following changes (TCP) can be made to maximise the product production.

- Temperature
- Concentration
- Pressure

1. Temperature

Increasing the temperature

Favours the endothermic reaction (ΔH = +)

Decreasing temperature

Favour the exothermic reaction (ΔH = -)

Example

$$NH_3(g) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

$$\Delta H = -36\,kJ\,mol^{-1}$$

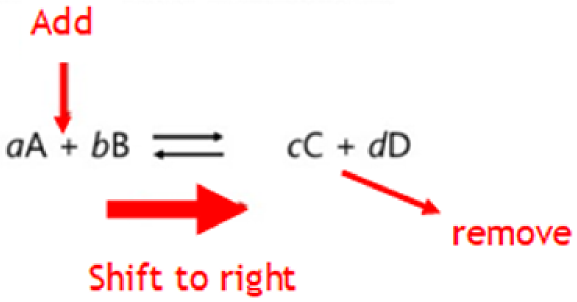
Lowering the temperature will promote the exothermic (ΔH = -) forward reaction.

Increasing the temperature will promote the endothermic (ΔH = +) backward reaction

2. Concentration

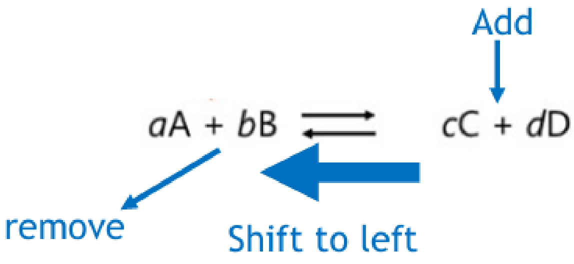
Removing the product/adding more reactant.

Shifts the equilibrium to the **right** (promotes forward reaction) thus producing more product.



Removing the reactant/adding more product .

Shifts the equilibrium to the left (promotes backward reaction) thus producing more reactant



Example $Cl_2(g) + H_2O(\ell) \rightleftharpoons 2H^+(aq) + ClO^-(aq) + Cl^-(aq)$

Forward reaction (shifts right) promoted by:

1. Increasing Cl_2 or H_2O molecules into the reaction
2. Removing H^+ /CLO^-/Cl^- ions from the reaction.

** H^+ ions can be removed by adding OH- ion into the reaction.**

Backward reaction (shifts left) promoted by:

1. Removing Cl_2 or H_2O molecules from the reaction
2. Adding H^+ /CLO^-/Cl^- ions into the reaction

Equilibrium

3. Pressure

Will only have an impact on equilibrium reactions which are carried out in a **sealed container.**

Increasing pressure

Favours the side of **least** **volume** of **gas.**

Decreasing the pressure

Favours the side of **greatest volume** of **gas.**

Same Volume each side

If both sides have the same volumes of gas they will be affected equally and the position of equilibrium will be **unaffected.**

Example 1

$$\underset{3\ \text{vol}}{2SO_2(g) + O_2(g)} \rightleftharpoons \underset{2\ \text{vol}}{2SO_3(g)}$$

Increasing the pressure
Promotes the forward reaction (shifts right) with the lower volume of gas.

Decreasing the pressure
Promotes the backward reaction (shifts left) to the side with the higher volume of gas.

Example 2

$$\underset{2\ \text{vol}}{NH_3(g) \quad + \quad O_2(g)} \quad \rightarrow \quad \underset{2\ \text{vol}}{NO(g) \quad + \quad H_2O(g)}$$

Increasing or decreasing the pressure does not affect the position of equilibrium as there is the same volume of gas on both sides.

Enthalpy

Enthalpy

The measure of the chemical energy in a substance.

1. Exothermic reaction

A reaction/process that releases heat energy.

In industry, exothermic reactions may **require heat to be removed** to prevent the temperature rising & explosions occurring.

2. Endothermic reactop

A reaction/process that takes in heat energy.

In industry, endothermic reactions may **incur costs in supplying heat energy** in order to maintain the reaction rate.

