Using Magnesium to Extract titanium

Titanium is a very useful metal because it is abundant, has a low density and is corrosion resistant – it is used for making strong, light alloys for use in aircraft for example.

Titanium cannot be extracted with carbon because titanium carbide (TiC) it is formed rather than titanium .
Titanium cannot be extracted by electrolysis because it has to be very pure.

Titanium is extracted by reaction with a more reactive metal (e.g. Mg).

Steps in extracting titanium

1. TiO_2 (solid) is converted to $TiCl_4$ (liquid) at 900°C:
2. The $TiCl_4$ is purified by fractional distillation in an argon atmosphere.
3. The Ti is extracted by Mg in an argon atmosphere at 500°C

$$TiO_2 + 2\,Cl_2 + 2\,C \rightarrow TiCl_4 + 2\,CO$$

$$TiCl_4 + 2Mg \rightarrow Ti + 2\,MgCl_2$$

Titanium is expensive because

1. The expensive cost of the Mg
2. This is a batch process which makes it expensive because the process is slower (having to fill up and empty reactors takes time) and requires more labour and the energy is lost when the reactor is cooled down after stopping
3. The process is also expensive due to the argon, and the need to remove moisture (because $TiCl_4$ is susceptible to hydrolysis).
4. High temperatures required in both steps

TiO_2 is converted to $TiCl_4$ as it can be purified by fractional distillation, $TiCl_4$ being molecular (liquid at room temperature) rather than ionic like TiO_2 (solid at room temperature).

This all makes titanium expensive even though it is a relatively abundant metal.
It is only therefore used to a limited amount even though it has useful properties.

Solubility of Hydroxides

Group II hydroxides become **more soluble** down the group.

All Group II hydroxides when not soluble appear as white precipitates.

Magnesium hydroxide is classed as insoluble in water.

Simplest Ionic Equation for formation of $Mg(OH)_2$ (s)

$$Mg^{2+} (aq) + 2OH^{-}(aq) \rightarrow Mg(OH)_2 (s).$$

A suspension of magnesium hydroxide in water will appear slightly alkaline (pH 9) so some hydroxide ions must therefore have been produced by a very slight dissolving.

Magnesium hydroxide is used in medicine (in suspension as milk of magnesia) to neutralise excess acid in the stomach and to treat constipation.

$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$

It is safe to use because it is so weakly alkaline. It is preferable to using calcium carbonate as it will not produce carbon dioxide gas.

Calcium hydroxide is classed as partially soluble in water and will appear as **a white precipitate** It is used in agriculture to neutralise acidic soils.

A suspension of calcium hydroxide in water will appear more alkaline (pH 11) than magnesium hydroxide as it is more soluble so there will be more hydroxide ions present in solution.

An aqueous solution of calcium hydroxide is called lime water and can be used a test for carbon dioxide. The limewater turns cloudy as white calcium carbonate is produced.

$Ca(OH)_{2\,(aq)} + CO_{2\,(g)} \rightarrow CaCO_{3\,(s)} + H_2O_{(l)}$

Barium hydroxide would easily dissolve in water. The hydroxide ions present would make the solution strongly alkaline.

$Ba(OH)_{2\,(S)} + aq \rightarrow Ba^{2+} (aq) + 2OH^{-}(aq)$

Solubility of Sulfates

Group II sulfates become **less soluble** down the group.

$BaSO_4$ is the least soluble.

An equation for the formation of the precipitate can be written as a full equation or simplest ionic equation

Full equation : $SrCl_2(aq) + Na_2SO_4 (aq) \rightarrow 2NaCl (aq) + SrSO_4 (s)$

Ionic equation: $Sr^{2+} (aq) + SO_4^{2-}(aq) \rightarrow SrSO_4 (s).$

$BaSO_4$ is used in medicine as a ‘Barium meal’ given to patients who need x-rays of their intestines. The barium absorbs the x-rays and so the gut shows up on the x-ray image. Even though barium compounds are toxic, it is safe to use here because its low solubility means it is not absorbed into the blood.

If barium metal is reacted with sulfuric acid it will only react slowly as the insoluble barium sulfate produced will cover the surface of the metal and act as a barrier to further attack.

$Ba + H_2SO_4 \rightarrow BaSO_4 + H_2$

The same effect will happen to a lesser extent with metals going up the group as the solubility increases.

