Reaction Types

Radical (Chain) Reactions

Reactions involving radicals — an atom or compound with an **unpaired electron**. Often, one of the **products** of the reaction is also a radical which can perform further reactions. This makes the process a **chain reaction**.

Redox

This is the name for a reaction that involves both **reduction** and **oxidation** processes. It is usually used to describe reactions that just involve **electron transfer**.

reduction

e.g. $Fe_{(s)} + Cu^{2+}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$

oxidation

Reduction

There are two possible definitions for this — the best is the **gain of electrons**. The other useful one is the loss of oxygen. Important point: oxidation and reduction **ALWAYS** happen **together** — it is impossible to have one without the other.

Reversible

This is the name given to any chemical reaction that can go **forwards** and **backwards** at the **same time**. That means that the reactants will form the products, but that the products will also react (or decompose) to give the reactants.

e.g. $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

Substitution

This is simply a reaction in which an atom (or group of atoms) in a molecule is **swapped** for a different atom (or group of atoms).

Thermal Decomposition

This is where one compound **breaks down**, under **heating**, into two or more simpler compounds. A classic example is the breakdown of any carbonate compound,

e.g. $CaCO_3 + \text{heat} \rightarrow CaO + CO_2$

Cracking of hydrocarbons is also an example.

I'm in the middle of a chain reaction...

1) Write down all the different types of reaction that each of the following could be classed as.
   a) burning ethanol
   b) iron + copper sulfate → iron sulfate + copper
   c) hydrochloric acid + sodium hydroxide → sodium chloride + water + heat
   d) propene (C_3H_6) + H_2 → propane (C_3H_8)

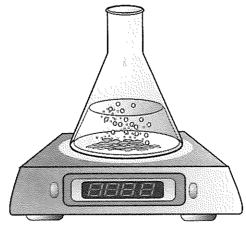
Reaction Rates

Measuring the Rate of a Reaction

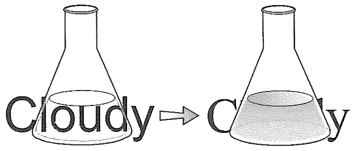
The **rate** of reaction is just a measure of how **fast** a particular reaction is going.

You need to know some of the ways that you can follow the rates of different reactions. They're all about measuring how fast the **reactants** are being **used up**, or measuring how fast the **products** of the reaction are **forming**.

There are lots of ways of measuring the rate of a reaction:



1) You can measure the **change in mass** that occurs during a reaction where gas is released as one of the products.

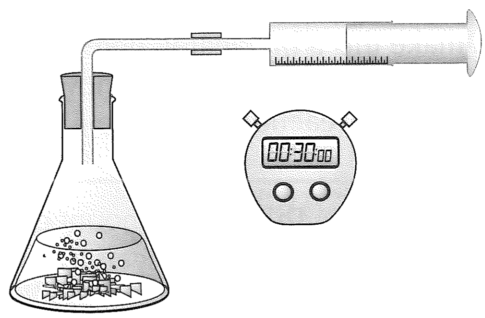


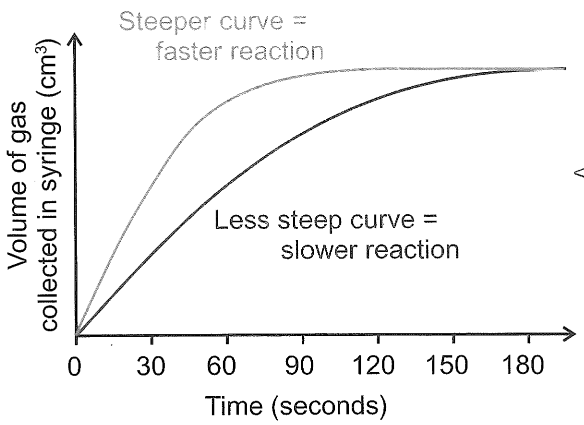
2) You can follow the **colour change** of a reaction. This includes precipitation reactions, where the solution turns cloudy as more of the product is made.
3) You can measure changes in **temperature** or **pH** that occur during the reaction.
4) You can measure the **volume of gas** produced during a reaction.

EXAMPLE: Measuring the rate of reaction between hydrochloric acid and magnesium metal.

magnesium + hydrochloric acid → magnesium chloride + hydrogen

- Use a **gas syringe** to collect the hydrogen gas that is given off during the reaction.
- Use a **stopwatch** to **time** the reaction.
- At **timed intervals**, say every 30 seconds, **record** how much hydrogen gas has been produced.





Plotting graphs lets you compare rates of reactions.

(Another way to measure the rate of this reaction would be to measure the decrease in **mass** as hydrogen gas is lost from the reaction container.)

My rate of chocolate biscuit consumption is worryingly high...

1) Describe how you could measure the rates of the following reactions:
   a) The endothermic reaction between citric acid and sodium bicarbonate to give carbon dioxide, water and a sodium salt.
   b) The precipitation reaction between sodium thiosulfate and hydrochloric acid to form a sulfur precipitate, sulfur dioxide gas, sodium chloride and water.
   c) The reaction between solid calcium carbonate and hydrochloric acid to produce calcium chloride and carbon dioxide gas.