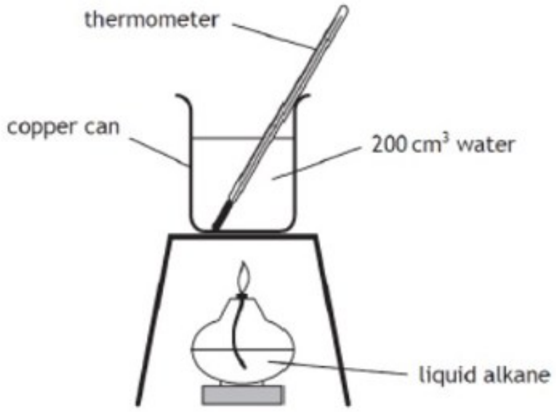
The enthalpy of combustion

The enthalpy change when one mole of the substance burns completely in oxygen.

$$\boxed{C_4H_{10}} \quad + \quad \mathbf{6.5}\ O_2 \quad \rightarrow \quad \mathbf{4}CO_2 \quad + \quad \mathbf{5}H_2O$$

The enthalpy change associated with a reaction can be calculated from the quantity of heat energy released experimentally using

$$E_h = c\ m\ \Delta T$$



$$E_h = cm\Delta T$$

Where E_h = energy gained by the water

c = specific heat capacity of water = $4.18\ kJ^{o}C^{-1}kg^{-1}$

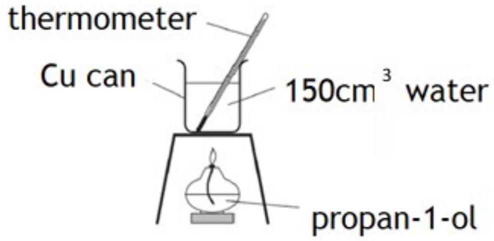
m = mass of water (1 litre = 1kg)

ΔT= change in temperature of water

Enthalpy of Combustion: Experimental Data

Example

The following data was obtained by a student trying to determine the enthalpy of combustion of propan-1-ol (C_3H_7OH).



Mass of burner + alcohol before = *184·25 g*
Mass of burner + alcohol after = *183·95 g*
Volume of water = *150 cm^3*
Temperature of water before heating = *21 °C*
Highest temperature of water after heating = *32 °C*

$E_h = c\ m\ \Delta T$
$= 4.18 \times 0.150 \times 11$
$= 6.896$ kJ

Extra step at higher

Need to get enthalpy of combustion = 1 mole

Mass of propan-1-ol burned: 0.3g

Moles of propan-1-ol burned : 0.3 / 60 = 0.005 moles —> 6.896 kJ

$1 \text{ mole} \longrightarrow \dfrac{1 \times 6.896}{0.005}$

$\longrightarrow 1378 \text{ kJ mol}^{-1}$

$\underline{\Delta H = -1378 \text{kJ mol}^{-1}}$

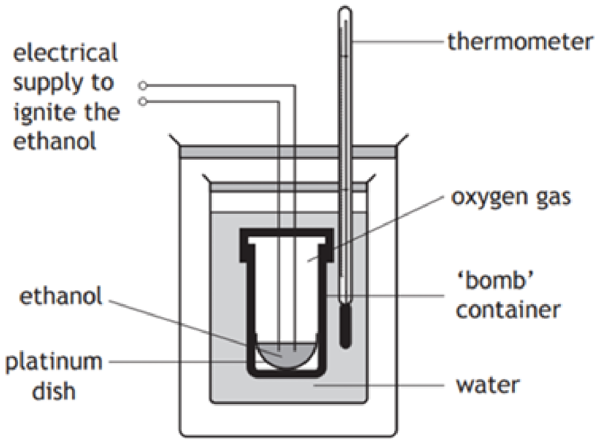
This value is lower than that quoted in the data book

1. Heat is lost to the surroundings
2. Incomplete combustion of the alcohol
3. Alcohol lost through evaporation

Bomb Calorimeter

An accurate measure of the enthalpy change can be found using a bomb calorimeter.

1. **Water completely surrounds fuel** so no heat is lost to surroundings when it burns
2. **Pure oxygen is used** to prevent incomplete combustion of fuel.
3. As **alcohol is encased in a closed vessel** no loss of alcohol via evaporation.

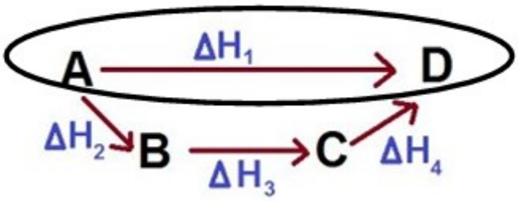


Hess's Law

Hess's Law state's that the overall enthalpy change for a reaction will be the same regardless of route taken.

