

## IGCSE CHEMISTRY REVISION NOTES

### TOPIC 7: Chemical Reactions

#### Rates of reaction

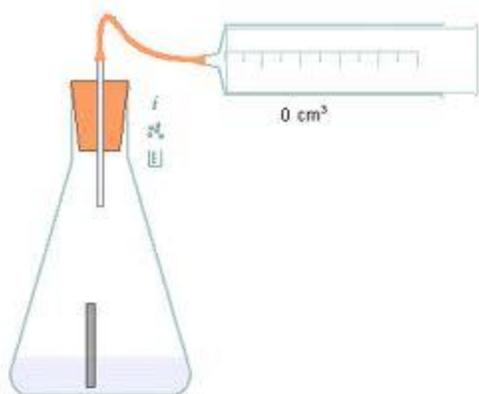
Rate is a measure of how fast a reaction goes, its speed.

The rate of a chemical reaction is the speed at which the reactants are used up or the speed at which new products are formed.

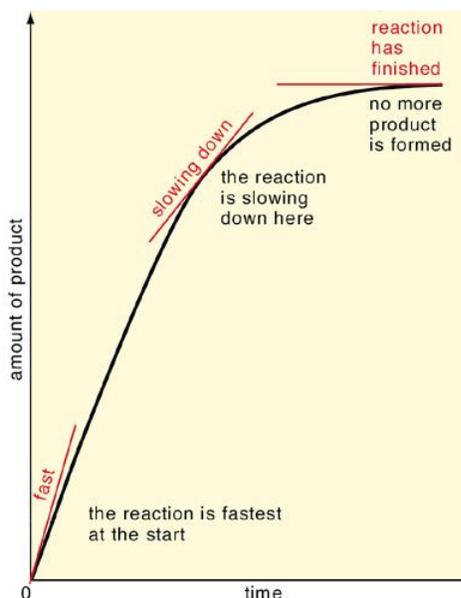
In order to measure the rate of a reaction you need to measure a change (for example in mass, or volume or colour) over a period of time.

#### Examples

In an experiment where a gas is produced (e.g., the reaction of magnesium with hydrochloric acid which produces hydrogen), the volume of gas produced can be measured with time.



The results of such experiments can be used to plot graphs. The rate of the reaction is the slope (gradient) of the graph. Reactions are always fastest at the start because this is when the



concentration of the reactants is high. The rate of reaction slows down as the reactants get used up. Eventually no more product (in our example gas) is formed and the reaction has finished.

The average rate of reaction can be calculated by:

$$\text{Average rate} = \frac{\text{volume change}}{\text{Time}}$$

if the volume was in  $\text{cm}^3$  the units would be  $\text{cm}^3/\text{s}$ .

Equally it could be a change in mass with time.

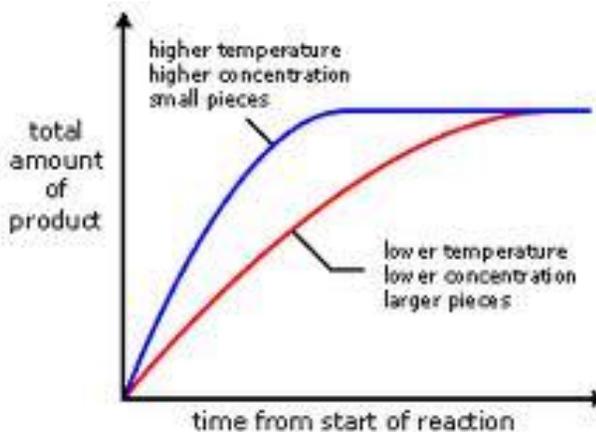
## Changing the rate of a reaction

When a chemical reaction occurs the reacting particles need to meet up with each (collide) in order to form new bonds. They also need to have enough energy to react (the activation energy) when they collide. These ideas together are called collision theory and they can be used to explain how we can make reactions go more slowly or more quickly.

There are four ways in which we can increase the rate of a chemical reaction.

