

METALS

IMPORTANCE OF METALS

More than three quarters of elements in the periodic table are metals.
We cook food, heat the food on metal stoves.
We travel in metallic vehicles like cars, airplanes, ships.
The concrete walls and roofs of houses are reinforced with metal rods.
Coins, taps, cutlery, pins, needles, paper pins etc. are all made up of metals.

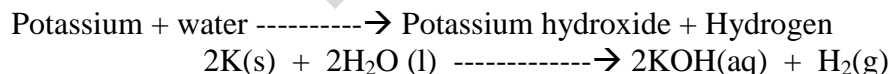
PHYSICAL PROPERTIES OF METALS

PROPERTY	REASON
Metals are malleable and ductile i.e. they can be drawn into sheets and wires	Layers of metals can slide over each other.
Metals conduct electricity.	They have free moving electrons in their outer most shell.
Metals usually have high melting points, high boiling points and high densities	Strong metallic bond

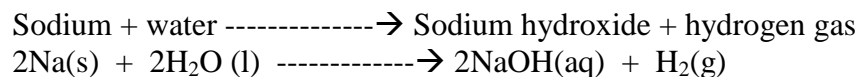
CHEMICAL PROPERTIES OF METALS

Reaction with water

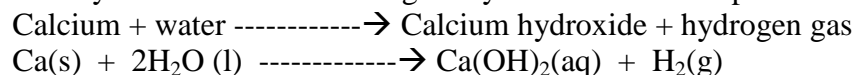
Potassium reacts vigorously with cold water to form salt and hydrogen gas. The reaction is so exothermic that the hydrogen gas produced, burns in air.



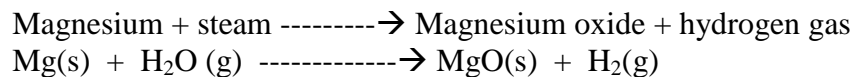
Sodium reacts with cold water in the same way.



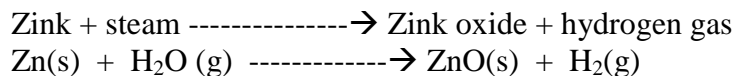
Calcium reacts readily with cold water and vigorously with hot water to produce salt and hydrogen gas.



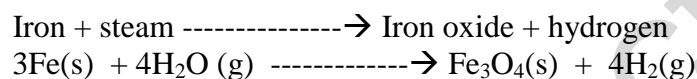
Magnesium reacts very slowly with cold water but vigorously with steam to form salt and hydrogen gas.



Zinc do not react with cold water but reacts slowly with steam to form zinc oxide and hydrogen gas.



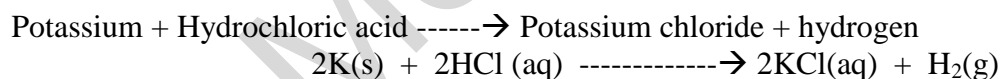
Iron do not react with cold water but rusting occur very slowly in the presence of oxygen. Red hot iron reacts very slowly with steam to produce salt and hydrogen gas.



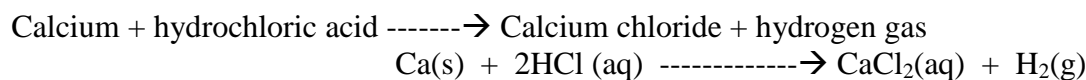
Copper do not react with water under any condition
Silver do not react with water in any condition.

Reaction with Hydrochloric acid

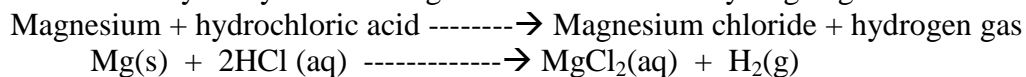
Potassium and sodium reacts explosively to form salt and hydrogen gas. The reaction is so exothermic that the hydrogen gas produced, burn in air.



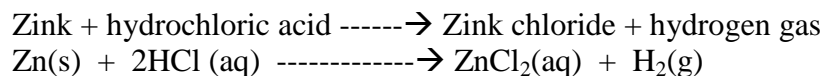
Calcium reacts vigorously` to produce calcium chloride and hydrogen gas.



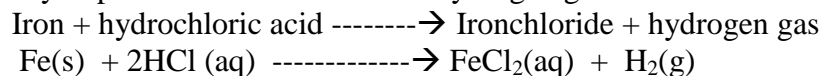
Magnesium reacts very fastly to form magnesium chloride and hydrogen gas.



Zinc reacts moderately to form zinc chloride and hydrogen gas.