Example 1

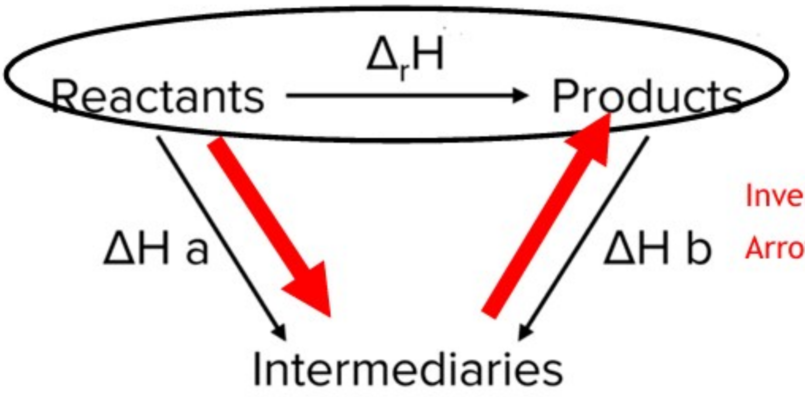
For route A to D.



$$\Delta H_1 = \Delta H_2 + \Delta H_3 + \Delta H_4$$

Example 2

For reactants to products.



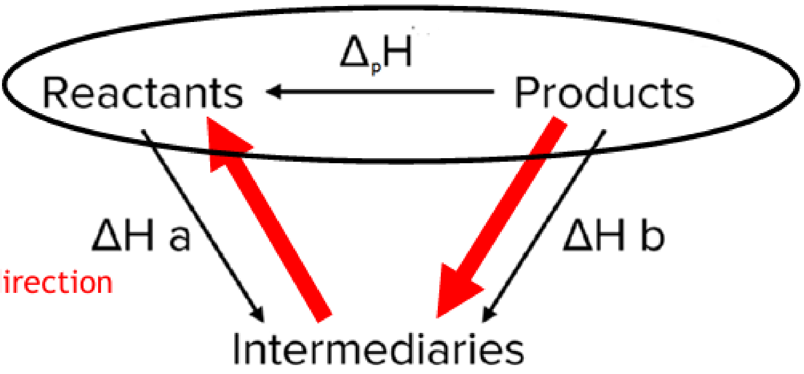
Invert b to **-b**

Arrow faces wrong direction

$$\Delta_r H = \Delta H\ a - \Delta H\ b$$

Example 3

For products to reactants.



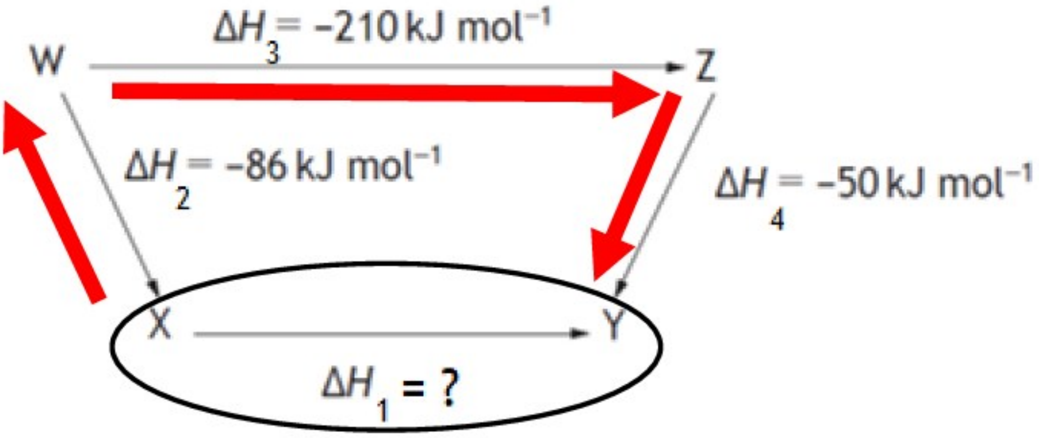
Invert a to **-a**

Arrow faces wrong direction

$$\Delta_p H = \Delta H\ b - \Delta H\ a$$

Hess's Law

Numerical Worked Example—Calculate the enthalpy change associated with **X to Y.**



$$\Delta H_1 = \Delta H_3 + \Delta H_4 - \Delta H_2$$

$= -210 + (-50) - (-86)$

$= -210 - 50 + 86$

$= -174\ \text{kJmol}^{-1}$

Example

Calculate the enthalpy change for the reaction using the following enthalpy changes:

$$C_{(s)} + Cl_{2(g)} + H_{2(g)} \longrightarrow CH_2Cl_{2(g)}$$

1. $CH_2Cl_{2(g)} + O_{2(g)} \longrightarrow CO_{2(g)} + 2HCl_{(g)}$ ΔH = -446 kJ mol-1
2. $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ ΔH = -394 kJ mol-1
3. $H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$ ΔH = -184 kJ mol-1

