Rate of reaction:

A) Measuring mass of reactants

Downward curve:

- Gradient is at the steepest at the start, indicating that reactants are reacting at their maximum speed.
- Eventually become less steep because reactants are being used up and so the speed of reaction decreases.
- Eventually the reaction curve will end with a 0 gradient that indicates the completion of reaction when either of the reactants are used up.

B) Measuring change in volume of gaseous products

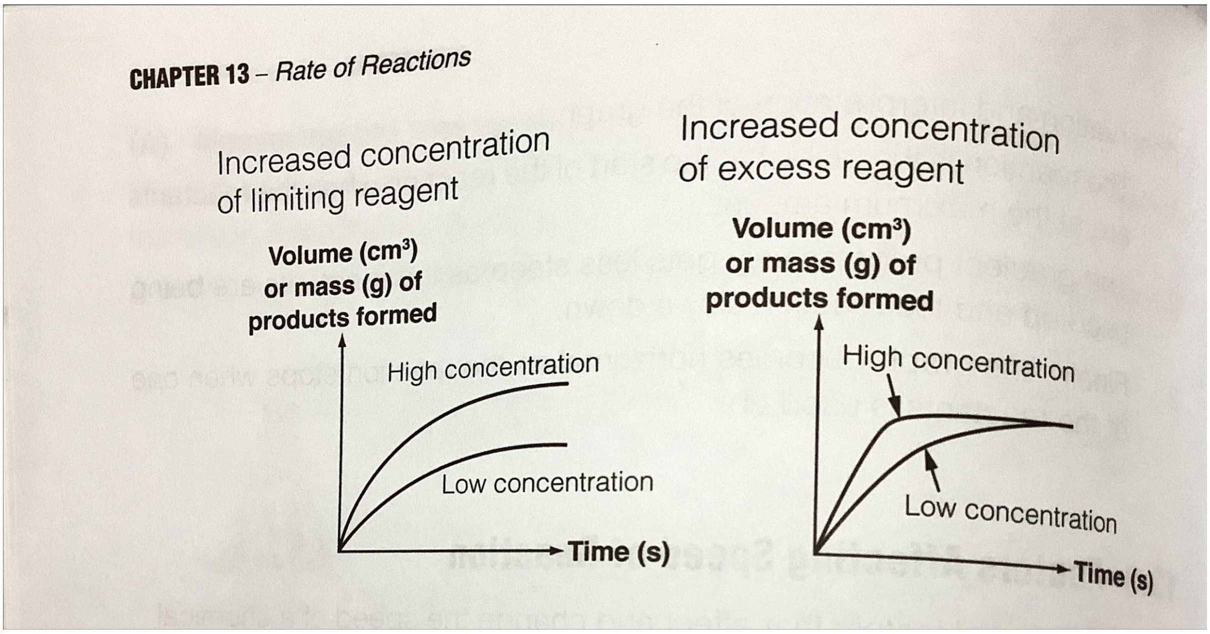
- Reaction is the fastest at the start of the reaction when the reactants are at the maximum amount.
- The gradient progressively gets less steep as the reactants are being used up and reaction slows down.
- Graph becomes horizontal as the reaction stops when one of the reactants is used up.

Factors affecting speed of reaction:

1. Concentration of the reactants (solutions)
2. Pressure (gas)
3. Particle size of reactants (solid)
4. Temperature

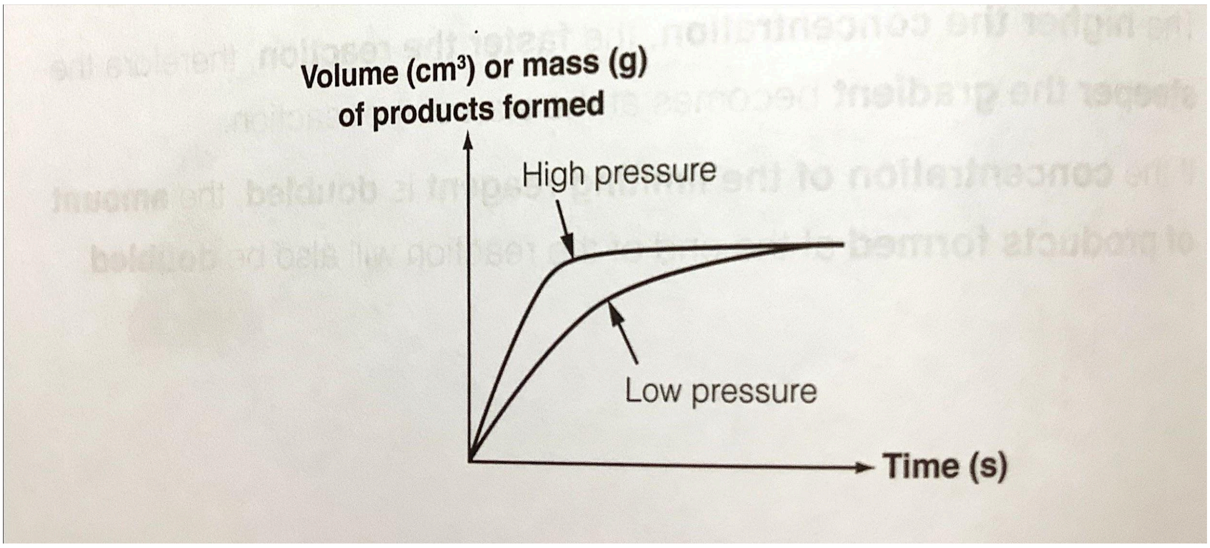
A) Concentration of reactants:

- Higher the concentration, the faster the speed of reaction.
- Increasing concentration -> Increasing number of reactants particles per unit volume.
- Less distance between reactant particles and higher frequency of collisions between them, causing increase in rate of reaction.
- Steeper the gradient.
- If concentration of a limiting reagent is doubled, the amount of product formed will also double.



B) Pressure of gaseous reactants:

- Higher the pressure, faster the reaction.
- Increasing pressure of gases will decrease the distance between the reactant particles, thus increasing the number of reactants per particle unit volume. (Same as concentration)
- Higher frequency of effective collision between reactant particles.
- Steeper gradient at the start of reaction.
- Amount of product formed remains the same.

