WJEC Wales Chemistry GCSE

2.3: Metals and their extraction

Detailed notes

Metal extraction

Metal reactivity

Metals are extracted from ores which are rocks found on the Earth's crust which contain metal compounds.

- Different methods can be used to extract the pure metal from ores
- The method used depends on the metal's position in the reactivity series
- The more reactive a metal is, the more stable its metal compound is so the harder it is to extract pure metal.
- Metals more reactive than carbon are usually extracted by a process called electrolysis.
- Metals less reactive than carbon can be extracted by reduction with carbon which is a cheaper process than electrolysis, so is favoured.
- Metals less reactive than hydrogen tend to be so unreactive they are often found pure in their native form.

The reactivity series

Most reactive metal, most stable metal compound

K	Potassium
Na	Sodium
Ca	Calcium
Mg	Magnesium
Al	Aluminium
C	Carbon
Zn	Zinc
Fe	Iron
Sn	Tin
Pb	Lead
H	Hydrogen
Cu	Copper
Ag	Silver
Au	Gold
Pt	Platinum

Least reactive metal, least stable metal compound

Displacement reactions

- A displacement reaction is when a more reactive metal displaces a less reactive metal from a compound
- These reactions can be used to investigate a metal's relative reactivity
- Some displacement reactions can be observed - for instance in the reaction between copper sulfate and magnesium, magnesium is more reactive than copper so displaces it and forms magnesium sulfate. Magnesium sulfate is colourless whereas copper sulfate solution is blue, so as the reaction proceeds the blue colour fades and you know a displacement reaction has occurred.
    - The chemical equation for this reaction is:
      Copper sulfate (aq) + magnesium → magnesium sulfate (aq) + copper

- Conversely, if copper was added to magnesium sulfate there would be no colour change as no reaction would occur
- The thermite reaction is a displacement reaction used in industry to produce pure iron
  - The balanced chemical equation is:
    $2Al\ (s) + Fe_2O_3\ (s) \rightarrow 2Fe\ (s) + Al_2O_3\ (s)$
  - The reaction is highly exothermic (it releases a lot of heat energy)

Reduction and oxidation

Reduction and oxidation can be defined in terms of loss/gain of electrons, oxygen and hydrogen.

	Oxidation	Reduction
Oxygen transfer	Gain of oxygen	Loss of oxygen
Electron transfer	Loss of electrons	Gain of electrons
Hydrogen transfer	Loss of hydrogen	Gain of hydrogen

OIL RIG is a handy mnemonic to remember oxidation in terms of electrons
Oxidation Is Loss (of electrons)
Reduction Is Gain (of electrons)

- A reaction in which reduction and oxidation both occur is called a redox reaction
- The species that is reduced is known as the oxidising agent
- The species that is oxidised is known as the reducing agent

Extraction of iron in the blast furnace

Iron ores contain iron oxide compounds such as haematite, Fe_2O_3. To obtain pure iron the oxygen must be removed from these compounds, meaning they get reduced. The reduction process is carried out in a blast furnace which contains the following materials:
- Iron ore
- Limestone, $CaCO_3$ - undergoes thermal decomposition to form calcium oxide, CaO, which reacts with impurities
- Coke - reacts with oxygen to form carbon monoxide which then reduces iron ore
- Oxygen - oxygen levels are controlled so carbon monoxide forms and not carbon dioxide, as oxygen reacts with the carbon in coke.

<u>The reactions</u>

1) $C\ (s) + O_2\ (g) \rightarrow CO_2\ (g)$
   Coke reacts with oxygen in the air to form carbon dioxide. This reaction is exothermic and heats up the furnace.

2) CO_2 (g) + C (s) → 2CO (g)
   The carbon dioxide reacts with more coke to form carbon monoxide. The carbon dioxide is reduced in this reaction.

3) $2Fe_2O_3$ (s) + 3C (s) → 4Fe (l) + $3CO_2$ (g)
   Fe_2O_3 (s) + 3CO → 2Fe (l) + $3CO_2$ (g)
   The iron oxide reacts with coke or carbon monoxide and is reduced to form molten iron.