Iron reacts slowly to produce iron chloride and hydrogen gas.



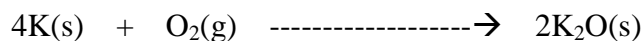
Copper do not react with dilute HCl

Silver do not react with dilute HCl

Reaction with oxygen

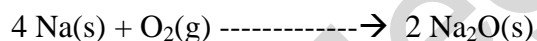
Potassium tarnishes in the presence of oxygen to form potassium oxide K_2O

Potassium + oxygen \longrightarrow Potassium oxide



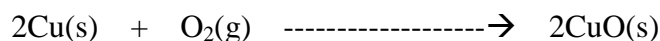
Sodium burns with a yellow flame to produce odium oxide Na_2O

Sodium + Oxygen \longrightarrow Sodium Oxide



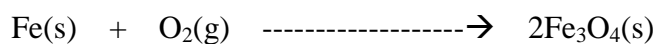
Copper powder burns with dull red glow to form copper oxide. CuO

Copper + Oxygen \longrightarrow Copper oxide



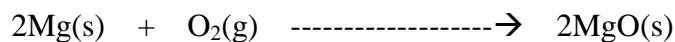
Iron powder or wire burns with a bright yellow flame to form iron oxide Fe_3O_4

Iron + Oxygen \longrightarrow Iron oxide



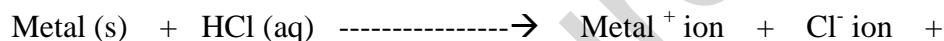
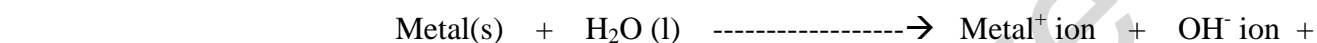
Magnesium burns with a bright white flame to produce white solid magnesium oxide. MgO

Magnesium + oxygen -----→ Magnesium oxide



REACTIVITY SERIES

Metals can be arranged in order of their chemical reactivity. The reactivity series is based on the reaction of metals with water or dilute hydrochloric acid. When metal reacts with acid or water, the metal atom loses an electron to become an ion.



The more readily a metal gives up electrons to form ions, the more reactive it is.

A metal high up in the reactivity series

Reacts vigorously with chemicals

Readily gives up electrons in reactions to form positive ions

Corrode easily

A metal low down in the reactivity series

Does not Reacts vigorously with chemicals

Does not Readily gives up electrons in reactions to form positive ions

Does not Corrode easily

Hydrogen is sometimes placed in the reactivity series. Metals below hydrogen in the series do not react with acids to produce hydrogen gas.

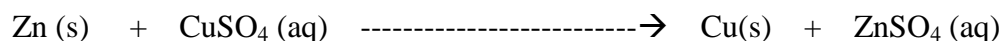
Please -----→ Potassium
Send -----→ Sodium
Cats -----→ Calcium
Monkeys -----→ Magnesium
And -----→ Aluminium
Zebras -----→ Zinc
In -----→ Iron
Large -----→ Lead
Hired -----→ Hydrogen
Cages -----→ Copper
Make -----→ Mercury
Sure -----→ Silver
Good -----→ Gold
Padlock -----→ Platinum

DISPLACEMENT OF METALS

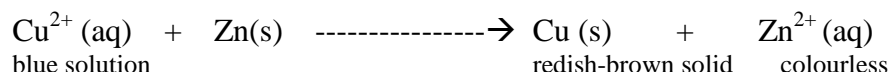
Displacement of metals from solutions

A more reactive metal will displace the ions of any less reactive metal in the reactivity series, from solution.

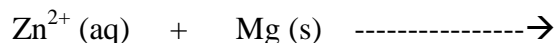
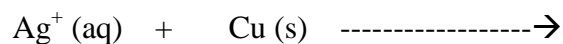
Zinc + copper (II) sulphate solution -----→ Copper + zinc sulphate solution.



Zinc displace copper from the copper sulphate solution because it is more reactive than copper and readily give up electrons to form positive ions. The electrons are transferred from zinc atom to copper (II) ions.

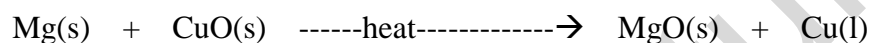
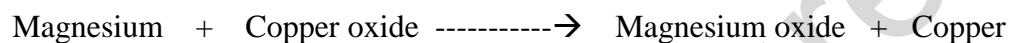


Other examples:

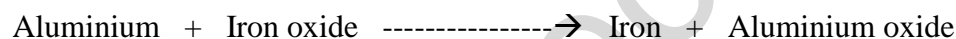


Displacement of metals from metallic oxides by more reactive metals

A metal will take oxygen from the oxide of any metal below it in the reactivity series. For example, when magnesium powder and copper (II) oxide powder is heated there is a vigorous exothermic reaction. The magnesium takes oxygen from copper (II) oxide to form magnesium oxide and copper metal.