The same effect does not happen with other acids like hydrochloric or nitric as they form soluble group 2 salts.

Testing for Presence of a Sulfate ion

$BaCl_2$ solution acidified with hydrochloric acid is used as a reagent to test for sulphate ions.

If acidified **barium chloride** is added to a solution that contains sulfate ions a **white precipitate** of barium sulfate forms.

Other anions should give a negative result which is no precipitate forming.

Simplest ionic equation

$Ba^{2+} (aq) + SO_4^{2-}(aq) \rightarrow BaSO_4 (s).$

The hydrochloric acid is needed to react with carbonate impurities that are often found in salts which would form a white barium carbonate precipitate and so give a false result. You could not use sulfuric acid because it contains sulfate ions and so would give a false positive result.

$2HCl + Na_2CO_3 \rightarrow 2NaCl + H_2O + CO_2$

Fizzing due to CO_2 would be observed if a carbonate was present.

More on Insoluble salts and Precipitation reactions

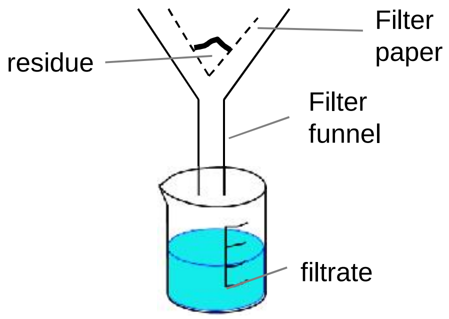
Insoluble salts can be made by mixing appropriate solutions of ions so that a **precipitate** is formed
barium nitrate (aq) + sodium sulfate (aq) → **barium sulfate (s)** + sodium nitrate (aq)
These are called **precipitation** reactions. A **precipitate is a solid**

There are some common rules for solubility of salts. No syllabus requires these to be learnt but a good chemist does know them.

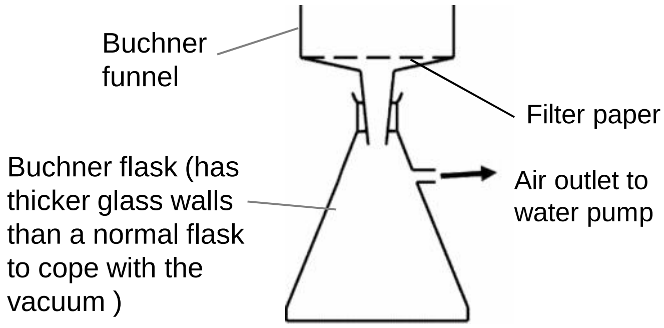
Soluble salts	Insoluble salts
All sodium, potassium and ammonium salts	
All nitrates	
Most chlorides, bromides, iodides	Silver, lead chlorides, bromides iodides
Most sulfates	Lead, strontium and barium sulfates
Sodium, potassium and ammonium carbonates	Most other carbonates
Sodium, potassium and ammonium hydroxides	Most other hydroxides

When making an insoluble salt, normally the salt would be removed by **filtration,** washed with **distilled water to remove soluble impurities** and then **dried on filter paper**

Filtration



This is gravitational filtration. Use if small amounts of solid are formed.



This is vacuum filtration. The apparatus is connected to a water pump which will produce a vacuum. Use if larger amounts of solid are formed.

For both types of filtration apparatus AQA expect filter paper to be drawn on the diagram

Writing Ionic equations for precipitation reactions

We usually write ionic equations to show precipitation reactions. Ionic equations only show the ions that are reacting and leave out spectator ions.

Spectator ions are ions that are
- Not changing state
- Not changing oxidation number

Take full equation $Ba(NO_3)_2\ (aq) + Na_2SO_4\ (aq) \rightarrow BaSO_4\ (s) + 2\ NaNO_3\ (aq)$

Separate (aq) solutions into ions $Ba^{2+}{}_{(aq)} + 2NO_3^-{}_{(aq)} + 2Na^+{}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)} + 2\ Na^+{}_{(aq)} + 2NO_3^-{}_{(aq)}$

Cancel out spectator ions leaving the simplest ionic equation $Ba^{2+}\ (aq) + SO_4^{2-}(aq) \rightarrow BaSO_4\ (s)$.