2. $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ ΔH = -394 kJ mol-1
3. $H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$ ΔH = -184 kJ mol-1

Rev 1. $CO_{2(g)} + 2HCl_{(g)} \longrightarrow CH_2Cl_{2(g)} + O_{2(g)}$ ΔH = 446 kJ mol-1

Turn -446 into +446 as product appears as reactant in original equation

Cancel out any species which exists on both sides of equations:

2. $C_{(s)} + \cancel{O_{2(g)}} \longrightarrow \cancel{CO_{2(g)}}$ ΔH = -394 kJ mol-1
3. $H_{2(g)} + Cl_{2(g)} \longrightarrow \cancel{2HCl_{(g)}}$ ΔH = -184 kJ mol-1
1. $\cancel{CO_{2(g)}} + \cancel{2HCl_{(g)}} \longrightarrow CH_2Cl_{2(g)} + \cancel{O_{2(g)}}$ ΔH = 446 kJ mol-1

$C_{(s)} + H_{2(g)} + Cl_{2(g)} \longrightarrow CH_2Cl_{2(g)}$ ΔH = -132 kJ mol-1

Bond Enthalpies

Bond Enthalpies

Breaking bonds in a molecule is an **endothermic process**, i.e energy must be applied.

$$\text{H—H} \longrightarrow 2\,\text{H} \qquad \Delta H = 286 \text{ kJ mol-1}$$

Making bonds in a molecule is an **exothermic process**, i.e energy is released to the surroundings.

$$2\text{H} \longrightarrow \text{H-H} \qquad \Delta H = -286 \text{ kJ mol-1}$$

Bond enthalpies are found in the data book.

Bond enthalpies: Values quoted for diatomic bonds

Mean Bond Enthalpies: Average values quoted for bonds found in a range of molecules.

The bond enthalpy gives an indication of the strength of a bond, For example,

H-F	H-Cl
570 kJ mol-1	432 kJ mol-1

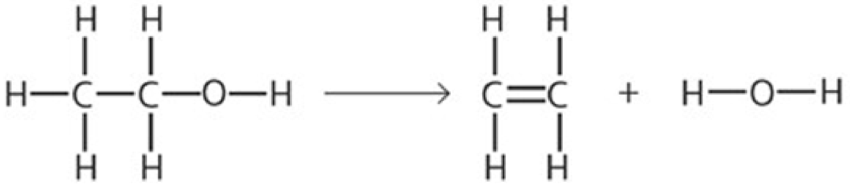
More energy would be required to break a H-F bond compared to a H-Cl bond.

Bond Enthalpies can be used to calculate the ΔH for a reaction.

Example

Calculate the enthalpy for the reaction below using bond enthalpy data from the data book.

ethanol → ethene + water



Bond Breaking (+)	Bond Making (-)
5 (C-H) = 5 (412) = 2060	4 (C-H) = 4 (412) = 1648
(C-C) = 348	(C=C) = 612
(C-O) = 360	2 (O-H) = 2 (463) = 926
(O-H) = 463	
3231	-3186

$$\Delta H = 3231 + (-3186)$$
$$= \mathbf{45\ kJ\ mol^{-1}}$$

Oxidising/Reducing Agents

Oxidation and reduction reactions play important roles in chemistry and are collectively called **redox** reactions.