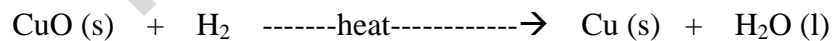
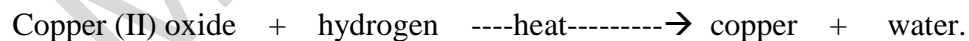
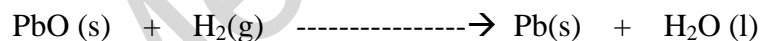
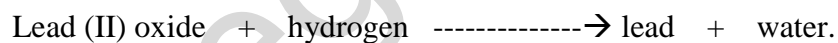


Thermite reaction reaction.



Reaction of metallic oxides with hydrogen

Hydrogen can take oxygen from metallic oxides, producing the metal and water. For example when hydrogen is passed over hot lead (II) oxide, lead metal and water are produced.

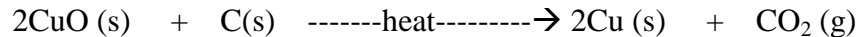


The less reactive the metal, the easier it is for hydrogen to take oxygen from its oxide. The oxides of very reactive metals such as aluminium oxide and sodium oxide cannot be reduced to the metal by hydrogen.

Reaction of metallic oxides with carbon.

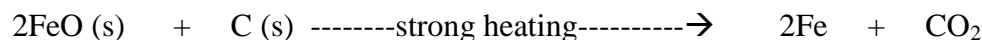
Carbon can take up oxygen from the oxide of metals which are not too high in the reactivity series. For example a mixture of charcoal and copper (II) oxide reacts when heated together

Copper (II) oxide + Carbon $\xrightarrow{\text{heat}}$ copper + carbon dioxide.



The more reactive the metal the more harder it for carbon to take oxygen from its oxide. Iron is more reactive than copper, iron oxide and carbon must be heated very strongly for the reaction to take place.

Iron (II) oxide + Carbon $\xrightarrow{\text{strong heating}}$ Iron + carbon dioxide.



Carbon is unable to take oxygen from the oxides of very reactive such as calcium and sodium.

THE EXTRACTION OF METALS

Most of the metals are found as compounds called minerals. Minerals are usually found mixed with large amounts of impurities. These impure minerals are called ores.

A **ROCK** is a mixture of minerals from which useful substances can be made.

A **MINERAL** is a solid element or compound found naturally in the Earth's crust.

A **METAL ORE** is a mineral or mixture of minerals from which economically viable amounts of metal can be obtained. **Two important ores to know:**

Haematite for Iron [contains iron(III) oxide, Fe₂O₃]

Bauxite for Aluminium [contains aluminium oxide, Al₂O₃]


Some important minerals

Name of Mineral	Chemical Name	Formula	Metal extracted	Usual method of Extraction
Bauxite	Aluminium oxide	Al ₂ O ₃	Aluminium	Electrolysis of oxide dissolved in molten cryolite.
Galena	Lead sulphide	PbS	Lead	Sulphide is roasted in air and the oxide produced is
Haematite	Iron (III) oxide	Fe ₂ O ₃	Iron	Heat oxide with carbon
Sphalerite	Zinc Sulphide	ZnS	Zinc	Sulphide is roasted in air and the oxide produced is heated with carbon.
Copper pyrite	Copper iron sulphide	CuFeS ₂	Copper	Sulphide ore is roasted in air

TABLE 23.1 Principal Mineral Sources of Some Common Metals

Metal	Mineral	Composition
Aluminum	Bauxite	Al_2O_3
Chromium	Chromite	$FeCr_2O_4$
Copper	Chalcocite	Cu_2S
	Chalcopyrite	$CuFeS_2$
	Malachite	$Cu_2CO_3(OH)_2$
Iron	Hematite	Fe_2O_3
	Magnetite	Fe_3O_4
Lead	Galena	PbS
Manganese	Pyrolusite	MnO_2
Mercury	Cinnabar	HgS
Molybdenum	Molybdenite	MoS_2
Tin	Cassiterite	SnO_2
Titanium	Rutile	TiO_2
	Ilmenite	$FeTiO_3$
Zinc	Sphalerite	ZnS

