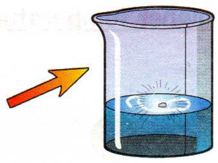
Energy Transfer in Reactions

Whenever chemical reactions occur energy is transferred to or from the surroundings.

In an Exothermic Reaction, Heat is Given Out

An EXOTHERMIC reaction is one which transfers energy to the surroundings, usually in the form of heat and usually shown by a rise in temperature.

1) The best example of an exothermic reaction is burning fuels — also called COMBUSTION. This gives out a lot of heat — it's very exothermic.
2) Neutralisation reactions (acid + alkali) are also exothermic — see page 65.
3) Many oxidation reactions are exothermic. For example, adding sodium to water produces heat, so it must be exothermic — see page 75. The sodium emits heat and moves about on the surface of the water as it is oxidised.



4) Exothermic reactions have lots of everyday uses. For example, some hand warmers use the exothermic oxidation of iron in air (with a salt solution catalyst) to generate heat. Self heating cans of hot chocolate and coffee also rely on exothermic reactions between chemicals in their bases.

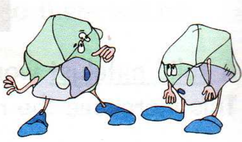
In an Endothermic Reaction, Heat is Taken In

An ENDOTHERMIC reaction is one which takes in energy from the surroundings, usually in the form of heat and is usually shown by a fall in temperature.

Endothermic reactions are much less common. Thermal decompositions are a good example:

Heat must be supplied to make calcium carbonate decompose to make quicklime.

$$CaCO_3 \rightarrow CaO + CO_2$$



Endothermic reactions also have everyday uses. For example, some sports injury packs use endothermic reactions — they take in heat and the pack becomes very cold. More convenient than carrying ice around.

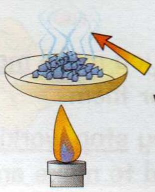
Reversible Reactions Can Be Endothermic and Exothermic

In reversible reactions (see page 55), if the reaction is endothermic in one direction, it will be exothermic in the other direction. The energy absorbed by the endothermic reaction is equal to the energy released during the exothermic reaction. A good example is the thermal decomposition of hydrated copper sulfate.

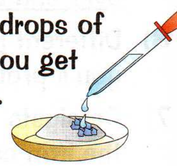
$$\text{hydrated copper sulfate} \underset{\text{exothermic}}{\overset{\text{endothermic}}{\rightleftharpoons}} \text{anhydrous copper sulfate + water}$$

"Anhydrous" just means "without water", and "hydrated" means "with water".

1) If you heat blue hydrated copper(II) sulfate crystals it drives the water off and leaves white anhydrous copper(II) sulfate powder. This is endothermic.



2) If you then add a couple of drops of water to the white powder you get the blue crystals back again. This is exothermic.



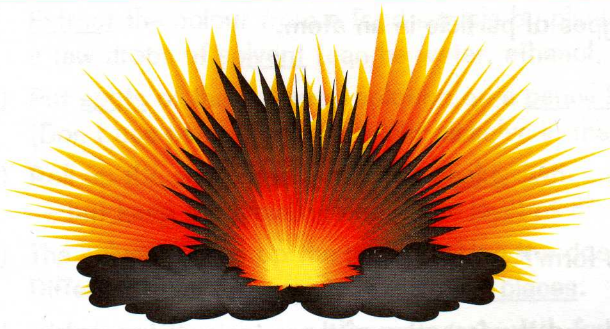
Right, so burning gives out heat — really...

This whole energy transfer thing is a fairly simple idea — don't be put off by the long words. Remember, "exo-" = exit, "-thermic" = heat, so an exothermic reaction is one that gives out heat. And "endo-" = erm... the other one. Okay, so there's no easy way to remember that one. Tough.

Rate of Reaction

Reactions can be fast or slow — you've probably already realised that. But you need to know what affects the rate of a reaction, as well as what you can do to measure it. You'll be on the edge of your seat. Honest.