Oxidation reactions	Reduction reactions
A species loses electrons	A species **gains** electrons

Reducing agents (electron donor)

- Usually a metal/ negative ion
- Donates electrons to another element or ion **reducing** the **other species**.
- Are themselves **oxidised**

Example

Sodium is a **reducing agent** which is **itself oxidised** as follows.

$$Na(s) \rightarrow Na^{+}(aq) + e^{-}$$

Strongest reducing agents

Alkali metals (Group 1) as they have **low electronegativities** and **lose electrons** very **easily.**

Oxidising agents (electron acceptor)

- Normally a non-metal/positive ion
- Gain electrons from other atoms **oxidising** the **other species.**
- Are themselves **reduced**

The electrochemical series represents a series of reduction reactions found in the Data Bookl

Example
Chlorine & copper ions are both **oxidising agents** which are themselves reduced as follows:

$$Cl_2(g) + 2e^{-} \rightarrow 2Cl^{-}(aq)$$

$$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$$

Strongest oxidising agents

Halogens (group 7) which are **highly electronegative** and **gain** electrons **easily.**

Oxidising Agent Uses

1. Turn alcohols into aldehydes/ketone & then carboxylic acids in the lab.
2. Can kill fungi and bacteria & inactivate viruses.
2. Used to break down coloured compounds in hair/clothes (hydrogen peroxide H_2O_2)

Oxidising/Reducing Agents

Reaction		
$Li^+(aq) + e^-$	$\rightleftharpoons$	$Li(s)$
$Cs^+(aq) + e^-$	$\rightleftharpoons$	$Cs(s)$
$Rb^+(aq) + e^-$	$\rightleftharpoons$	$Rb(s)$
$K^+(aq) + e^-$	$\rightleftharpoons$	$K(s)$
$Sr^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Sr(s)$
$Ca^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Ca(s)$
$Na^+(aq) + e^-$	$\rightleftharpoons$	$Na(s)$
$Mg^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Mg(s)$
$Al^{3+}(aq) + 3e^-$	$\rightleftharpoons$	$Al(s)$
$2H_2O(l) + 2e^-$	$\rightleftharpoons$	$H_2(g) + 2OH^-(aq)$
$Zn^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Zn(s)$
$Cr^{3+}(aq) + 3e^-$	$\rightleftharpoons$	$Cr(s)$
$Fe^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Fe(s)$
$Ni^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Ni(s)$
$Sn^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Sn(s)$
$Pb^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Pb(s)$
$Fe^{3+}(aq) + 3e^-$	$\rightleftharpoons$	$Fe(s)$
$2H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2(g)$
$Sn^{4+}(aq) + 2e^-$	$\rightleftharpoons$	$Sn^{2+}(aq)$
$Cu^{2+}(aq) + e^-$	$\rightleftharpoons$	$Cu^+(aq)$
$SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$	$\rightleftharpoons$	$SO_3^{2-}(aq) + H_2O(l)$
$Cu^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Cu(s)$
$O_2(g) + 2H_2O(l) + 4e^-$	$\rightleftharpoons$	$4OH^-(aq)$
$I_2(s) + 2e^-$	$\rightleftharpoons$	$2I^-(aq)$
$Fe^{3+}(aq) + e^-$	$\rightleftharpoons$	$Fe^{2+}(aq)$
$Ag^+(aq) + e^-$	$\rightleftharpoons$	$Ag(s)$
$Hg^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Hg(l)$
$Br_2(l) + 2e^-$	$\rightleftharpoons$	$2Br^-(aq)$
$O_2(g) + 4H^+(aq) + 4e^-$	$\rightleftharpoons$	$2H_2O(l)$
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^-$	$\rightleftharpoons$	$2Cr^{3+}(aq) + 7H_2O(l)$
$Cl_2(g) + 2e^-$	$\rightleftharpoons$	$2Cl^-(aq)$
$MnO_4^-(aq) + 8H^+(aq) + 5e^-$	$\rightleftharpoons$	$Mn^{2+}(aq) + 4H_2O(l)$
$F_2(g) + 2e^-$	$\rightleftharpoons$	$2F^-(aq)$

Redox Equations

Equations for redox reactions can be produced by adding together the two ion-electron equations representing each half-step either reduction or oxidation.

Oxidation Ion Electron Equation

Magnesium metal is oxidised (loses electrons) to form magnesium ions. The ion-electron equation for the oxidation step is:

$$Mg(s) \rightarrow Mg^{2+} + 2e^-$$

Reduction Ion Electron Equation

The reduction reaction involves copper ions in the solution being reduced (gaining electrons) to form copper metal, and is shown by the following ion-electron equation:

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

Redox Equation

Adding the two half equations so that the **electrons cancel out** gives the equation for the redox reaction.

$Mg(s) \rightarrow Mg^{2+}(aq) + \cancel{2e^-}$ Oxidation half step

$Cu^{2+}(aq) + \cancel{2e^-} \rightarrow Cu(s)$ Reduction half step

---

$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + CU(s)$ Redox equation

Magnesium is the reducing agent (oxidised) Copper is the oxidising agent (reduced).