Introduction

- The Earth's **crust** contains many **different rocks**. Rocks are a **mixture of minerals** and from some we can make useful substances.
- A **mineral** can be a **solid metallic or non-metallic element** or a **compound** found naturally in the Earth's crust.
- A **metal ore is a mineral or mixture of minerals from which economically viable amounts of metal can be extracted**, i.e. its got to have enough of the metal, or one of its compounds, in it to be worth digging out! **Ores are often oxides, carbonates or sulphides**. They are all finite resources so we should use them wisely!
- In order to extract a metal, the ore or compound of the metal must undergo a process called reduction to free the metal** (i.e. the positive **metal ion gains negative electrons to form the neutral metal atom**, or the **oxide loses oxygen**, to form the **free metallic atoms**).
- Generally speaking the **method of extraction** depends on the metals **position in the reactivity series**.

- The reactivity series of metals can be presented to include two non-metals, carbon and hydrogen, to help predict which method could be used to extract the metal.
 - lower **Pt Au Ag Cu (H) Pb Sn Fe Zn (C) Al Mg Ca Na K** higher in series
 - RULE: Any element higher in the series can displace any other lower element**
- Metals above zinc and carbon** in the reactivity series **cannot usually be extracted with carbon or carbon monoxide**. **They are usually extracted by electrolysis** of the purified molten ore or other suitable compound
 - eg aluminium** from molten aluminium oxide or **sodium** from molten sodium chloride.

- The ore or compound must be **molten** or dissolved in a **solution** in an **electrolysis cell** to allow **free movement of ions** (electrical current).
- **Metals below carbon can be extracted by heating the oxide with carbon or carbon monoxide.** The non-metallic elements **carbon will displace the less reactive metals** in a smelter or blast furnace **e.g. iron** or zinc and metals lower in the series.
 - **Metals below hydrogen will not displace hydrogen from acids.** Their oxides are easily reduced to the metal by heating in a stream of hydrogen, though this is an extraction method rarely used in industry. In fact most metal oxides below carbon can be reduced when heated in hydrogen, even if the metal reacts with acid.
- Some metals are **so unreactive** that they do not readily combine with oxygen in the air or any other element present in the Earth's crust, and so can be **found as the metal itself**. For example **gold** (and sometimes copper and silver) and no chemical separation or extraction is needed. In fact all the metals below hydrogen can be found as the 'free' or 'native' element.
- Other methods are used in special cases using the **displacement rule**. **A more reactive metal can be used to displace and extract a less reactive metal** but these are costly processes since the more reactive metal also has to be produced in the first place! **See Titanium** or see at the end of the section on **copper extraction**
- Sometimes **electrolysis is used to purify less reactive metals** which have previously been extracted using carbon or hydrogen (**eg copper** and **zinc**). Electrolysis is also used to plate one metal with another.
- The demand for raw materials does have **social, economic and environmental implications** eg conservation of mineral resources by recycling metals, minimising pollution etc.
- Historically **as technology and science have developed** the methods of extraction have improved to the point where all metals can be produced. The reactivity is a measure of the ease of compound formation and stability (ie more reactive, more readily formed stable compound, more difficult to reduce to the metal).
 - The **least reactive metals** such as gold, silver and copper have been used for the past 10000 years because the pure metal was **found naturally**.
 - **Moderately reactive metals** like iron and tin have been extracted using **carbon based smelting** for the past 2000-3000 years.
 - BUT it is only in the last 200 years that **very reactive metals** like sodium or aluminium have been extracted by **electrolysis**.