Reactions Can Go at All Sorts of Different Rates



1) One of the slowest is the rusting of iron (it's not slow enough though — what about my little MGB).
2) A moderate speed reaction is a metal (like magnesium) reacting with acid to produce a gentle stream of bubbles.
3) A really fast reaction is an explosion, where it's all over in a fraction of a second.

The Rate of a Reaction Depends on Four Things:

1) Temperature
2) Concentration — (or pressure for gases)
3) Catalyst
4) Surface area of solids — (or size of solid pieces)

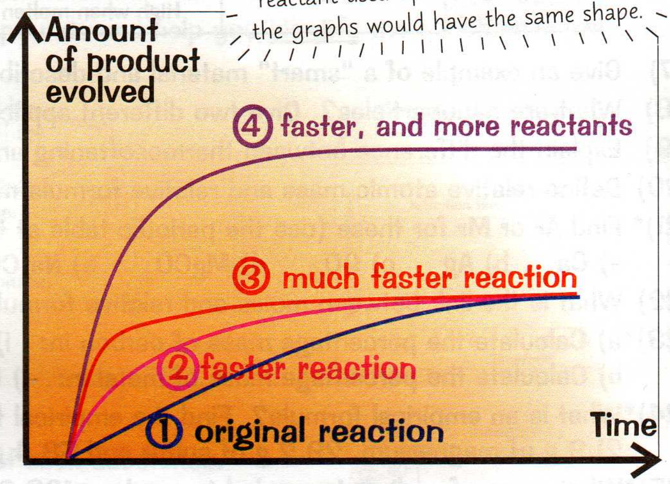
LEARN THEM!

Typical Graphs for Rate of Reaction

The plot below shows how the rate of a particular reaction varies under different conditions. The quickest reaction is shown by the line with the steepest slope. Also, the faster a reaction goes, the sooner it finishes, which means that the line becomes flat earlier.

1) Graph 1 represents the original fairly slow reaction. The graph is not too steep.
2) Graphs 2 and 3 represent the reaction taking place quicker but with the same initial amounts. The slope of the graphs gets steeper.
3) The increased rate could be due to any of these:
   a) increase in temperature
   b) increase in concentration (or pressure)
   c) catalyst added
   d) solid reactant crushed up into smaller bits.

You could also show the amount of reactant used up over time instead — the graphs would have the same shape.



4) Graph 4 produces more product as well as going faster. This can only happen if more reactant(s) are added at the start. Graphs 1, 2 and 3 all converge at the same level, showing that they all produce the same amount of product, although they take different times to get there.

How to get a fast, furious reaction — crack a wee joke...

Industrial reactions generally use a catalyst and are done at high temperature and pressure. Time is money, so the faster an industrial reaction goes the better... but only up to a point. Chemical plants are quite expensive to rebuild if they get blown into lots and lots of teeny tiny pieces.

Measuring Rates of Reaction

Ways to Measure the Rate of a Reaction

The rate of a reaction can be observed either by measuring how quickly the reactants are used up or how quickly the products are formed. It's usually a lot easier to measure products forming.

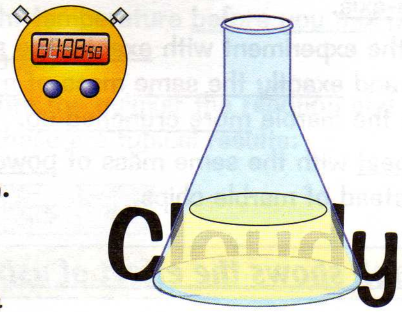
The rate of reaction can be calculated using the following formula:

$$\text{Rate of Reaction} = \frac{\text{Amount of reactant used or amount of product formed}}{\text{Time}}$$

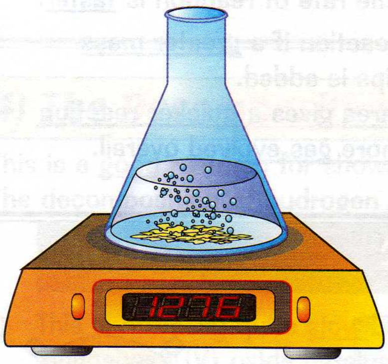
There are different ways that the rate of a reaction can be measured. Learn these three:

1) Precipitation

1) This is when the product of the reaction is a precipitate which clouds the solution.
2) Observe a mark through the solution and measure how long it takes for it to disappear.
3) The quicker the mark disappears, the quicker the reaction.
4) This only works for reactions where the initial solution is rather see-through.
5) The result is very subjective — different people might not agree over the exact point when the mark 'disappears'.



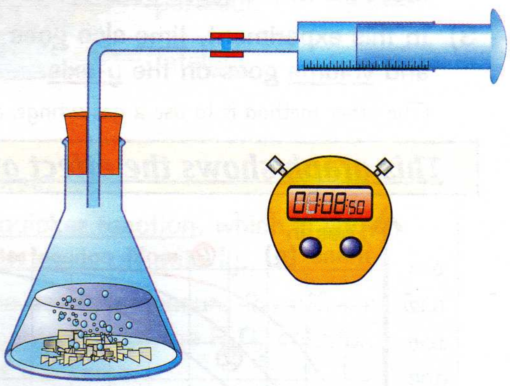
2) Change in Mass (Usually Gas Given Off)



1) Measuring the speed of a reaction that produces a gas can be carried out on a mass balance.
2) As the gas is released the mass disappearing is easily measured on the balance.
3) The quicker the reading on the balance drops, the faster the reaction.
4) Rate of reaction graphs are particularly easy to plot using the results from this method.
5) This is the most accurate of the three methods described on this page because the mass balance is very accurate. But it has the disadvantage of releasing the gas straight into the room.

3) The Volume of Gas Given Off

1) This involves the use of a gas syringe to measure the volume of gas given off.
2) The more gas given off during a given time interval, the faster the reaction.
3) A graph of gas volume against time elapsed could be plotted to give a rate of reaction graph.
4) Gas syringes usually give volumes accurate to the nearest millilitre, so they're quite accurate. You have to be quite careful though — if the reaction is too vigorous, you can easily blow the plunger out of the end of the syringe!



OK have you got your stopwatch ready *BANG!* — oh...

Each method has its pros and cons. The mass balance method is only accurate as long as the flask isn't too hot, otherwise you lose mass by evaporation as well as by the reaction. The first method isn't very accurate, but if you're not producing a gas you can't use either of the other two. Ah well.

Collision Theory

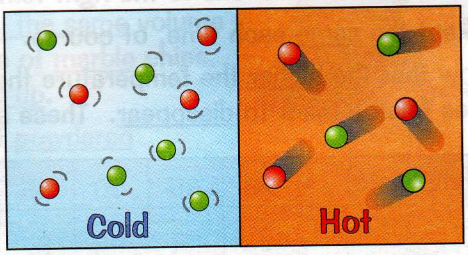
Reaction rates are explained by collision theory. It's really simple. It just says that the rate of a reaction simply depends on how often and how hard the reacting particles collide with each other. The basic idea is that particles have to collide in order to react, and they have to collide hard enough (with enough energy).

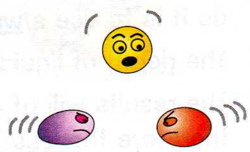
More Collisions Increases the Rate of Reaction

The effects of temperature, concentration and surface area on the rate of reaction can be explained in terms of how often the reacting particles collide successfully.

1) HIGHER TEMPERATURE increases collisions

When the temperature is increased the particles all move quicker.
If they're moving quicker, they're going to collide more often.

