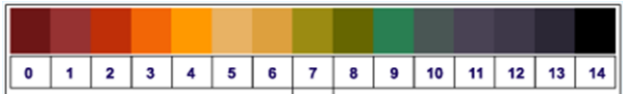
## pH and Acids + Alkalis




Acids produce $H^+$ (as $H_3O^+$) ions in water and they taste sour. They also corrode metals and have a pH of less than 7. They also turns blue litmus paper to red.

Alkalis produce $OH^-$ ions in water and they taste bitter with a pH greater than 7. Alkalis turns red litmus paper to blue.

A solution is defined as an acid if the concentration of $H^+$ ions is greater than the concentration of $OH^-$ ions. $[H^+] > [OH^-]$

A solution is defined as alkali/base if the concentration of hydrogen ions is less than the concentration of hydroxide ions. $[H^+] < [OH^-]$

## Strong and weak acids

A strong acid is completely ionised in aqueous solution. $HCl + H_2O \rightarrow H^+ + Cl^-$

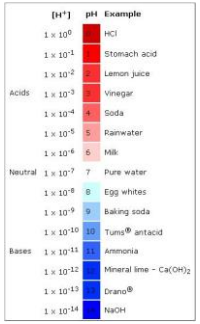
Examples of strong acids are hydrochloric, nitric and sulfuric acids.

A weak acid is only partially ionised in aqueous solution. $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H^+$

Examples of weak acids are ethanoic, citric and carbonic acids.

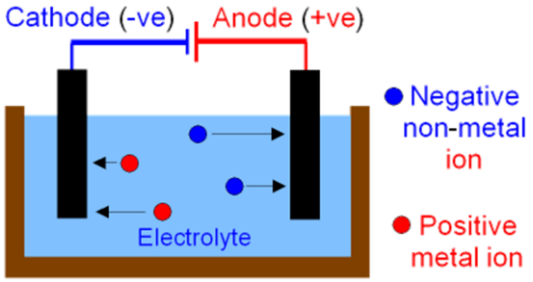
For a given concentration of aqueous solutions, the stronger an acid, the lower the pH.

As the pH decreases by one unit, the hydrogen ion concentration of the solution increases by a factor of 10.



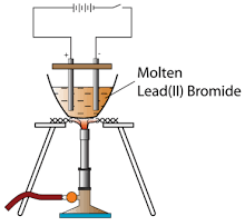

## Electrolysis

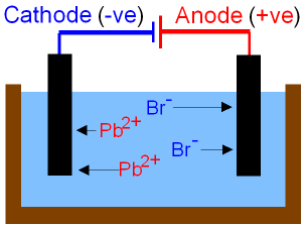
When an ionic compound is melted or dissolved in water, the **ions** are free to move about within the liquid or solution. These liquids and solutions are able to conduct electricity and are called electrolytes. Passing an electric current through electrolytes causes the ions to move to the electrodes. Positively charged ions move to the negative electrode (the cathode), and negatively charged ions move to the positive electrode (the anode).




## Electrolysis of molten ionic compounds

When a simple ionic compound (eg lead bromide) is electrolysed in the molten state using inert electrodes, the metal (lead) is produced at the cathode and the non-metal (bromine) is produced at the anode.




Cathode (-ve electrode)

$Pb^{2+} + 2e^- \rightarrow Pb$

Anode (+ve electrode)

$2Br^- - 2e^- \rightarrow Br_2$

OR $2Br^- \rightarrow Br_2 + 2e^-$

## Electrolysis Extended

At the negative electrode, hydrogen is produced if the metal is more reactive than hydrogen.
At the positive electrode oxygen is produced unless the solution contains halide ions when the halogen is produced.

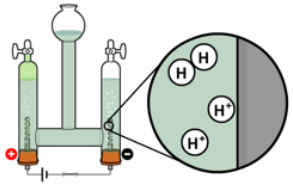
This is due to water molecules breaking down in aqueous solution to form hydrogen and hydroxide ions.

At the cathode positively charged ions gain electrons, whereas as the negatively charged ions lose electrons at the anode. These are both examples of oxidation and reduction. These can be represented as half equations.

$2H+ + 2e- \rightarrow H_2$

$4OH- \rightarrow O_2 + 2H_2O + 4e-$

$4OH- - 4e- \rightarrow O_2 + 2H_2O$




**At the cathode**

Whether hydrogen or a metal is produced at the cathode depends on the position of the metal in the metal **reactivity series**:

- the metal is produced at the cathode if it is less **reactive** than hydrogen
- hydrogen is produced at the cathode if the metal is more reactive than hydrogen

**Rules for determining products**

**At the anode**

Oxygen is produced (from hydroxide ions), unless **halide** ions (chloride, bromide or iodide ions) are present. In that case, the negatively charged halide ions lose electrons and form the corresponding **halogen** (chlorine, bromine or iodine).