21.2 Metallurgy

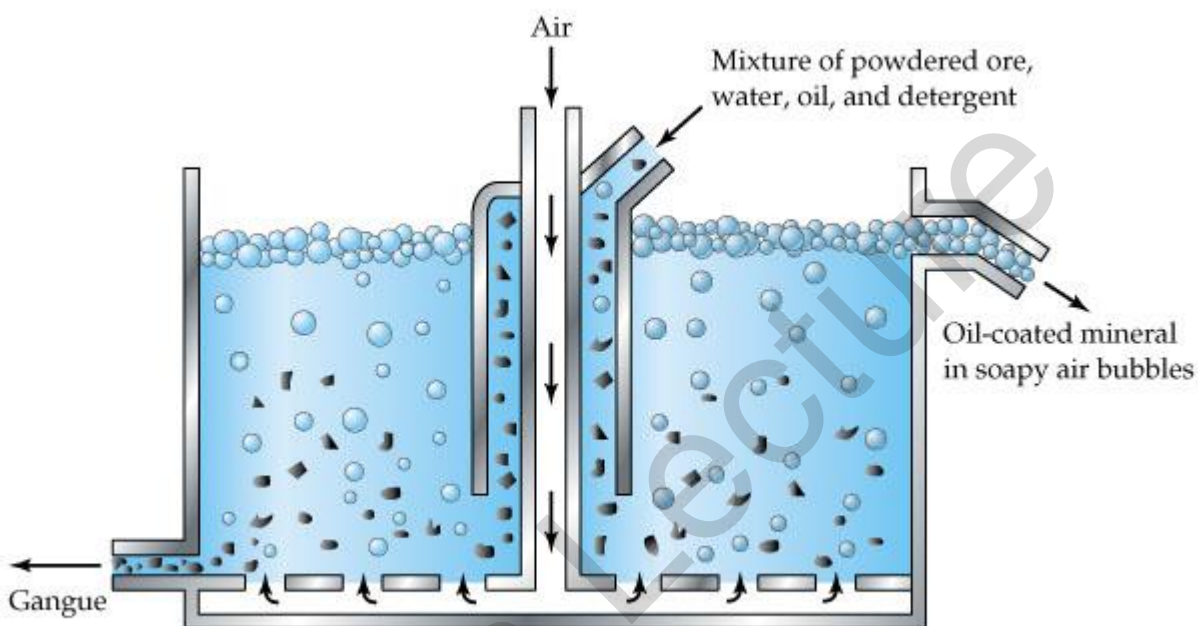
Metallurgy is the combination of science and technology used to extract metals from their ores. Ores are complex mixtures of metal-containing material and useless impurities called **gangue**. The steps involved in extracting a metal include the following:

- concentrating the ore, and chemically treating it if necessary
- reducing the mineral to free metal
- refining and purifying the metal.

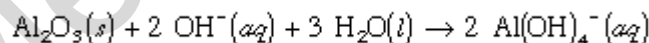
The metal may be mixed with other elements to modify its properties or to form an **alloy**, a metallic solution of two or more elements

Concentration and Chemical Treatment of Ores

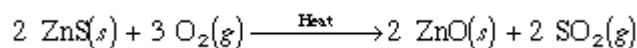
The different physical properties of the mineral and the gangue, such as density and magnetic charge, can be used to concentrate the mineral and remove the gangue. Metal sulfide ores are concentrated by **flotation**, a process that exploits differences in the ability of water and oil to wet the surfaces of the mineral and the gangue. Mineral particles float to the top of the tank along with soapy air bubbles, while the gangue sinks to the bottom.



Ores can also be concentrated by chemical means. In the Bayer process, the Al_2O_3 in bauxite is separated from Fe_2O_3 impurities by treating the ore with NaOH .

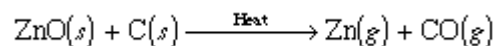


Roasting, or heating in air, is another chemical treatment used to convert minerals to compounds that are more easily reduced to the metal.

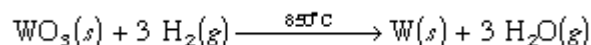


Reduction

The more active metals are obtained by reducing their ores with a chemical reducing agent. Zinc is obtained by reducing ZnO with coke, a form of carbon.

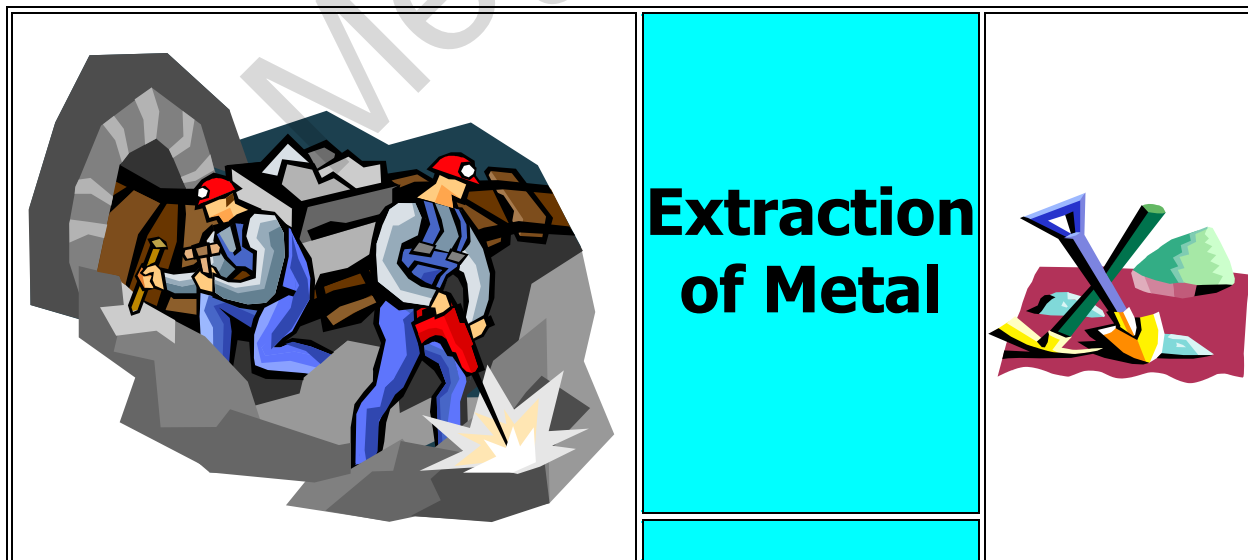
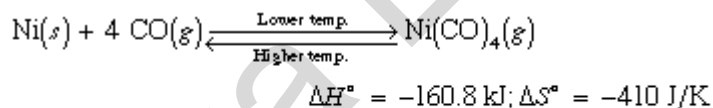