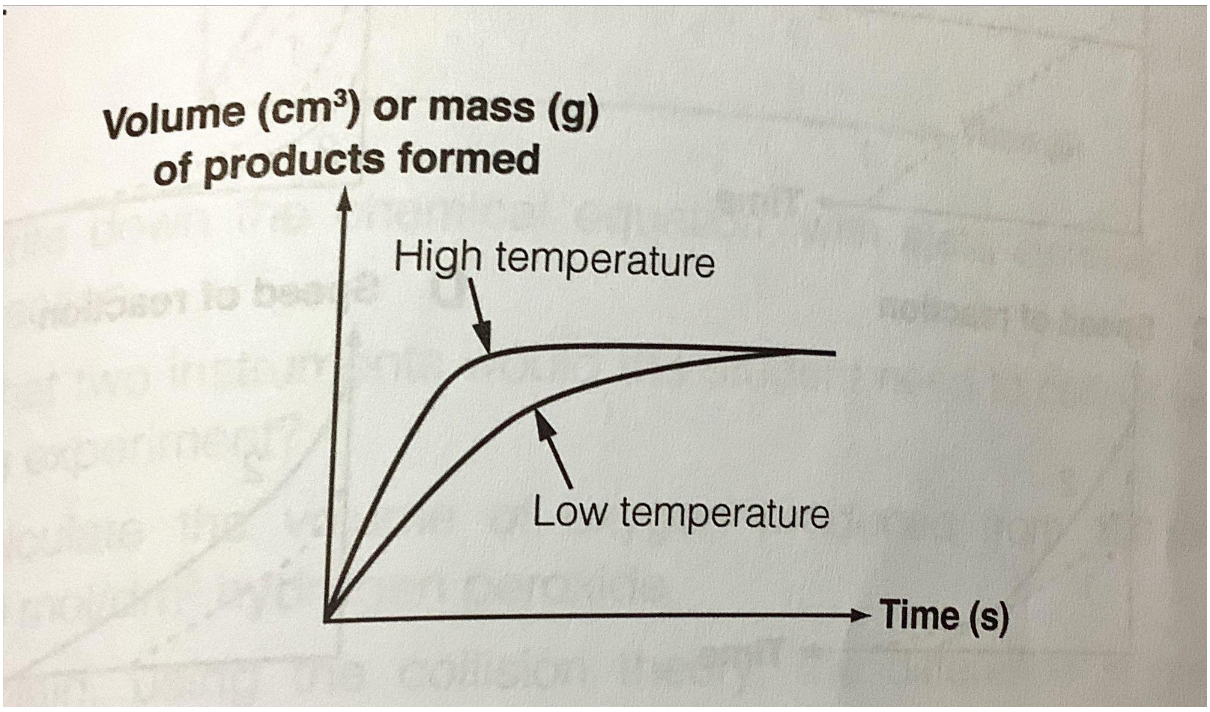


C) Particle size of solid reactants:

- The smaller the reactant particles, the faster rate of reaction.

- Smaller particles —> Surface area for reaction is larger —> more chances of effective collisions between reactant particles.
- Amount of product formed remains the same.

D) Temperature:
- Temperature rises, the particles of the reactants gain more kinetic energy and move faster.
- Causes the number of particles with energy higher than the activation energy to increase. (Able to overcome the activation energy)
- Increases the frequency of effective collisions between the reacting particles. Hence, speed of reaction increases.
- Steeper gradient at the start.
- Amount of product formed remains the same.



Qualitative analysis:

- Cations are identified using two main alkaline reagents:
  -> Aqueous sodium hydroxide
  -> Aqueous ammonia

A) Test for cations using aqueous sodium hydroxide:
- Sodium hydroxide will dissociate in water to form sodium hydroxide ions.
  -> NaOH(aq) —> Na+(aq) + OH-(aq)

- Changes: Precipitation and effervescent[1]

Cation present (+)	adding a few drops	adding in excess
Aluminium Al^{3+}	white precipitate is formed.	white precipitate dissolves to give a colourless solution. (Soluble in excess)
Zinc Zn^{2+}	white precipitate is formed.	white precipitate dissolves to give a colourless solution. (Soluble in excess)
Ca^{2+}	white precipitate is formed.	white precipitate is insoluble.
Copper (ll), Cu^{2+}	Light blue precipitate is formed.	light blue precipitate is insoluble.
Iron (ll), Fe^{2+}	green precipitate is formed.	green precipitate is insoluble.
Iron (lll), Fe^{3+}	reddish-brown precipitate is formed.	reddish-brown precipitate is insoluble.
Ammonium, NH^{4+}	No precipitate is formed, the solution remains colourless. Upon heating, a pungent gas is evolved which turns red litmus paper blue.	No visible reaction.

B) <u>Test for cations using aqueous ammonia:</u>

- Ammonia dissociates partially in water to form ammonium and hydroxide ions. Hence, concentration of OH- (aq) ions in aqueous is low.
  -> $NH^{3+}(g) + H_2O(l)$ <—> $NH^{4+}(aq) + OH^{-}(aq)$

[1] The vigorous release of gas.

Cation present	adding a few drops	adding in excess
Aluminium, Al^{3+}	white precipitate is formed.	white precipitate is insoluble
Zinc Zn^{2+}	white precipitate is formed	white precipitate dissolves to give a colourless solution. (Soluble in excess)
Calcium Ca^{2+}	no visible reaction	Due to the low concentration of hydroxide ions present in aqueous ammonia.
Iron(ll), Fe^{2+}	Green precipitate is formed	green precipitate is insoluble.
Iron(lll), Fe^{3+}	Reddish-brown precipitate is formed	reddish-brown precipitate is insoluble.

- Anions are usually identified using 4 types of test reagents:
  -> Dilute nitric acid
  -> Aqueous Silver nitrate/nitric acid
  -> Aqueous barium nitrate/nitric acid
  -> Aqueous sodium hydroxide/aluminium foil

- Changes:
  -> Colour of precipitation.
  -> Gas given off.

Anion (-)	Test	observation
Carbonate	1. Add dilute acid. 2. Pass any gas given off into limewater.	Effervescence is observed. Colourless, odourless, gas evolved to form white precipitate with lime water. Gas is carbon dioxide.

