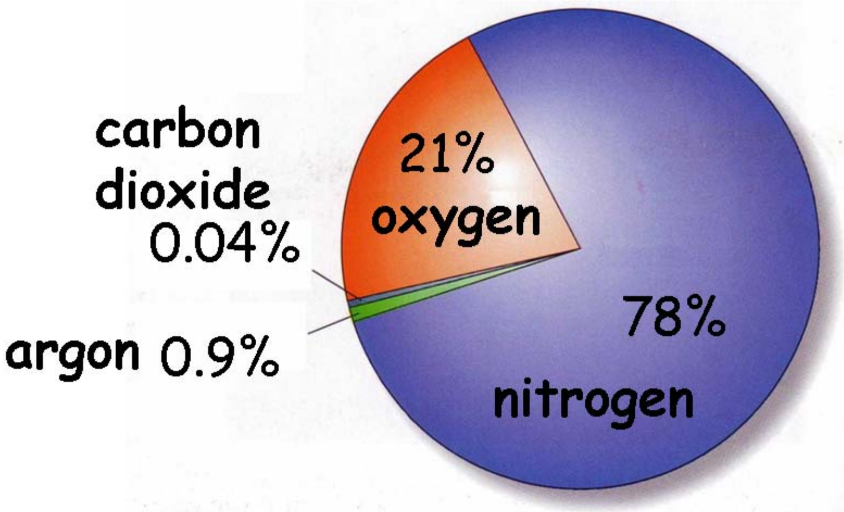
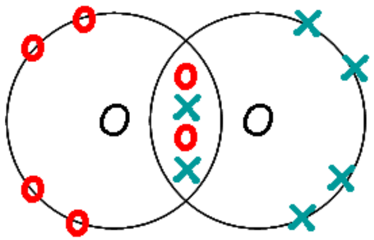
Composition of the air

Air is a **mixture** of gases. The approximate amount if each gas in dry air is shown in the pie chart (right), but you should be aware that air also contains a variable amount of water vapour which makes the air humid. There are also many trace gases present in extremely small amounts.



Oxygen

Formula: O_2
Bonding: covalent
Appearance: colourless gas



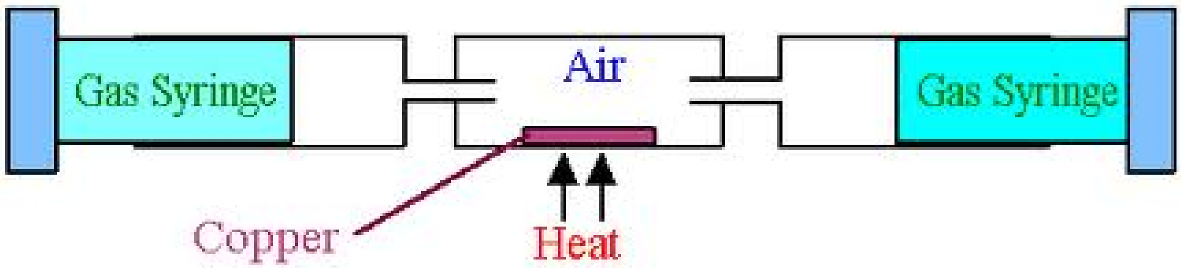
Oxygen is one of the two main gases in our atmosphere, the other being nitrogen.

Measuring the % of oxygen in air

We can show that air contains about 1/5th oxygen by reacting the oxygen with another substance to remove it from the air. The substance the oxygen is reacting with must be **in excess** so that **all the oxygen reacts** and is removed.

We would measure the volume of air before and after, and show the volume had decreased by the % of oxygen.

Using copper to measure %oxygen in air



When heated, the copper reacts with the oxygen to form solid black copper(II) oxide, removing the oxygen from the air. $2\ Cu_{(s)} + O_{2(g)} \rightarrow 2\ CuO_{(s)}$

e.g.