Carbon cannot be used to reduce metals that form stable carbides, such as tungsten. Tungsten(VI) oxide is reduced with hydrogen gas.

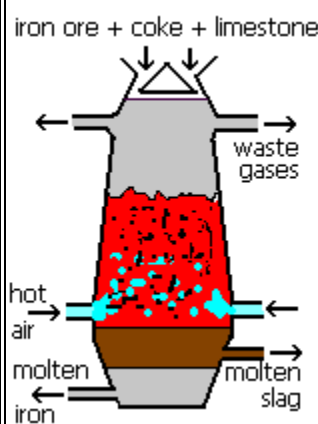
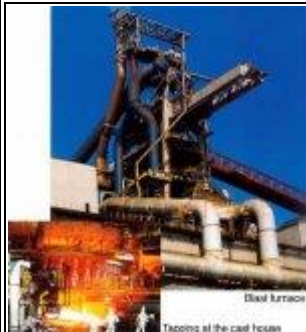


The most active metals cannot be reduced with chemical reducing agents, so these metals are produced by electrolytic reduction,.

The metals obtained from reducing ores generally require purification. Some metals, including zinc, can be purified by distillation. Nickel is purified using the **Mond process**, a chemical method in which $\text{Ni}(\text{CO})_4$ is formed and then decomposed at a higher temperature. The equilibrium shift at the higher temperature favors pure nickel.



The Extraction of Iron



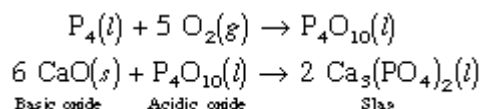
Raw Materials:

- **Iron Ore** eg **haematite** ore [iron(III) oxide, Fe_2O_3]
- **coke** (carbon, C)
- **hot air** (for the O_2 in it)
- **limestone** (calcium carbonate, CaCO_3)