Nitrate	1. Add aqueous NaOH, followed by small pieces of aluminium foil/powder. 2. Warm the mixture and test any gas given off with moist red litmus paper.	Effervescence is observed. Colourless, pungent gas evolved to turn moist red litmus paper blue. Gas is ammonia gas.
Chloride	Add dilute nitric acid (to removed any carbonate present), followed by aqueous silver nitrate. *Dilute hydrochloric acid cannot be used,	White precipitate is formed.
Iodine	Add dilute nitric acid followed by aqueous silver nitrate.	Yellow precipitate is formed.
Sulfate	Add nitric acid followed by aqueous barium nitrate.	white precipitate is formed.

- Identification of gas:
  —> Colour, odour, reaction with specific chemical test.

A) Chemical test:

Gas	Test	Observations and explanations
Ammonia	Place moist red litmus paper at the mouth of the tube. Note*: Litmus paper must be moist otherwise no observation can be made.	Moist red litmus paper turns blue. NH_3, gas dissociates partially in water to form hydroxide ions, OH-, OH+ ions turns red litmus paper blue.

Chlorine	Place moist blue litmus paper at the mouth of the test tube.	Moist blue litmus paper turns red, and then bleach white. Cl_2
Sulfur dioxide	Place a piece of filter paper dipped in acidified potassium manganate (Vll) at the mouth of the test tube.	Purple acidified potassium manganate (Vll) turns colourless. Redox reaction takes place. Sulfur dioxide is a reducing agent. Reduces purple MnO_{4-} to colourless Mn^{2+} ions in solution.
Carbon dioxide	Pass gas evolved into limewater, $Ca(OH)_2$	White precipitate is formed. White precipitate is formed is, $CaCO_3$. $Ca(OH)_2(aq) + CO_2(g)$ —> $CaCO_3$ (S) + H_2O (l)
Oxygen	Insert a glowing splint into its test tube.	Glowing splint is relighted/rekindled. Oxygen supports burning combustion.
Hydrogen	Insert a lighted splint into the test tube.	Lighted splint is extinguished[2] with a ‘pop’ sound.

Reactivity series of metal:

[2] Key word for answering.

- The higher the metal is in the series, the more reactive it is.
- The more reactive a metal is, the greater its tendency to donate its valence electrons to form a positive ion with a stable electronic configuration.
- More reactive a metal, the more difficult it is to be extracted from its ore.

P — Potassium (K)
S — Sodium (Na)
C — Calcium (Ca)
M — Magnesium (Mg)
A — Aluminium [3] (Al)
C — Carbon [C]
Z — Zinc (Zn)
I — Iron (Fe)
L — Lead (Pb)
H — Hydrogen [H]
C — Copper (Cu)
S — Silver

- Determining reactivity of metal:
- This can be established by studying the reaction of metals with cold water or steam and acids.

A) Reaction of metals with cold water or steam.
- Some metals react with cold water to form metal hydroxides and hydrogen gas.
- Sodium + cold water —> sodium hydroxide + hydrogen
  -> E.g. $2Na\ (s) + 2H_2O\ (l) \longrightarrow 2NaOH\ (aq) + H_2\ (g)$
- Some metals however do not react with cold water but steam to form metal oxides and hydrogen gas.
- Zinc + steam —> zinc oxide + hydrogen
  -> $Zn\ (s) + H_2O\ (g) \longrightarrow ZnO\ (s) + H_2\ (g)$

Metal	Reaction with water	Reaction with steam
K - Potassium	Reacts explosively	React explosively
Na - Sodium	Reacts violently	Reacts explosively

[3] Not in the 2024 syllabus.