Redox Equations

Balancing electrons in Redox Equations

Oxidation Ion Electron Equation

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$

Reduction Ion Electron Equation

$$Ag^+(aq) + e^- \rightarrow Ag(s)$$

Redox Equation

Imbalance in the **number of electrons**

This means that the ion-electron equation involving silver (the reduction step) must be **multiplied by two** before the equations are added together.

The **electrons** must also be **cancelled out** before writing the final redox reaction.

$$Zn(s) \rightarrow Zn^{2+}(aq) + \cancel{2e^-}$$ oxidation

$$2\,Ag^+(aq) + \cancel{2e^-} \rightarrow 2\,Ag(s)$$ reduction x 2

$$Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$$ redox

Zinc is the reducing agent (oxidised) Silver is the oxidising agent (reduced).

Writing Complex ion-electron equations

1. Balance the atoms that are already present on both sides.
2. If oxygen atoms are not balanced, add the same number of water molecules to the side with the lower total oxygen.
3. Add twice the number of hydrogen (H^+) ions as water molecules to the OTHER side (side with higher original oxygen atoms).
4. Complete the equation by adding the total number of electrons needed to balance the overall electrical charge.

Redox Equations

What is the ion-electron equation for the reduction of dichromate ions to chromium(III) ions?

The basic change involved from the question can be written.

$$Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq)$$

1. Firstly, we can balance the chromium ions on both sides.

$$\underline{Cr}_2O_7^{2-}(aq) \rightarrow \boxed{2}Cr^{3+}(aq)$$

2. As there is oxygen present on the left hand side, we can balance this by adding water molecules to the right hand side.

$$Cr_2\underline{O_7}^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + \boxed{7H_2O(l)}$$

The addition of water has added hydrogen into the equation, which must also be balanced.

This can be achieved by adding hydrogen ions to the left hand side.

$$Cr_2O_7^{2-}(aq) + \boxed{14H^+}(aq) \rightarrow 2Cr^{3+}(aq) + \underline{7H_2}O(l)$$

Finally, the electric charge must be balanced on both sides. Totalling the charge on the left hand side gives an overall charge of 12+ (the total of the charges on the negative dichromate ion and positive hydrogen ions), while the right hand side is 6+ (from the two chromium ions).

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

$14^+ + 2^- = \underline{12^+}$ $\quad$ $2 \times 3^+ = \underline{6^+}$

The charge can be balanced by adding electrons to the left hand side to give a balanced complex ion-electron equation:

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) \boxed{+ 6e^-} \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

$\underline{6^+}$ $\quad$ $\underline{6^+}$

Redox Titration

Redox Titrations

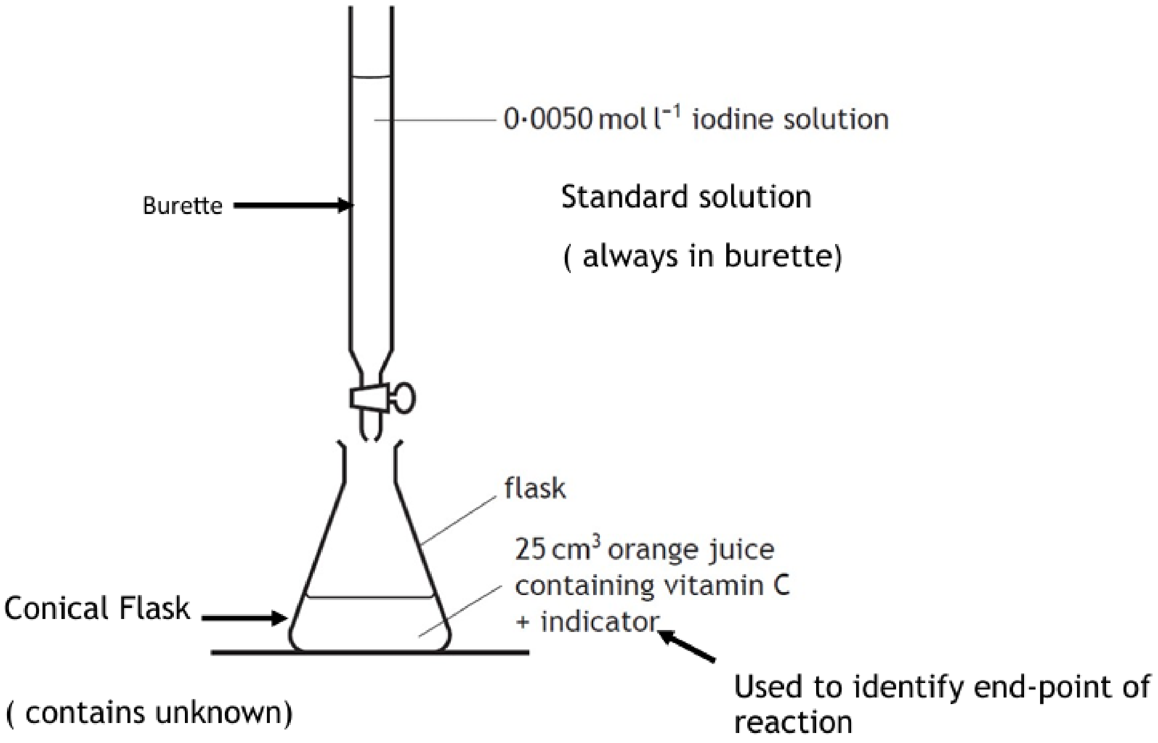
Redox equations can be used in titration reactions to determine the concentration of a solution.