Volume of air before reaction: $= 200cm^3$

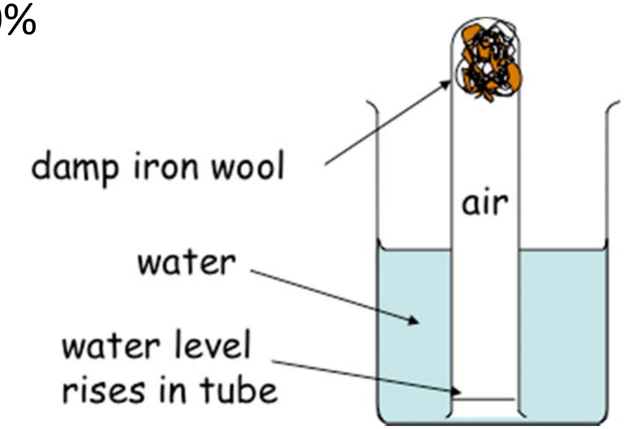
Volume of air after reaction: $= 160cm^3$

Amount of oxygen removed from air: $= 200 - 160 = 40cm^3$

%oxygen in the air: $= (40 \div 200) \times 100 = 20\%$

Using iron to measure %oxygen in air

When moist, the iron reacts with the oxygen to form rust, removing it from the air.



e.g.

Volume of air before reaction: $= 25cm^3$

Volume of air after reaction: $= 20cm^3$

Amount of oxygen removed from air: $= 25-20 = 5cm^3$

%oxygen in the air: $= (5 \div 25) \times 100 = 20\%$

Note: we can also do this experiment with phosphorus, but phosphorus is quite dangerous to handle.

Sources of error

Even if all the measurements are made correctly and there are no errors in the calculation, it is possible to get anomalous results.

Definition: Anomaly – a result that doesn't fit the pattern of other results.

We need to consider what the main ways in which our experiments could produce anomalous results, and what the effect of each of these sources of error might be. In doing so we will not consider human errors such as misreading measurements or incorrect calculations; we should assume that the experimenter can do the experiment correctly.

Examples:

Error: heating the copper for too short a period of time

Effect: not all the copper reacts, so not all the oxygen is removed from the air, so the volume of air after heating is too large, which makes the %oxygen too small.

Error: only using a small amount of iron (not in excess)

Effect: even when all the iron has reacted there is still some oxygen left in the test tube, so the volume of air after reaction is too large, which makes the %oxygen too small.

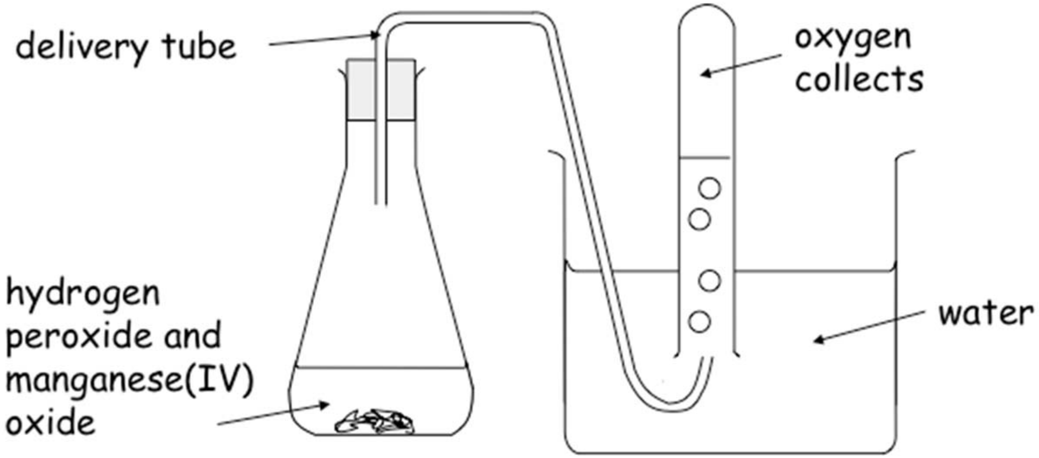
Error: doing the iron experiment at constant 40°C rather than room temperature