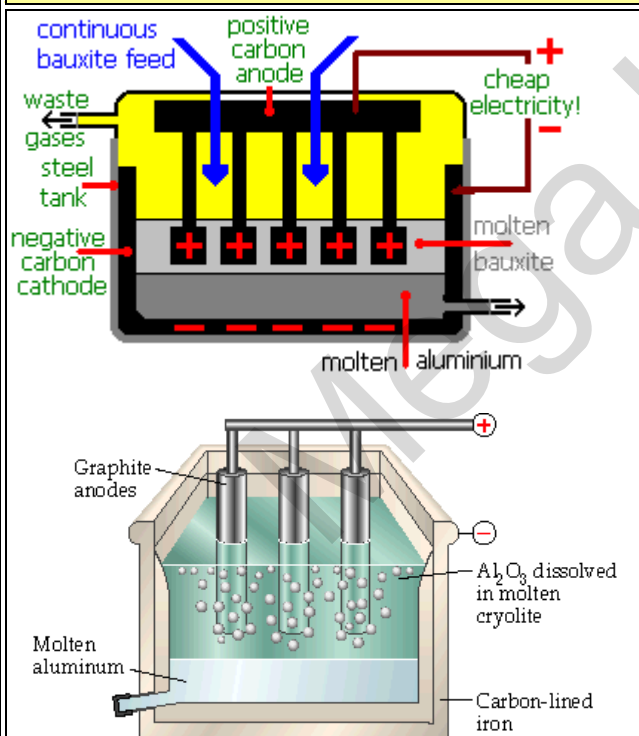
- The solid mixture of **haematite ore**, **coke** and **limestone** is continuously fed into the top of the blast furnace.
- The **coke is ignited** at the base and **hot air** blown in to burn the coke (carbon) to form carbon dioxide in an **oxidation** reaction (**C gains O**).
- The **heat energy** is needed from this very **exothermic reaction** to raise the temperature of the blast furnace to over 1000°C to effect the ore reduction. The furnace contents must be heated.
 - **carbon + oxygen ==> carbon dioxide**
 - $\text{C}_{(s)} + \text{O}_{2(g)} ==> \text{CO}_{2(g)}$
- at high temperature the carbon dioxide formed, reacts with more coke (carbon) to form carbon monoxide
 - **carbon dioxide + carbon ==> carbon monoxide**
 - $\text{CO}_{2(g)} + \text{C}_{(s)} ==> 2\text{CO}_{(g)}$
 - (note: CO_2 reduced by O loss, C is oxidised by O gain)
- The **carbon monoxide** is the molecule that actually removes the oxygen from the iron oxide ore. This a **reduction reaction** (Fe_2O_3 loses its O, or Fe^{3+} gains three electrons to form Fe) and the **CO** is known as the **reducing agent** (the O remover and gets oxidised in the process).
- This frees the iron, which is molten at the high blast furnace temperature, and trickles down to the base of the blast furnace. The main reduction reaction is ...
 - **iron(III) oxide + carbon monoxide ==> iron + carbon dioxide**
 - $\text{Fe}_2\text{O}_{3(s)} + 3\text{CO}_{(g)} ==> 2\text{Fe}_{(l)} + 3\text{CO}_{2(g)}$
 - note, as in the two reactions above, **oxidation and reduction always go together!**
 - Other possible ore reduction reactions are ...
 - $\text{Fe}_2\text{O}_{3(s)} + 3\text{C}_{(g)} ==> 2\text{Fe}_{(l)} + 3\text{CO}_{(g)}$
 - $2\text{Fe}_2\text{O}_{3(s)} + 3\text{C}_{(g)} ==> 4\text{Fe}_{(l)} + 3\text{CO}_{2(g)}$
- The original ore contains **acidic mineral impurities** such as **silica** (SiO_2 , silicon dioxide). These react with the **calcium carbonate (limestone)** to form a **molten slag** of e.g. calcium silicate.
 - calcium carbonate + silica ==> calcium silicate + carbon dioxide
 - $\text{CaCO}_3 + \text{SiO}_2 ==> \text{CaSiO}_3 + \text{CO}_2$
 - this is sometimes shown in two stages:
 - $\text{CaCO}_3 ==> \text{CaO} + \text{CO}_2$
 - $\text{CaO} + \text{SiO}_2 ==> \text{CaSiO}_3$
- The molten slag forms a layer above the more dense molten iron and they can be both separately, and regularly, drained away. The iron is cooled and cast into pig iron ingots OR transferred directly to a steel producing furnace.
- Iron from a blast furnace is **ok for very hard cast iron objects** BUT is too brittle for many applications due to too high a carbon content from the coke. So it is converted into **steel alloys** for a wide range of uses.
- The **waste slag is used for** road construction or filling in quarries which can then be landscaped or making cement.

21.3 Iron and Steel

The cast iron or pig iron produced in a blast furnace must be purified. In the **basic oxygen process**, molten iron from the blast furnace is mixed with pure oxygen gas in a furnace lined with basic oxides. The impurities in the iron are oxidized and the acidic oxides react with CaO to yield a molten slag. Phosphorus impurities react in this process to form a calcium phosphate slag.



The Extraction of Aluminium



- The purified **bauxite ore of aluminium oxide** is continuously fed in. **Cryolite** is added to lower the melting point and dissolve the ore.
- Ions must be free to move to the electrode connections** called the **cathode (-)**, attracting positive ions eg Al^{3+} , and the **anode (+)** attracting negative ions eg O^{2-} .
- When the d.c. current is passed through **aluminium forms at the positive cathode (metal*)** and sinks to the bottom of the tank.
- At the negative **anode, oxygen gas is formed (non-metal)**. This is quite a problem. At the high temperature of the electrolysis cell it burns and oxidises away the carbon electrodes to form toxic carbon monoxide or carbon dioxide. So the **electrode is regularly replaced** and the waste gases dealt with!
- It is a costly process (6x more than Fe!) due to the **large quantities of expensive electrical energy** needed for the process.