### 1. Temperature

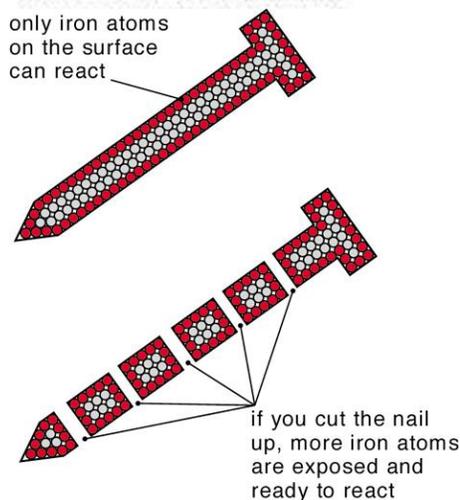
Increasing the temperature increases the rate of the reaction. When the temperature is increased the particles all move more quickly and have more energy (more will have the activation energy). Thus they collide more often and with more energy so the rate of reaction increases. This can be shown on a graph



### 2. Size of solid particles (surface area)

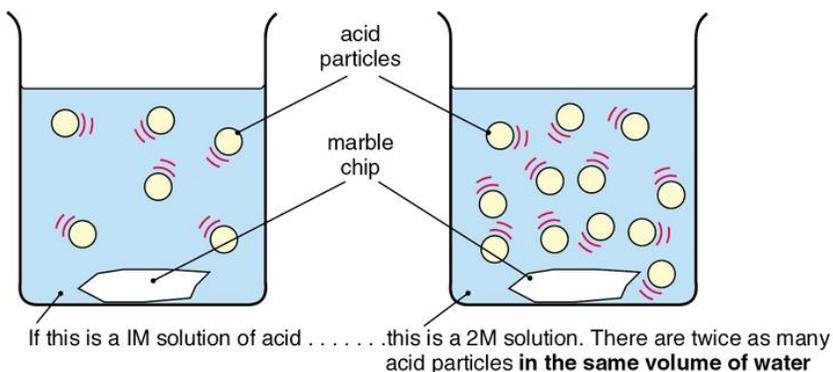
If one of the reactants is a solid then breaking it up into smaller pieces will increase its surface area. This means that the particles around it will have more area to work on and there will be more collisions.

In flour mills and coal mines this can lead to explosions. The air can fill with fine flour dust (or coal dust) with a large total surface area. A spark could cause the dust to ignite and explode



### 3. Concentration (or pressure)

If a solution is made more concentrated that means there are more reactant molecules between the water

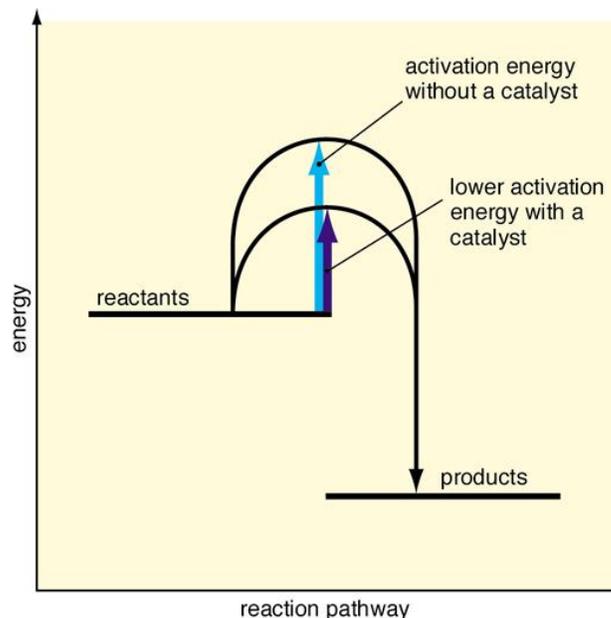


molecules. This makes collisions between the reacting molecules more likely. In a gas increasing the pressure means that the particles are squashed closer together and so they will collide more often.

4. **Catalyst** A catalyst increases the rate of a chemical reaction without being used up in the reaction. It can work by lowering the activation energy for the reaction and thus it increases the number of molecules with enough energy to react. The activation energy is the energy needed to start the reaction.