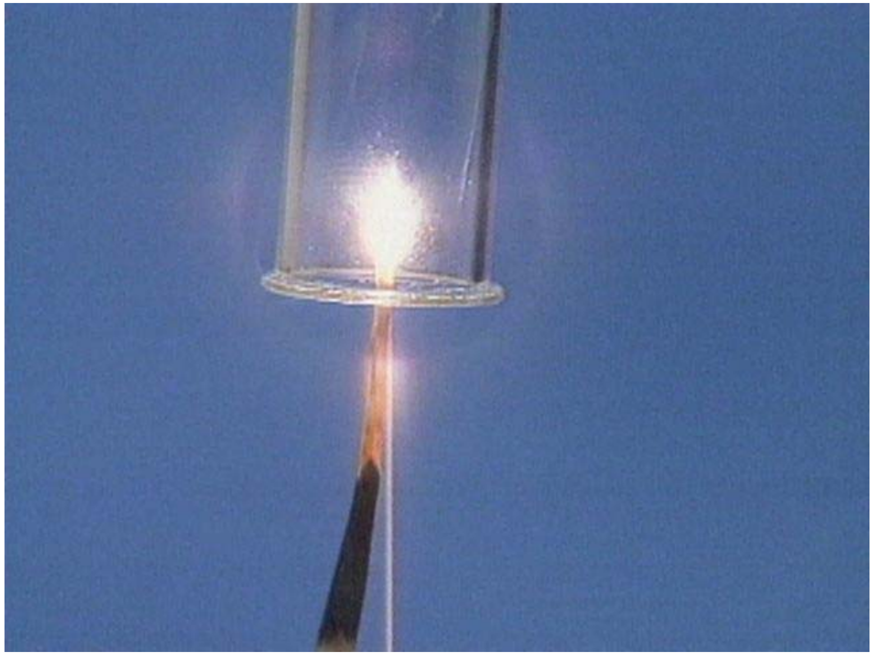
Effect: the rate of reaction is faster, but all the oxygen still reacts and the correct volume is removed from the air. This means the %oxygen will still be correct.

Making oxygen in the laboratory
Oxygen can be made by breaking down hydrogen peroxide solution, using **manganese(IV) oxide** as a catalyst. This type of reaction is called **catalytic decomposition**.

$$2\ H_2O_{2(aq)} \rightarrow 2\ H_2O_{(l)} + O_{2(g)}$$

hydrogen peroxide → water + oxygen





How the test for oxygen works:

We can show that a sample of gas is oxygen by lighting a splint, blowing it out so it is just glowing, then putting it into a tube of the gas. **Oxygen will re-light the glowing splint.**

The extra oxygen increases the rate of burning, generating enough heat to re-ignite a flame.

Reactions of oxygen – forming oxides

Oxygen will combine with many elements to form oxides. Some react slowly e.g. iron rusting.

iron + oxygen → iron oxide

Other reactions with oxygen are more rapid, releasing a lot of heat energy. This is what we mean by **burning**.

Burning is not just heating something up !

Example 1:

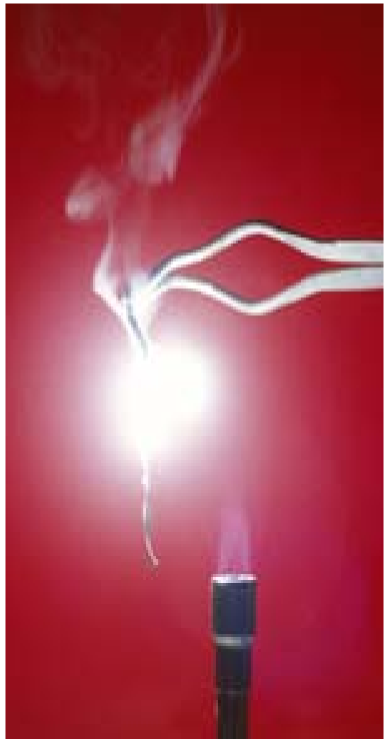
Magnesium burns with a bright white flame to produce powdery white magnesium oxide, which is a base.

$$2\ Mg_{(s)} + O_{2(g)} \rightarrow 2\ MgO_{(s)}$$

Like most metal oxides, magnesium oxide is almost insoluble in water. A small amount does dissolve to produce a slightly alkaline solution of magnesium hydroxide, indicating that magnesium is a metallic element.

$$MgO_{(s)} + H_2O_{(l)} \rightarrow Mg(OH)_{2(aq)}$$

The alkaline nature of the solution is due to the presence of hydroxide ions, OH^-.



Example 2:
Sulphur burns with a blue flame to produce colourless, poisonous sulphur dioxide gas.

$$S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$$

Like most non-metal oxides, sulphur dioxide dissolves in water, reacting to produce an acidic solution of sulphurous acid. This shows that sulphur is a non-metallic element.

$$SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$$

The acidic nature is due to the presence of H^+ ions in the solution.