Ca - Calcium	Reacts readily	Reacts explosively
Mg - Magnesium	Reacts very slowly	Reacts violently
Zn - Zinc	No reaction	Reacts readily
Fe - Iron	No reaction	Reacts slowly

B) Reaction of metals with dilute hydrochloric acid (HCL)

1. Some metals react with dilute hydrochloric acid to form a salt and hydrogen gas.

- Sodium + hydrochloric acid —> 2Nacl (aq)+ H2 (g)

Metal	Reaction with dilute hydrochloric acid
K - Potassium	Reacts explosively
Na - Sodium	Reacts explosively
Ca - Calcium	Reacts violently
Mg - Magnesium	Reacts rapidly
Zn - Zinc	Reacts moderately fast
Fe - Iron	Reacts slowly
Pb - Lead	Reacts very slowly

- Metals below lead in the reactivity series do not react with dilute hydrochloric acid.

C) Displacement reactions of metals in reactivity series.

- A higher metal which has higher reactivity can displace any metal below it.
- Displacement reactions are useful for extraction of metal. A metal can be extracted from:
  -> soluble aqueous salt of the metal.
  -> oxide of the metal.

- two observation are needed;
  -> colour of metal deposited and change in colour of solution.

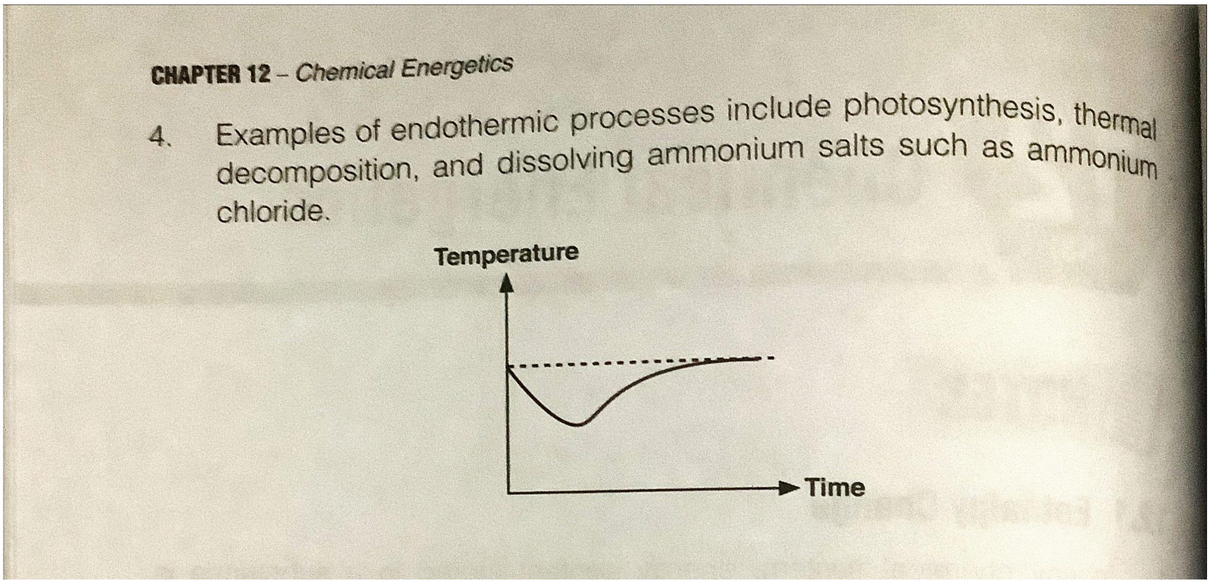
Metal	Colour of metal	Colour of salt solution
Magnesium	grey	colourless
Zinc	grey	colourless
Copper	red-brown	blue
Lead	grey	colourless
Iron	grey	Iron (ll): green Iron (lll): brown

D) Rusting of iron.

- Rusting of iron is the slow oxidation of iron to form hydrated iron (lll) oxide (rust).
- Presence of both air (oxygen) and water are needed for rusting to occur.
- Rust prevention:
  —> painting, plastic coating, oiling/greasing.

Chemical energetics:

- Endothermic reaction;
- Energy is absorbed or taken in from the surrounding by the reading system.
- Results in the cooling of its surroundings and a rise in energy content of the reacting system.



*examples in the picture.

- Exothermic reactions;
- Heat energy is released or given out from the reacting system to the surroundings.
- Results in the warming of its surroundings and a drop in energy content of the reacting system.

4. Examples of exothermic processes include neutralisation, combustion, respiration, and dissolving anhydrous salts such as sodium carbonate.