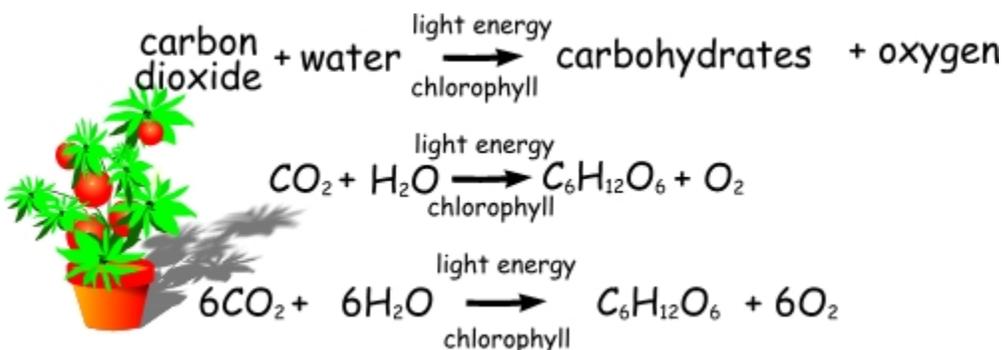
Catalysts are important in industry to speed up reactions and save time and money.

Enzymes are proteins made by living cells and they act as biological catalysts. They are used in detergents to dissolve stains and the enzymes in yeast are used in bread making. Enzymes differ from other catalysts because they only work in limited ranges of pH and temperature



### The effect of light

Some reactions obtain the energy they need from light. They are called photochemical reactions. Photosynthesis is a photochemical reaction. Carbon dioxide reacts with water in the leaves of plants using chlorophyll as a catalyst and light from the sun to produce glucose and oxygen. If you increase the light intensity you increase the reaction rate



Silver halides (silver chloride, silver bromide and silver iodide) decompose in the presence of light forming silver.

Half equation:  $\text{Ag}^+(\text{s}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$

This was used in photographic film. The reaction speeds up as light intensity increases.

### Reversible Reactions

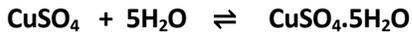
Many chemical reactions are reversible. This means they can go in either direction;