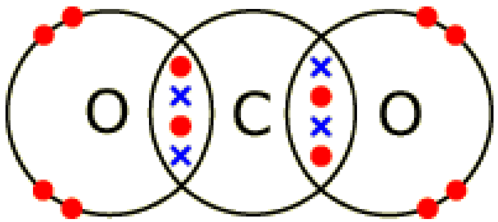
Example 3:
Carbon burns with a yellow-orange flame to produce colourless carbon dioxide gas. (The orange-yellow flame is due to particles of carbon (soot) produced as a result of incomplete combustion.)

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

Carbon dioxide dissolves to some extent in water to produce a slightly acidic solution. This shows that carbon is a non-metallic element.

Carbon Dioxide

Formula: CO_2
Bonding: covalent



dry ice is frozen carbon dioxide

Appearance: colourless, odourless gas
Density: denser than air

Carbon dioxide levels in the atmosphere are currently around 0.04% - but they were only 0.03% a few hundred years ago. This increase is due to our use of fossil fuels.

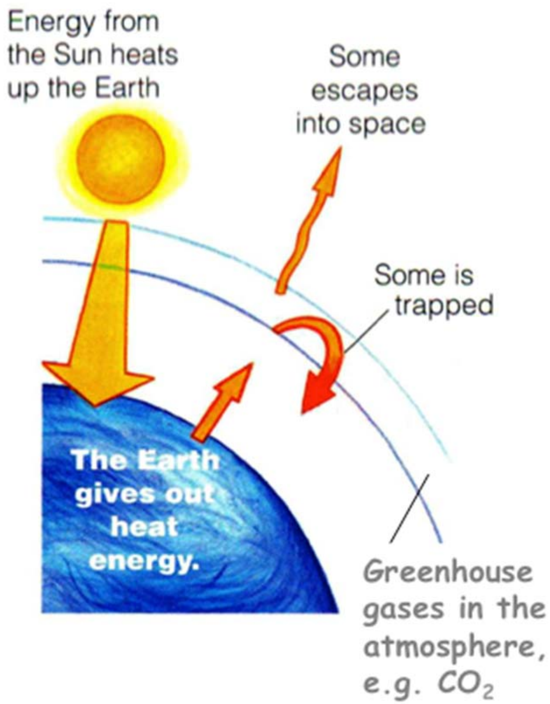
Role of carbon dioxide in climate change

Carbon dioxide is one of the most effective **greenhouse gases**.

When the surface of the Earth has been heated by the Sun, it re-radiates heat as infrared energy.

The greenhouse gases in the atmosphere absorb this infrared energy that would otherwise be lost into space. This keeps the Earth from having too cold a climate.

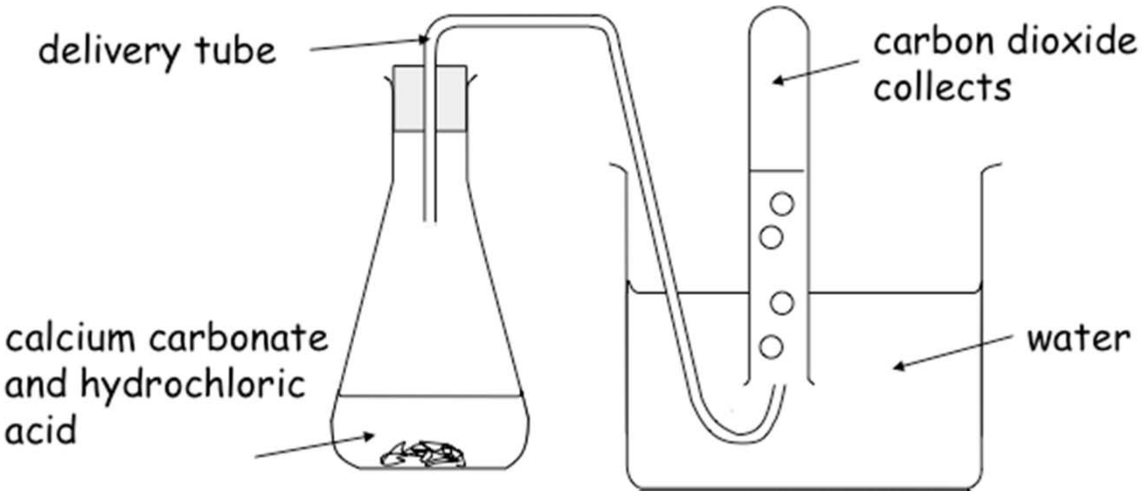
If the amount of greenhouse gases in the atmosphere gets too large, too much heat energy is retained. This absorbed heat energy is transferred to other molecules in the air, warming the whole atmosphere. This leads to **climate change**.