Removing impurities with limestone

- The main impurity in the mixture is silicon dioxide (sand)
- Limestone undergoes thermal decomposition into calcium oxide which reacts with silicon dioxide to form solid calcium silicate which can be removed from the furnace
  $CaCO_3$ (s) → CaO (s) + CO_2 (g)
  CaO (s) + SiO_2 (l) → $CaSiO_3$ (s)
- This is a neutralisation reaction

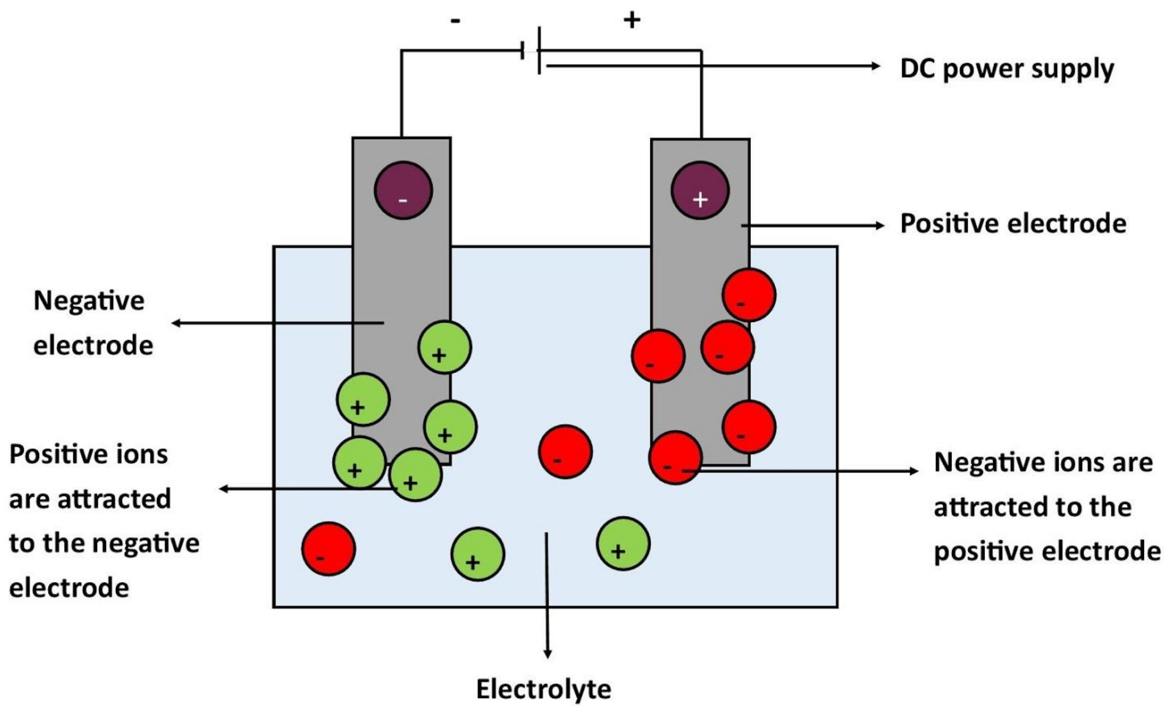
Electrolysis

The need for electrolysis

- Metals that are more reactive than carbon e.g aluminium are extracted by electrolysis of molten compounds.
  - Too reactive to be extracted by reduction with carbon
- Metals that are less reactive than carbon can be extracted by electrolysis as well
- Large amounts of energy are used in the extraction process to melt the compounds and to produce the electrical current
  - This makes it an expensive method, so for metals that can be extracted using reduction carbon this is used in preference

Electrolysis setup

- When a metallic compound is melted or dissolved, the ions are free to move about within the liquid or solution.
- Passing a current through substances that are molten or solution means that the solution can be broken down into elements. This is electrolysis, and the substance being broken down is the electrolyte.
- During electrolysis, positively charged ions move to the negative electrode (cathode), and negatively charged ions move to the positive electrode (anode).
  - A useful tool to remember the charge on the electrodes is PANiC which stands for Positive Anode Negative (i) Cathode
- Ions are discharged at the electrodes producing elements



Half equations

At the positive and negative electrodes ions transfer electrons to/from the electrode to form their uncharged element. These are oxidation and reduction reactions and can be represented using half equations.