Reactants  $\rightarrow$  Products called the forward reaction

AND Products  $\rightarrow$  Reactants called the reverse reaction

This is represented by Reactants  $\rightleftharpoons$  Products

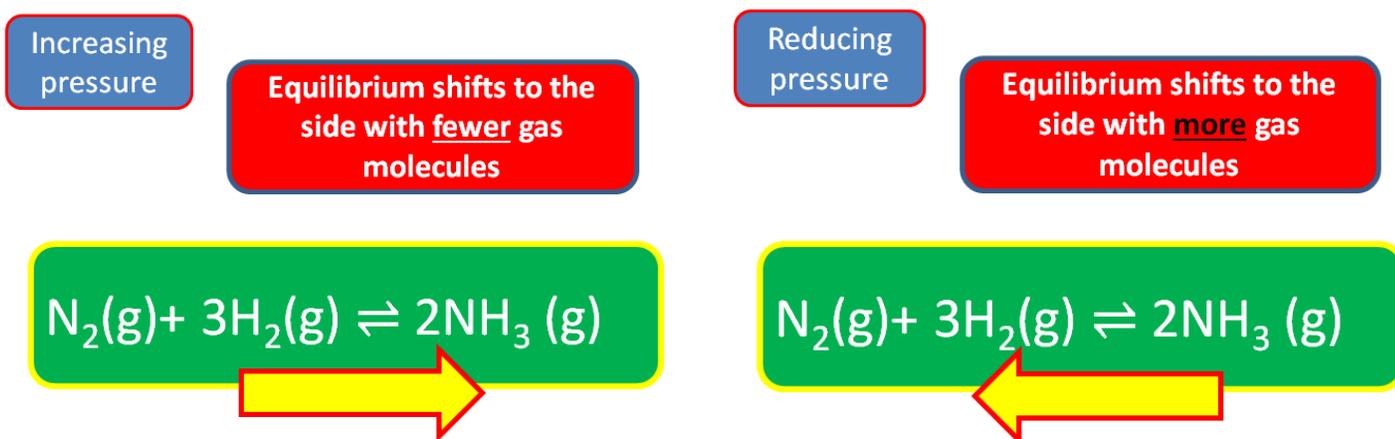
An example is



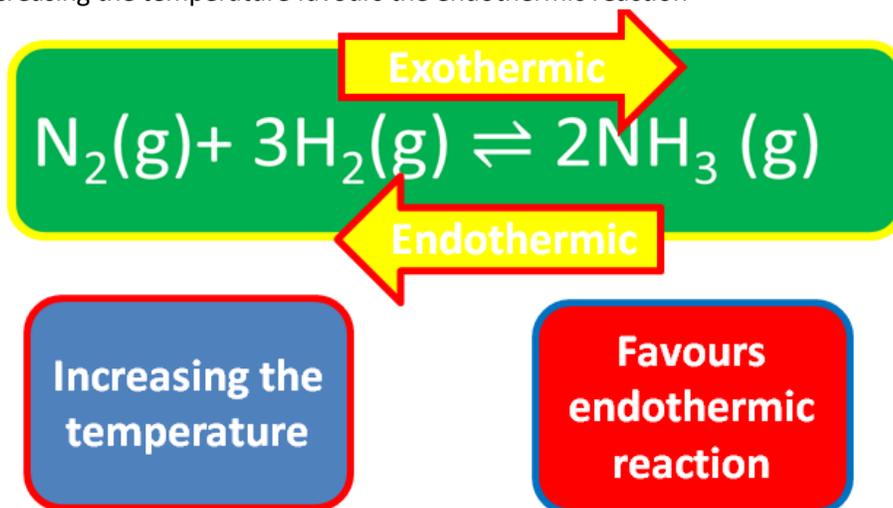
Reversible reactions can reach a point called equilibrium. Here the rate of the forward reaction is equal to the rate of the reverse reaction. The reactions continue but the concentration of reactants and products remains constant.

If the reaction conditions are changed this can affect the position of equilibrium:

1. In gaseous reactions the equilibrium can be affected by changing the pressure. Increasing the pressure favours the reaction with the least number of gas molecules.  
e.g.



2. Increasing the temperature favours the endothermic reaction



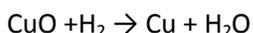
3. Increasing the concentration of the reactants or removing products favours the forward reaction
4. Increasing the concentration of products and reducing the reactants favours the reverse reaction

Adding catalysts do not affect the position of equilibrium as they affect the rate of both the forward and reverse reactions. They reduce the time taken to reach equilibrium.

### Oxidation and reduction (redox reactions)

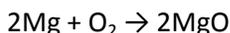
Many reactions in chemistry are redox reactions. If a reaction is not a precipitation reaction or a neutralization reaction then the chances are it's a redox reaction. The word redox is short for reduction and oxidation. In simple terms reduction means loss of oxygen and oxidation is gain of oxygen. In their broader definitions reduction is the gain of electrons and oxidation is the loss of electrons.

When copper (II) oxide is heated with hydrogen, the black copper oxide turns the red brown colour of copper. The following reaction occurs.

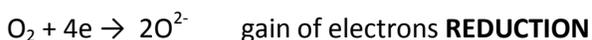


The copper oxide loses oxygen to become copper so is reduced. The hydrogen gains oxygen and is oxidized.

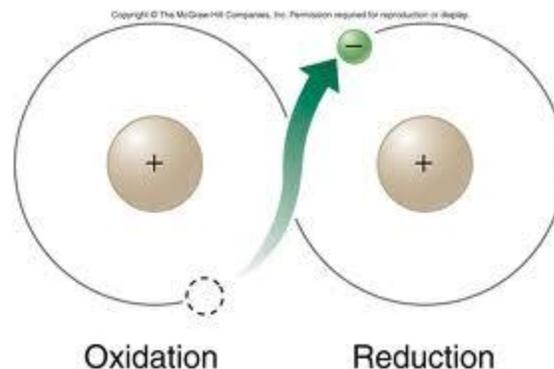
When magnesium is burned in air it reacts with the oxygen in the air to form magnesium oxide.



The magnesium has been oxidized as it gains oxygen however the magnesium has also lost electrons to form magnesium ions and the oxygen has gained electrons. Oxidation can be defined as the loss of electrons and reduction as the gain of electrons. The famous **OILRIG** This can be shown by half equations



Oxidation and reduction always occur together. There are always two half equations one showing reduction and one showing oxidation.