Collision Theory

Particles Need to *Collide* in Order to *React*

Reaction rates are explained by **collision theory**. It's based on the idea that particles in liquids and gases are always **moving around** and **colliding** with each other.
Not every **collision** results in the particles **reacting**. The following **conditions** need to be right:

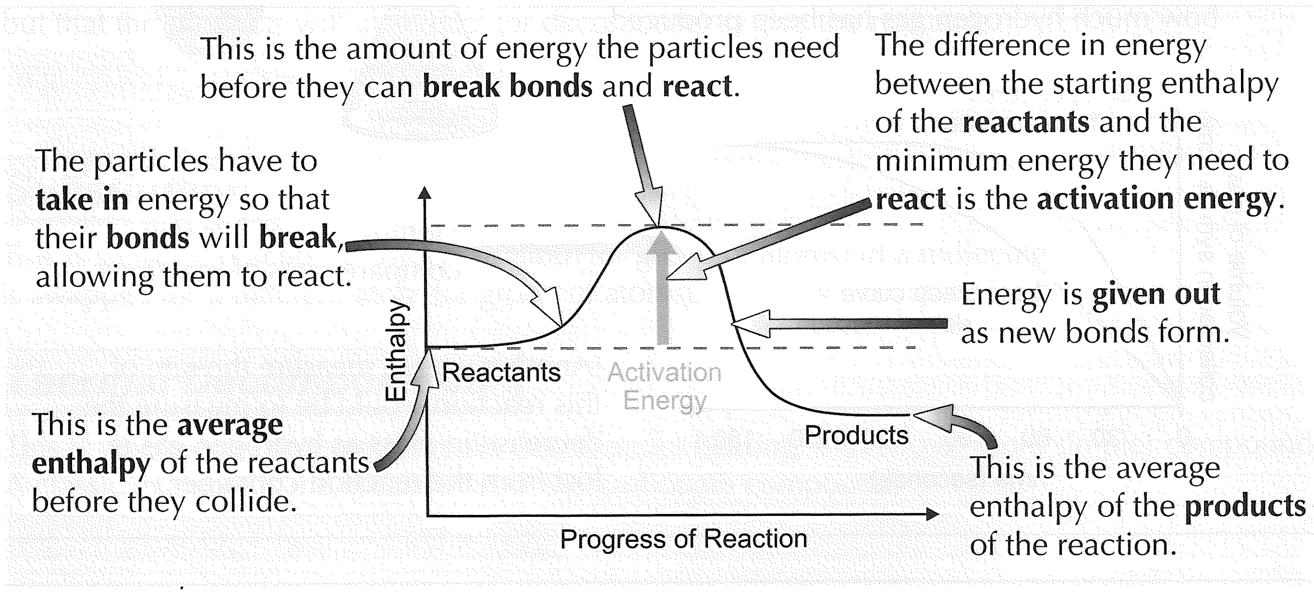
- The particles need to collide in the **right direction**. They need to be **facing** each other the right way.
- The particles need to collide with at least a certain **minimum** amount of **energy**.

Collision theory states that the **more collisions** there are, and the **more energy** these collisions have, the **more likely** particles are to react.

Particles Need Enough *Energy* to React

1) The **minimum amount of kinetic** (movement) **energy** particles need to react is known as the **activation energy**. This energy is used to **break the bonds** to start the reaction.
2) Reactions with **low activation energies** often happen **pretty easily**, but reactions with **high activation energies** don't — you have to give the particles **extra energy** (e.g. by **heating** them).

To make this a bit clearer, here's an **enthalpy profile diagram** — enthalpy is just a fancy word for energy. These diagrams show how the **enthalpy** of the reacting particles changes over the course of a reaction (see page 41 for more on enthalpy changes).



Toast and a large cup of tea — that's my morning activation energy...

1) Two particles in a reaction vessel collide but don't react. Give two reasons why the reaction may not have happened.
2) What is the activation energy of a reaction?
3) Draw an enthalpy profile diagram for a reaction. On your diagram, label the reactants, products and activation energy.

Reaction Rates and Catalysts

Changing the Rate of Reaction

The rate of reaction depends on **how often** particles **collide** (see page 32) and how **likely** the collisions are to be **successful**.
More frequent **successful** collisions mean a **faster** rate of reaction.

These factors all **increase** the rate of reaction:

1) **Increasing temperature** — the particles tend to have more kinetic energy. This means that they move around faster, and so are more likely to collide with each other **and** have enough energy to react.
2) **Increasing concentration (or pressure in gases)** — this means that the particles of reactant will be closer together, so they will be more likely to collide.
3) **Increasing the surface area of a solid reactant** — this increases the number of particles of the solid reactant able to come into contact with other reactants.

EXAMPLE: Predict whether magnesium dust or magnesium ribbon will react faster with hydrochloric acid.

Magnesium dust has a larger **surface area** than magnesium ribbon.
Increasing the surface area of a solid reactant increases the rate of reaction, so **magnesium dust** will react faster than magnesium ribbon.

Catalysts Speed Up Reactions

You met **activation energy** on the last page — it's just the **minimum** amount of energy needed for a reaction to happen.
A **catalyst increases** the **rate** of a reaction by **lowering** its activation energy.

A **catalyst** is any substance which changes the **rate** of a reaction, without being **changed** or **used up** itself.

Catalysts are also very **specific** — different reactions will only be sped up by **certain catalysts**.

There are loads of advantages to using catalysts:

1) Catalysts reduce the need for **high temperatures** and **pressures** in industrial reactions, like hydrocarbon cracking (see page 28) and ethanol production (see page 36). This makes these processes **cheaper** to run.
2) Using lower temperatures also means less **energy demand**, and so lower CO_2 emissions.

Tabbys are number one on my cat list...

1) Describe two things you could do to increase the rate of a reaction between aqueous species.
2) Why does increasing the pressure increase the rate of a reaction between gases?
3) What's a catalyst?
4) Give two advantages of using a catalyst in industrial reactions.