- At the negative electrode positive ions gain electrons and are therefore reduced.
  - The ion will gain the same number of electrons as the magnitude of its charge
  - Example half equations for this:
    $Mg^{2+} + 2e^- \rightarrow Mg\ (s)$
    $Li^+ + e^- \rightarrow Li\ (s)$
- At the positive electrode negative ions lose electrons and are therefore oxidised.
  - The ion will lose the same number of electrons as the magnitude of its charge
  - Example half equations for this:
    $2Cl^- \rightarrow Cl_2 + 2e^-$
    $2O^{2-} \rightarrow O_2 + 4e^-$

Electrolysis of molten ionic compounds

If ionic compounds are molten it is much more simple to predict the products of electrolysis as there are no ions present except those in the ionic compound:

- Identify which ions there are within the ionic compound
- The + ions will go to the cathode
- The - ions will go to the anode
- EXAMPLE: Molten lead bromide
  - Pb^{2+} is positively charged so these ions move to the cathode, solid lead is produced and coats the cathode
  - The half equation for the reaction occurring at the cathode is:
    $Pb^{2+} + 2e^- \rightarrow Pb\ (s)$

- Br^- is negatively charged so these ions move to the anode, where 2 bromide ions lose 2 electrons and combine to form liquid bromine Br_2 (l)
- The half equation for the reaction occurring at the anode is:
  $2Br^- \rightarrow Br_2 + 2e^-$

Industrial electrolysis of aluminium

- Aluminium oxide is melted so electricity can pass through it
  - The melting point of aluminium oxide is very high which makes it expensive to melt
  - Aluminium oxide is dissolved in a substance called cryolite which lowers the melting point
  - The use of molten cryolite as a solvent reduces some of the energy costs involved in extracting aluminium
- The negative electrode (cathode) and positive electrode (anode) are made of graphite, a form of carbon
- Aluminium metal forms at the negative electrode and sinks to the bottom of the tank and it is tapped off here
- Oxygen forms at the positive electrode and it reacts with the carbon forming carbon dioxide, which bubbles out of the tank
  - As the electrode itself reacts with oxygen the positive electrode gradually burns away
  - Therefore, the positive electrode has to be replaced often, adding to the cost of the process

Electrolysis of water

Electrolysis can be used to separate water into hydrogen gas, H_2 (g), and oxygen gas, O_2 (g).

- The overall equation for this reaction is:
  $2H_2O\ (l) \rightarrow 2H_2\ (g) + O_2\ (g)$
- Positive hydrogen ions move to the negative electrode (cathode) and gain electrons in a reduction reaction.
  - The half equation is:
    $4H^+\ (aq) + 4e^- \rightarrow 2H_2\ (g)$
- Negative OH^- ions move to the positive electrode (anode) and lose electrons to form water and oxygen in an oxidation reaction.
  - The half equation is:
    $4OH^-\ (aq) \rightarrow 2H_2O\ (l) + O_2\ (g) + 4e^-$

HIGHER TIER ONLY - Electrolysis of aqueous solutions

- **When you have an ionic solution (NOT a molten ionic compound), your solution will contain: the ions that make up the ionic compound, and the ions in water (OH^- and H^+)**
- **At the cathode (-):**
  - **Hydrogen (from H^+ in water) is produced UNLESS the + ions in the ionic compound are from a metal less reactive than hydrogen**
  - **If the metal is less reactive, it will be produced instead**
  - **The relative reactivities is shown in the reactivity series**

- **At the anode (+):**
  - **Oxygen (from OH^- in water) will be produced UNLESS the ionic compound contains halide ions (Cl^-, Br^-, I^-)**
  - **If there are halide ions, the halogen will be produced instead (e.g. Cl_2)**

Examples:

- Copper chloride solution
  - Cu^+ ions go to cathode, Cu (s) is produced (Cu is less reactive than hydrogen)
  - Cl^- ions go to anode, Cl_2 (g) is produced (Cl^- are halide ions)
- Sodium chloride solution
  - H^+ ions go to cathode, H_2 (g) is produced (Na is more reactive than hydrogen)
  - Cl^- ions go to anode, Cl_2 (g) is produced (Cl^- are halide ions)
- Sodium sulfate solution
  - H^+ ions go to cathode, H_2 (g) is produced (Na is more reactive than hydrogen)
  - OH^- ions go to anode, O_2 (g) is produced (SO_4^{2-} ions are not halide ions)
- Water acidified with sulfuric acid
  - H^+ to cathode, H_2 (g) is produced (these are the other ions present in sulfuric acid H_2SO_4)
  - OH^- ions go to anode, O_2 (g) is produced (SO_4^{2-} ions are not halide ions)