The table summarises the product formed at the anode during the electrolysis of different **electrolytes** in solution.

| Negative ion | Element given off at anode |
|---|---|
| Chloride, $Cl^-$ | Chlorine, $Cl_2$ |
| Bromide, $Br^-$ | Bromine, $Br_2$ |
| Iodide, $I^-$ | Iodine, $I_2$ |
| Sulfate, $SO_4^{2-}$ | Oxygen, $O_2$ |
| Nitrate, $NO_3^-$ | Oxygen, $O_2$ |

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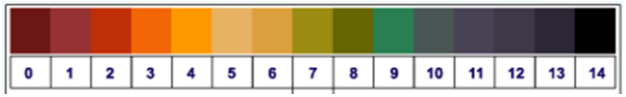


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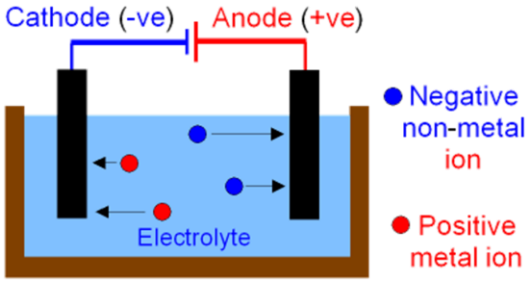


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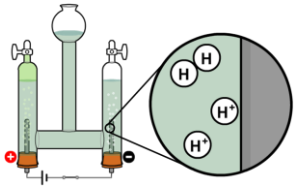
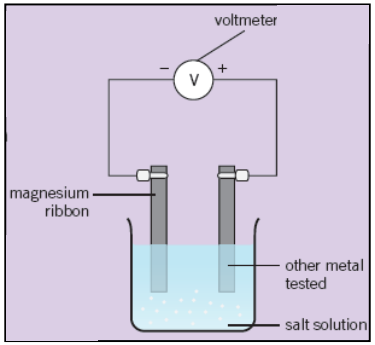



## Cells and batteries continued…

- Metals lose electrons and form positive ions.
- When 2 metals are dipped in a salt solution and joined by a wire, the more reactive metal will donate electrons to the less reactive metal. This forms a simple electrical cell.
- The greater the difference in reactivity between the 2 metals, the higher the voltage produced by the cell.

### Investigating chemical cells

This apparatus is used to investigate the voltage produced by different metals paired with magnesium ribbon. You can compare magnesium against zinc, iron, copper & tin in your electrical cells.




## Fuel Cells

Scientists are developing hydrogen as a fuel.

**Hydrogen + Oxygen → Water**

**$2H_2 + O_2 \rightarrow 2H_2O$**

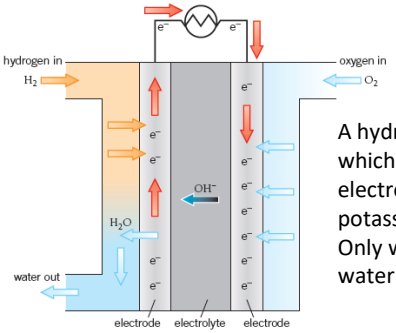
- The world relies on fossil fuels. However, they are non-renewable and they cause pollution.
- Hydrogen is one alternative fuel. It can be burned in combustion engines or used in fuel cells to power vehicles.
- Hydrogen gas is oxidised and provides a source of electrons in the hydrogen fuel cell, in which the only waste product is water.

Hydrogen gas is supplied as a fuel to the negative electrode. It diffuses through the graphite electrode and reacts with hydroxide ions to form water and provides a source of electrons to an external circuit.

$$2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$$

Oxygen is supplied to the positive electrode. It diffuses through the graphite and reacts to form hydroxide ions, accepting electrons from the external circuit.

$$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$




A hydrogen fuel cell which has an alkaline electrolyte, such as potassium hydroxide. Only waste product is water.

**Advantages of hydrogen fuel cells –**

1) Do not need to be electrically recharged
2) No pollutants are produced
3) Can be a range of sizes for different uses

**Disadvantages of hydrogen fuel cells–**

1) Hydrogen is highly flammable
2) Hydrogen is sometimes produced for the cell by non-renewable sources
3) Hydrogen is difficult to store