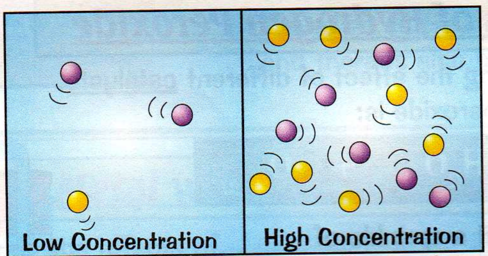




2) HIGHER CONCENTRATION (or PRESSURE) increases collisions

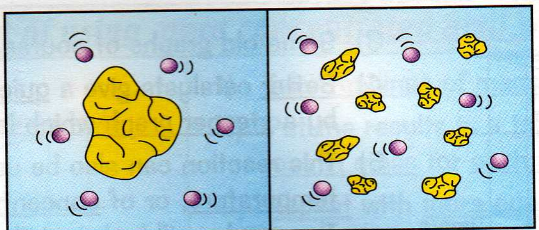
If a solution is made more concentrated it means there are more particles of reactant knocking about between the water molecules which makes collisions between the important particles more likely.

In a gas, increasing the pressure means the particles are more squashed up together so there will be more frequent collisions.



3) LARGER SURFACE AREA increases collisions

If one of the reactants is a solid then breaking it up into smaller pieces will increase the total surface area. This means the particles around it in the solution will have more area to work on, so there'll be more frequent collisions.



Collision theory — the lamppost ran into me...

Once you've learnt everything off this page, the rates of reaction stuff should start making a lot more sense to you. Isn't it nice when everything starts to fall into place... The concept's fairly simple — the more often particles bump into each other, and the harder they hit when they do, the faster the reaction happens.

Collision Theory and Catalysts

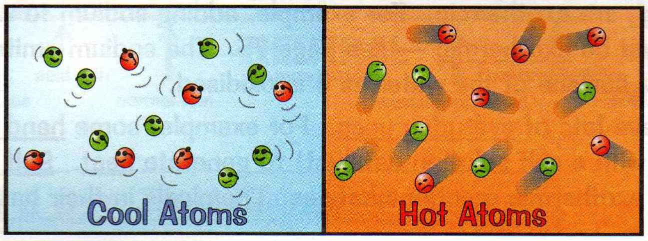
Without enough activation energy, it's game over before you start.

Faster Collisions Increase the Rate of Reaction

Higher temperature also increases the energy of the collisions, because it makes all the particles move faster.

Increasing the temperature causes faster collisions

Reactions only happen if the particles collide with enough energy.
The minimum amount of energy needed by the particles to react is known as the activation energy.
At a higher temperature there will be more particles colliding with enough energy to make the reaction happen.



Catalysts Speed Up Reactions

Many reactions can be speeded up by adding a catalyst.

A catalyst is a substance which speeds up a reaction, without being changed or used up in the reaction.

A solid catalyst works by giving the reacting particles a surface to stick to.
This increases the number of successful collisions (and so speeds the reaction up).

Catalysts Help Reduce Costs in Industrial Reactions

1) Catalysts are very important for commercial reasons — most industrial reactions use them.
2) Catalysts increase the rate of the reaction, which saves a lot of money simply because the plant doesn't need to operate for as long to produce the same amount of stuff.
3) Alternatively, a catalyst will allow the reaction to work at a much lower temperature. That reduces the energy used up in the reaction (the energy cost), which is good for sustainable development (see page 55) and can save a lot of money too.
4) There are disadvantages to using catalysts, though.
5) They can be very expensive to buy, and often need to be removed from the product and cleaned. They never get used up in the reaction though, so once you've got them you can use them over and over again.
6) Different reactions use different catalysts, so if you make more than one product at your plant, you'll probably need to buy different catalysts for them.
7) Catalysts can be 'poisoned' by impurities, so they stop working, e.g. sulfur impurities can poison the iron catalyst used in the Haber process (used to make ammonia for fertilisers). That means you have to keep your reaction mixture very clean.

Catalysts are like great jokes — they can be used over and over...

And they're not only used in industry... every useful chemical reaction in the human body is catalysed by a biological catalyst (an enzyme). If the reactions in the body were just left to their own devices, they'd take so long to happen, we couldn't exist. Quite handy then, these catalysts.