Making carbon dioxide in the laboratory

Carbon dioxide can be made by the reaction of hydrochloric acid with calcium carbonate.

$CaCO_{3(s)}$	+	$2HCl_{(aq)}$	→	$CaCl_{2(aq)}$	+	$H_2O_{(l)}$	+	$CO_{2(g)}$
calcium carbonate		**hydrochloric acid**		**calcium chloride**		**water**		**carbon dioxide**

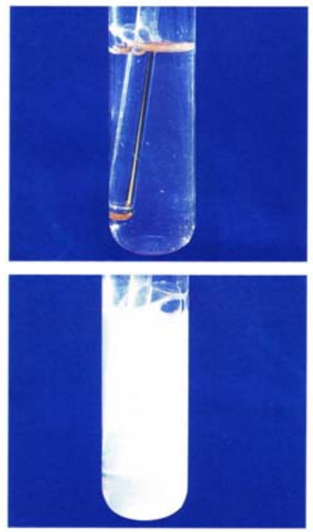


How the test for carbon dioxide works:

We can show that a sample of a gas is carbon dioxide by shaking it with (or bubbling it through) limewater (which is a solution of calcium hydroxide).

Carbon dioxide turns limewater milky by producing a white **precipitate of calcium carbonate**

$Ca(OH)_{2(aq)}$	+	$CO_{2(g)}$	→	$CaCO_{3(s)}$	+	$H_2O_{(l)}$
calcium hydroxide		**carbon dioxide**		**calcium carbonate**		**water**



Use of carbon dioxide in fizzy drinks

Carbon dioxide is slightly soluble and dissolves in water under pressure.

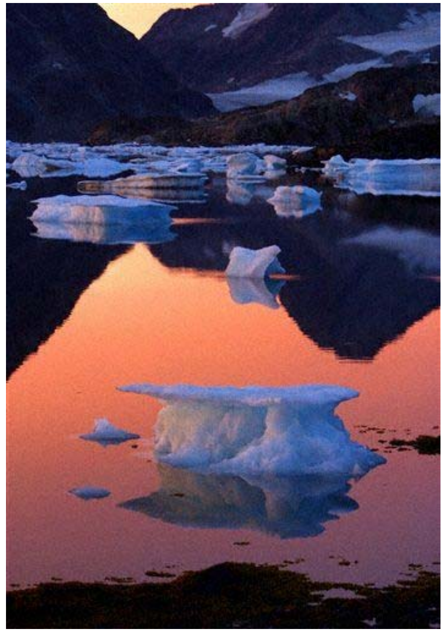
When the drink is opened and the pressure falls, bubbles of carbon dioxide come out of the solution making the drink fizzy.

The amount of a gas which dissolves depends on:

- the pressure – more dissolves at **higher pressure**
- the temperature – more dissolves at **lower temperature**

This is why fizzy drinks are kept cool, and why they go flat faster on a warm day.

Dissolved CO_2 in oceans



Oceans act as reservoirs for dissolved carbon dioxide, keeping down the level of carbon dioxide in the atmosphere.

When oceans get warmer, less CO_2 can be dissolved, so it ends up in the atmosphere.