In all redox titrations a **Standard solution** is used.

Standard Solution: A solution of **accurately** known concentration.

Vitamin C is found in fruits and vegetables.

Using iodine solution, a student carried out titrations to determine the concentration of vitamin C in orange juice.



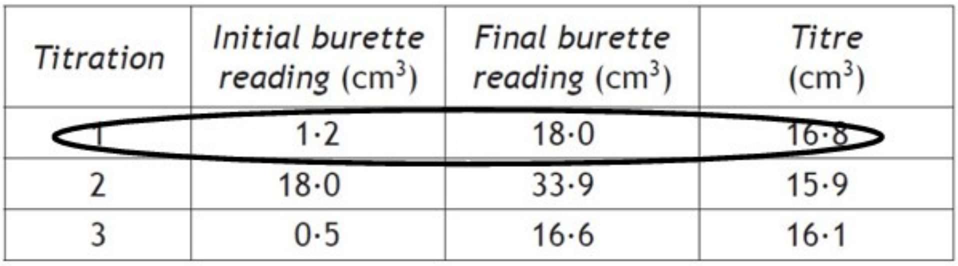
During the titration, the standard solution is added to the conical flask until there is a colour change observed. This is described as the end-point of the reaction.

The results of the titration are given in the table.

Titration	*Initial burette reading* (cm^3)	*Final burette reading* (cm^3)	*Titre* (cm^3)
1	1·2	18·0	16·8
2	18·0	33·9	15·9
3	0·5	16·6	16·1

Redox Titration

The results of the titration are given in the table.



Titration	*Initial burette reading* (cm^3)	*Final burette reading* (cm^3)	*Titre* (cm^3)
1	1·2	18·0	16·8
2	18·0	33·9	15·9
3	0·5	16·6	16·1

The average volume of iodine required to react with the Vit C is calculated using the 2 results which are concordant(within 0.2 cm^3)

The first set of results are inaccurate and considered rough results.

Average Volume required = $\frac{15.9 + 16.1}{2}$ = 16.0 cm^3

The equation for the reaction is

$$\underset{\text{vitamin C}}{C_6H_8O_6(aq)} + I_2(aq) \longrightarrow C_6H_6O_6(aq) + 2HI(aq)$$

Calculate the concentration, in mol l^{-1}, of vitamin C in the orange juice.

Always identify in the equation which reactant is the **standard solution** and which is the **unknown**.

$$\overset{\text{Unknown}}{\underset{\text{vitamin C}}{C_6H_8O_6(aq)}} + \overset{\text{Standard solution}}{I_2(aq)} \longrightarrow C_6H_6O_6(aq) + 2HI(aq)$$

The question should always provide the data to calculate the number of moles of the standard solution, I_2 reacting:

n = c x V

= 0.005 x 0.016

= 0.00008 moles of I_2 reacting with Vit C

As the mole: mole ratio is 1: 1

Moles of Vit C must be 0.00008 moles

Concentration of Vit C in the orange juice = Moles/ Volume

= 0.00008/ 0.025

= **0.0032 moll^{-1}**

Chromatography

Chromatography is a technique used to separate the components present within a mixture.