Uses of electrolysis

Electroplating

- Electrolysis can be used to cover the surface of one metal with another metal, such as when jewellery is coated in silver to make it silver-plated.
- Requirements:
  - The negative electrode is the object to be electroplated
  - The electrolyte must contain ions of the metal you want to cover the negative electrode in
  - The positive electrode is the metal you want to electroplate the negative electrode with
- Example - if you wanted to electroplate a piece of metal such as a fork with silver:
  - Electrolyte: needs to contain silver ions - silver nitrate solution
  - Positive electrode: pure silver metal
  - Negative electrode: the metal fork to be electroplated

HIGHER TIER ONLY - Purification of copper

Copper can be purified by electrolysis.

- **The set up:**
  - **The positive electrode is a rod of impure copper**
  - **The negative electrode is a rod of pure copper**
  - **The electrolyte contains copper ions from a solution like copper sulfate solution**
- **During electrolysis, copper ions from the positive electrode of impure copper dissolve and are deposited on the negative electrode**
- **The result is that the pure copper electrode grows in size**

- **The half equations are:**
    - **At the positive electrode - $Cu \rightarrow Cu^{2+} + 2e^-$ (oxidation)**
    - **At the negative electrode - $Cu^{2+} + 2e^- \rightarrow Cu$ (reduction)**

<u>HIGHER TIER ONLY - Manufacture of sodium hydroxide</u>

In the electrolysis of sodium chloride solution:

- **H^+ ions go to cathode, H_2 (g) is produced (Na is more reactive than hydrogen so hydrogen gas forms at the negative electrode)**
- **Cl^- ions go to anode, Cl_2 (g) is produced (Cl^- are halide ions)**
- **Sodium ions and hydroxide ions remain in the electrolyte, giving sodium hydroxide solution.**

Metals

<u>Properties of certain metals</u>

- Aluminium
    - Low density
    - Lightweight for their size
    - Very thin layer of their oxides on the surface, which stops air and water getting to the metal, so resists corrosion
    - Used for aircraft, trains, overhead power cables, saucepans and cooking foil
- Copper
    - Good conductor of electricity and heat
    - Soft, easily bent and shaped (i.e. malleable)
    - Resistant to corrosion (very unreactive)
    - Electrical wiring, gas pipes and water pipes, plumbing in houses (does not react with water)
- Iron
    - Malleable
    - An alloy is formed of iron and carbon
        - Steel – harder and stronger than iron and less likely to rust
        - Used to build cars
        - Steel is used in the construction industry
- Titanium
    - Low density
    - Lightweight for their size
    - Very thin layer of their oxides on the surface, which stops air and water getting to the metal, so resists corrosion
    - Fighter aircraft, artificial hip joints and pipes in nuclear power stations

<u>General properties of transition metals</u>

- High melting points
- Form coloured compounds
- Have the ability to form ions with different charges
    - For instance, iron can form Fe^{3+} and Fe^{2+} ions
- Most are malleable and ductile

- Good conductors of both heat and electricity due to their delocalised electrons
- Hard
- Less reactive than alkali metals (group 1 metals)

HIGHER TIER ONLY - Test for metal ions

When sodium hydroxide (NaOH) is added:

- **Copper(II) produces a blue precipitate**
- **Iron(II) produces a green precipitate**
- **Iron(III) produces a brown precipitate**

Alloys

- An alloy is a substance made of a mixture of 2 or more elements, of which at least one is a metal
- The composition of alloys can be changed to produce alloys with desired properties
- An example is bronze, which is a mixture of copper and tin
- Alloys are made by mixing molten metals

Evaluation of extraction processes

There are a number of factors that must be considered when extracting metals

- Site of extraction plants -
  - In areas with good roads and railway lines to allow for transport of the extracted metal
  - Near a town or city so workers can be sources
  - Away from built-up areas due to noise and pollution
  - Close to a power station so large amounts of energy can be supplied
- Method used -
  - Where possible lower energy and cost methods such as reduction with carbon should be used
    - Not possible with metals more reactive than carbon
- Recycling -
  - The possibility of using recycling plants to recycle and reuse materials as opposed to extract new raw materials