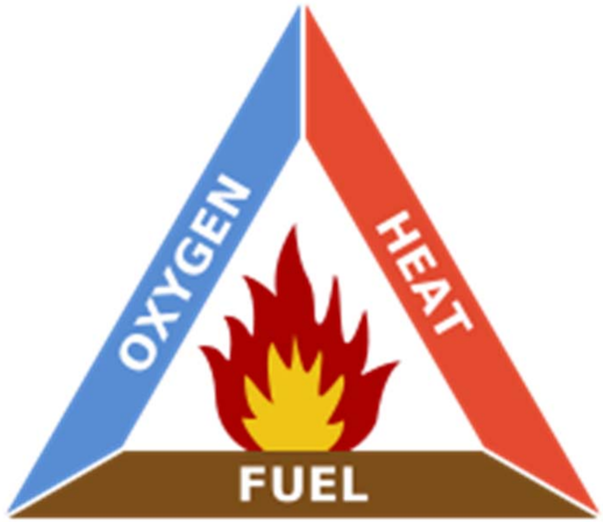
Increased atmospheric CO_2 increases global warming, increasing the temperature of the oceans…

Use of carbon dioxide in fire extinguishers

To continue burning, a fire needs:

- a supply of fuel
- a source of heat
- a supply of oxygen.

Taking any one of these away extinguishes the fire.



CO_2 is used to extinguish electrical or burning liquid fires, where using water could cause problems.

The dense CO_2 sinks onto the flames, keeping the oxygen from reaching the burning fuel.

Carbon dioxide from carbonates

Many metals form compounds which are carbonates

e.g.	sodium carbonate	Na_2CO_3
	magnesium carbonate	$MgCO_3$
	zinc carbonate	$ZnCO_3$

Limestone is composed mainly of calcium carbonate, $CaCO_3$

Limestone – calcium carbonate

These carbonates **thermally decompose** to produce the metal oxide and carbon dioxide when heated.

General equation:

metal carbonate → metal oxide + carbon dioxide

Note: You can tell a thermal decomposition reaction because there will only be one reactant (before the arrow).

Examples:

$$\mathbf{ZnCO_{3(s)} \rightarrow ZnO_{(s)} + CO_{2(g)}}$$

$$\mathbf{CuCO_{3(s)} \rightarrow CuO_{(s)} + CO_{2(g)}}$$

A Bunsen flame is **not hot enough** to decompose some, Group 1 carbonates e.g. potassium carbonate.

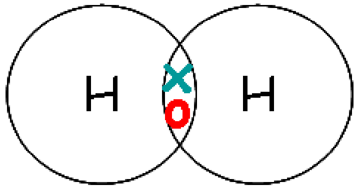
Practice: (answers at the end of the topic)

A sample of calcium carbonate weighing 2.50g was placed in a test tube and was heated strongly. It was then weighed again. After heating the sample weighed 2.15g.

i) Why did the sample weigh less ?

ii) What mass gas was produced in this reaction ?

Hydrogen
Formula: H_2
Bonding: covalent

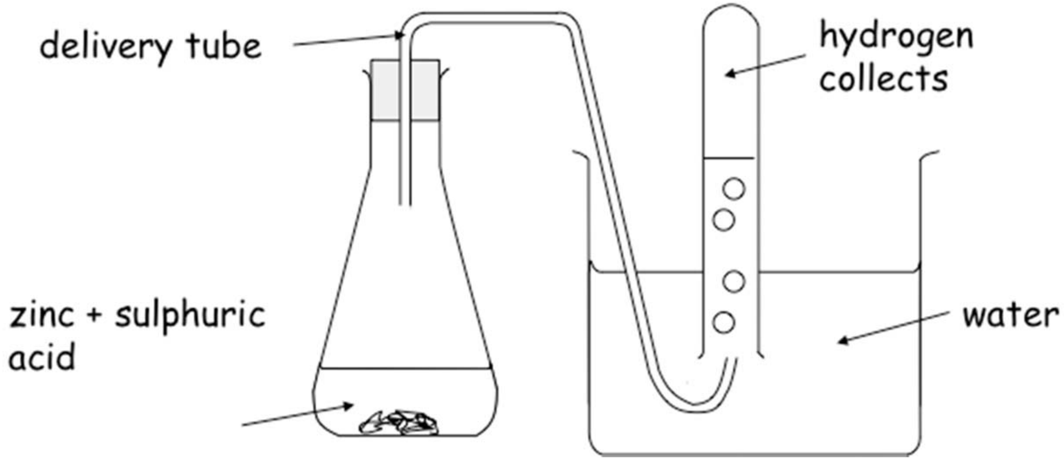


Why airships are not filled with hydrogen

Appearance: colourless, odourless gas – highly explosive
Density: much less dense than air

There is **no hydrogen** in our current atmosphere. Any hydrogen present when the Earth first formed escaped into space as it is such a ‘light’ (low density) gas.