Chromatography separates substances by differences in their

1. **molecular size.**
2. **polarity**

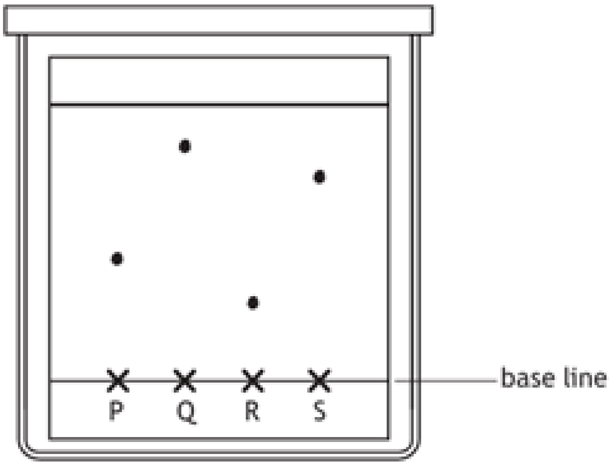
There are different types of chromatography.

1. Paper and thin-layer chromatography

   allows components to be identified by how far they have travelled in a solvent.

   This is referred to as the R_f value.

$$R_f = \frac{\text{distance moved by the substance}}{\text{maximum distance moved by the solvent}}$$



Smaller molecules will move further than larger molecules.

If the solvent is non-polar then the non-polar components will travel further up the chromatogram than polar components.

Chromatography

2. Gas-Liquid Chromatography

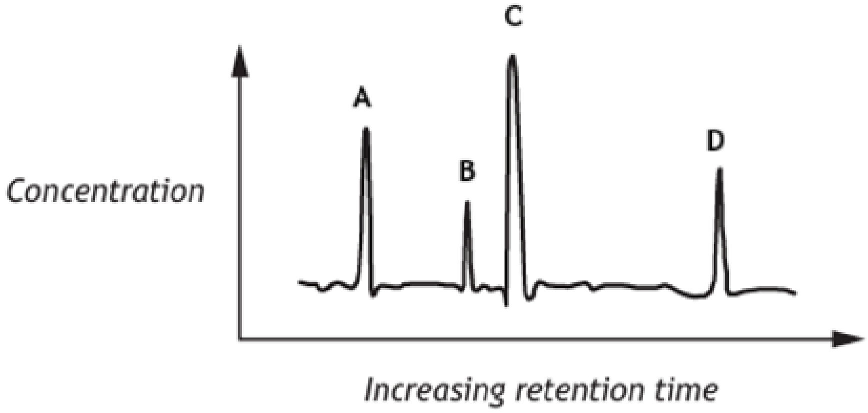
allows components to be identified by the time it takes to travel through a column containing either a polar or non-polar liquid.

This is known as the component's retention time R_t

The faster the R_t the less attracted the component is to the liquid in the column.

The longer the component takes to move through the column the more compatible the intermolecular bonding is between the component molecules and the liquid molecules held on the column.

The gas used to carry the sample along the chromatography column is helium. as it is **inert** and will not react with molecules being carried through the column.



Concentration of Substances

Substance B
Lowest concentration as it has the **smallest peak.**

Substance C
Highest concentration as it has the **largest peak.**

Retention Time

Substance A—shortest R_t

This means substance A passed through the column the **quickest** as it had the **weakest LDF's** with the liquid molecules in the column.

Substance D—longest R_t

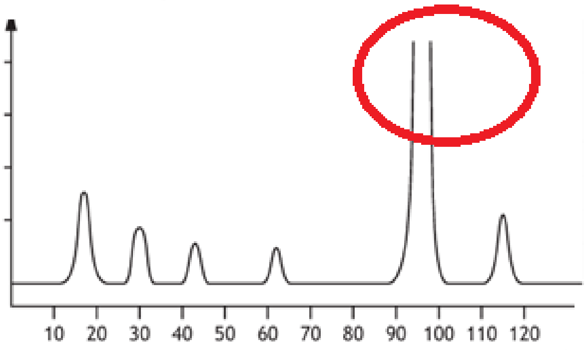
This means substance D passed through the column the **slowest** as it had **strongest LDF's** and attraction to molecules in the liquid within the column,

2. Gas-Liquid Chromatography

When the peak is too high

1. sample diluted

2. less sample used



peak too high