### Oxidation states

This tells you how many electrons each atom of an element has lost, gained or shared in forming a compound. The rules for oxidation states are:

1. All atoms in compounds have an oxidation state. Atoms in elements have an oxidation state of zero.
2. The oxidation states of each atom in the formula of a compound must add up to zero
3. Most elements have the same oxidation state in all their compounds. This can usually be determined from their group in the periodic table. e.g. group 1 metals have an oxidation state of +1 in compounds and group 7 elements have an oxidation state of -1.

4. Transition metals have variable oxidation states
5. In names the oxidation state is given as Roman numeral.

In assigning oxidation states always go with the elements in the compound you are sure of and then work out the rest.

Elements you can be sure of:

- Group 1 metals +1
- Group 2 metals +2
- Aluminium +3
- Group 7 elements (when not combined with oxygen) -1
- Oxygen (except in peroxides) -2

If oxidation states change during a reaction it is a redox reaction. Oxidation leads to an increase in oxidation number and reduction leads to a decrease in oxidation number

Oxidation states are used in the names of compounds if the oxidation state of an element varies. Transition metals have variable oxidation states so the oxidation state is given in the name. For example copper has the oxidation state of +1 and +2 so it can form two oxides copper (I) oxide ( $\text{Cu}_2\text{O}$ ) which is red-brown ) and copper (II) oxide ( $\text{CuO}$ ) which is black



### Oxidising and reducing agents

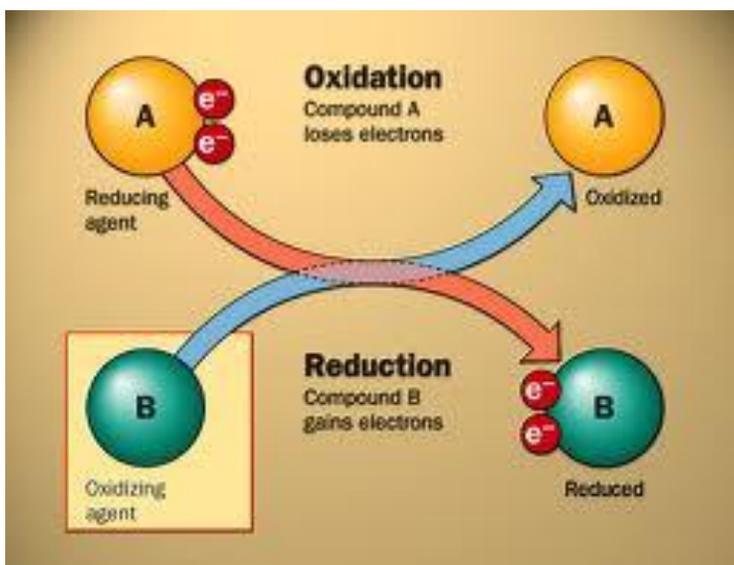
A reducing agent brings about reduction, and is itself oxidized  
 An oxidizing agent brings about oxidation , and is its self reduced

Reducing agents are sometimes called **reductants**. Oxidising agents are sometimes called **oxidants**.

Some substances are powerful oxidizing agents as they have a strong drive to gain electrons .  
 Examples are oxygen and chlorine

Some substances are powerful reducing agents because they have a strong drive to give up

are  
 carbon



electrons. Examples  
 hydrogen, carbon,  
 monoxide, and  
 reactive metals like  
 sodium.

### Tests for oxidizing agents and reducing agents

When some substances are oxidized and reduced there is a colour change. So these are useful for tests

#### Potassium manganate (VII)

This is a powerful oxidizing agent. When it is reduced the colour changes from purple to colourless.

$\text{MnO}_4^-$	→	$\text{Mn}^{2+}$
Oxidation state +VII (+7)		Oxidation state +II (+2)
Purple colour		colourless



#### Potassium iodide

This is a reducing agent. When the iodide ion is oxidized to iodine the colour changes from colourless to brown

$2\text{I}^-$	→	$\text{I}_2$
Oxidation state -1		Oxidation state 0
colourless		brown