Making hydrogen in the laboratory
Hydrogen can be made by the reaction of an acid such as hydrochloric or sulphuric acid with a reactive metal such as magnesium, zinc or iron. With zinc, we also add a little copper(II) sulphate to speed up the reaction.



e.g.

$$Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$$

zinc + sulphuric acid → zinc sulphate + hydrogen

How the test for hydrogen works:

We can show that a sample of gas is hydrogen by placing a lit splint into a test tube of the gas. If present, **hydrogen ignites with a squeaky 'pop'.**

The hydrogen reacts immediately with oxygen in the air, exploding to produce the 'pop' sound in the tube.

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

Use of hydrogen as a future fuel
Hydrogen can react with oxygen in a controlled combustion reaction which releases lots of energy. As a fuel it has less environmental impact than burning a fossil fuel.

The only product made when hydrogen is burned as fuel is water.

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$$

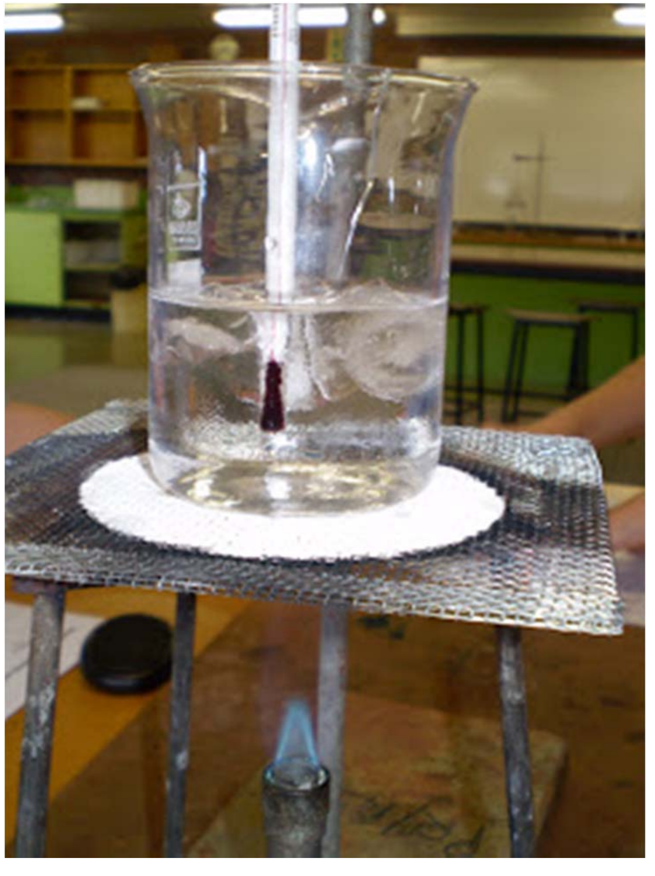
Practice:
Other than being 'clean', what else is needed for hydrogen to be a sustainable fuel for the future ? (answers at the end of the topic)

Tests for water

How can we prove that the liquid produced when hydrogen burns is pure water ?

We test to see if water is present by placing a few drops on some anhydrous copper sulphate in a watch glass.

If water is present, the white anhydrous copper sulphate turns blue.



We can determine if the water is pure by measuring its melting point and boiling point.

Pure water boils at exactly 100˚C
Pure water freezes at exactly 0˚C

Answers:

A sample of calcium carbonate weighing 2.50g was placed in a test tube and was heated strongly. It was then weighed again. After heating the sample weighed 2.15g.

i) Why did the sample weigh less ?

Because the calcium carbonate thermally decomposed, producing carbon dioxide gas, which escaped into the atmosphere, so wasn't weighed at the end of the experiment.

ii) What mass gas was produced in this reaction ?

The difference in mass is the mass of the escaped gas: 2.50 – 2.15 = 0.35g

Other than being 'clean', what else is needed for hydrogen to be a sustainable fuel for the future ?

There must sufficient supply of this fuel, accessible to all using it
It must safe to transport and store
There must be vehicles etc. designed to use this fuel

other answers are also possible, based on the idea that hydrogen as a fuel must be produced and made available to users without risk of pollution or explosion, excessive use of energy, or consumption of non-renewable resources.