WJEC England GCSE Chemistry

Topic 6: Reactivity series and extractions of metals

Notes

(Content in bold is for Higher Tier only)

Reactions of metals

- A few reactive metals will react with cold water
    - Products are a metal hydroxide (forming an alkaline solution) and hydrogen gas
    - E.g. with potassium: $2K + 2H_2O \rightarrow 2KOH + H_2$
- More metals react with acids
    - metal + acid → salt + hydrogen
- Almost all metals react with oxygen
    - metal + oxygen → metal oxide
- Only metal that does not react with any of the above is gold, because it is extremely unreactive
- You can therefore deduce the relative reactivity of some metals by seeing if they react with water (i.e. VERY reactive), acid (reactive), and oxygen (not that reactive)
- You can see if one metal is more reactive than another by using displacement reactions...
    - Easily seen when a salt of the less reactive metal is in the solution
        - More reactive metal gradually disappears as it forms a solution
        - Less reactive metal coats the surface of the more reactive metal
    - e.g. When you place a steel (iron) nail in a $CuCl_2$ solution, the iron first displaces the copper and forms $FeCl_2$, because iron is more reactive than copper.
    - e.g. thermite reaction:
        - aluminium + iron(III) oxide → iron + aluminium oxide
        - Because aluminium is more reactive than iron, it displaces iron from iron(III) oxide. The aluminium removes oxygen from the iron(III) oxide:

The reactivity series

- When metals react with other substances, metal atoms form positive ions
- Reactivity of a metal is related to its tendency to form positive ions
- Metals can be arranged in order of their reactivity in a reactivity series
    - Metals potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper can be put in order of their reactivity from their reactions with water and dilute acids
        - Metal + acid -> salt + hydrogen
    - Non-metals hydrogen and carbon are often included in the reactivity series

Extraction by heating with carbon (including iron)

- any metals less reactive than carbon can be extracted from their oxides by reduction with carbon (reduction= loss of oxygen)
- extraction of iron from iron ore:

- Iron oxide loses oxygen, and is therefore reduced. The carbon gains oxygen, and is therefore oxidised.
- $2Fe_2O_3(s) + 3C(s) \rightarrow 4Fe(l) + 3CO_2(g)$
- In the blast furnace, it is so hot that carbon monoxide can be used, in place of carbon, to reduce the iron(III) oxide:
    - Iron(III) oxide + carbon monoxide -> iron + carbon dioxide
    - $Fe_2O_3(s) + 3CO(s) \rightarrow 2Fe(l) + 3CO_2(g)$
- Limestone/calcium carbonate helps to remove acidic impurities from the iron by reacting with them to form molten slag

Oxidation and reduction

- oxidation and reduction in terms of oxygen:
    - oxidation is gain of oxygen
    - reduction is loss of oxygen
- **oxidation and reduction in terms of electrons:**
    - **oxidation is loss of electrons**
    - **reduction is gain of electrons**
    - **Try and remember this phrase: OIL RIG, it stands for Oxidation Is Loss and Reduction Is Gain (of electrons)**

The process of electrolysis

- When an ionic substance 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).
- Ions are discharged at the electrodes producing elements, this process is called electrolysis

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

- Molten lead bromide
    - Pb^{2+} to cathode, Pb (s) is produced (not in solution so these are the only + ions present)

- Br^- to anode, Br_2 (l) is produced (not in solution so these are the only - ions present)

Using electrolysis to extract metals

- 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 react with 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

Aluminium example:

- Aluminium oxide is melted so that electricity can pass through it
  - Expensive to melt it due to the very high melting point of aluminium oxide
  - Instead, it is dissolved in molten cryolite – an aluminium compound with a lower melting point than aluminium oxide
  - The use of molten cryolite as a solvent reduces some of the energy costs involved in extracting aluminium
- 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, therefore the positive electrode gradually burns away
  - Therefore, the positive electrode has to be replaced often, adding to the cost of the process

Electrolysis of aqueous solutions

- When you have a 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
- 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)

- Electrolysis of:
  - 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^- to anode, O_2 (g) is produced (SO_4^{2-} ions are not halide ions)

Alternative methods of metal extraction

- **Phytoextraction**
  - **Some plants absorb metal compounds through their roots**
  - **They concentrate these compounds as a result of this**
  - **The plants can be burned to produce an ash that contains the metal compounds**
- **Bacterial metal extraction**
  - **Some bacteria absorb metal compounds**
  - **Produce solutions called leachates which contain the metal**

Properties of certain metals

- 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 with 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

Tests for metal ions

- when NaOH (sodium hydroxide is added)
- Copper(II) produces a blue precipitate
- Iron(II) produces a green precipitate
- Iron(III) produces a brown precipitate

Practical assessments

NB: be prepared to carry out these assessments

- SP6A Determination of relative reactivities of metals through displacement reactions
- SP6B Investigation into electrolysis of aqueous solutions and electroplating