Raw materials for the electrolysis process:

- Bauxite ore of impure aluminium oxide

The redox details of the electrode processes:

- At the **negative (-) cathode, reduction**

<p>[Al₂O₃ made up of Al³⁺ and O²⁻ ions]</p> <ul style="list-style-type: none"> • Carbon (graphite) for the electrodes. • Cryolite reduces the melting point of the ore and saves energy, because the ions must be free to move to carry the current • Electrolysis means using d.c. electrical energy to bring about chemical changes eg decomposition of a compound to form metal deposits or release gases. The electrical energy splits the compound! • At the electrolyte connections called the anode electrode (+, attracts - ions) and the cathode electrode (-, attracts + ions). An electrolyte is a conducting melt or solution of freely moving ions which carry the charge of the electric current. 	<p>occurs (electron gain) when the positive aluminium ions are attracted to it. They gain three electrons to change to neutral Al atoms.</p> $\text{Al}^{3+} + 3\text{e}^{-} \Rightarrow \text{Al}$ <ul style="list-style-type: none"> • At the positive (+) anode, oxidation takes place (electron loss) when the negative oxide ions are attracted to it. They lose two electrons forming neutral oxygen molecules. $2\text{O}^{2-} \Rightarrow \text{O}_2 + 4\text{e}^{-}$ <ul style="list-style-type: none"> • Note: Reduction and Oxidation always go together! • The overall electrolytic decomposition is ... <ul style="list-style-type: none"> ○ aluminium oxide => aluminium + oxygen ○ 2Al₂O₃ ==> 4Al + 3O₂ ○ and is a very endothermic process, lots of electrical energy input!
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The original extraction of copper from copper ores

- **from copper carbonate ores* ...**
 - The ore can be roasted to concentrate the copper as its oxide.
 - Water is driven off and the carbonate thermally decomposed.
 - **copper(II) carbonate ==> copper oxide + carbon dioxide**
 - **CuCO_{3(s)} ==> CuO_(s) + CO_{2(g)}**
 - **The oxide can be smelted by heating with carbon (coke, charcoal) to reduce the oxide to impure copper**, though this method isn't really used much these days (the 'bronze age' method archaeologically!).
 - **copper(II) oxide + carbon ==> copper + carbon dioxide**
 - **2CuO_(s) + C_(s) ==> 2Cu_(s) + CO_{2(g)}**
- **from copper sulphide ores ...**
 - copper sulphide ores can be roasted in air to form impure copper
 - nasty sulphur dioxide gas is formed, this must be collected to avoid pollution and can be used to make sulphuric acid to help the economy of the process
 - **copper(I) sulphide + oxygen ==> copper + sulphur dioxide**
 - **Cu₂S_(s) + O_{2(g)} ==> 2Cu_(s) + SO_{2(g)}**
 - sulphur dioxide is a nasty toxic acidic gas, it is collected and used to make sulphuric acid, helps pay for the extraction process.
 - or ***CuS_(s) + O_{2(g)} ==> Cu_(s) + SO_{2(g)}**
 - *** the CuS might be part of an ore like chalcopyrite CuFeS₂ which is the principle ore copper is extracted from.**
- *** It is also possible to dissolve the carbonate ore or the oxide from roasted ore in dilute sulphuric acid and extracting copper by ...**
 - **(1) using electrolysis see purification by electrolysis above. or**
 - **(2) by adding a more reactive metal to displace it eg scrap iron or steel is used by**

adding it to the resulting copper(II) sulphate solution.

- iron + copper(II) sulphate ==> iron(II) sulphate + copper
- $\text{Fe}_{(s)} + \text{CuSO}_{4(aq)} \Rightarrow \text{FeSO}_{4(aq)} + \text{Cu}_{(s)}$

The Extraction of Titanium by Displacement

- Titanium ore is mainly the oxide TiO_2 , which is converted into titanium tetrachloride TiCl_4
- The chloride is then reacted with sodium or magnesium to form titanium metal and sodium chloride or magnesium chloride.
- This reaction is **carried out in an atmosphere of inert argon gas** so none of the metals involved becomes oxidised by atmospheric oxygen.
- $\text{TiCl}_4 + 2\text{Mg} \Rightarrow \text{Ti} + 2\text{MgCl}_2$ or $\text{TiCl}_4 + 4\text{Na} \Rightarrow \text{Ti} + 4\text{NaCl}$
- These are examples of **metal displacement reactions** eg the less reactive titanium is displaced by the more reactive sodium or magnesium.
- Overall the **titanium oxide ore is reduced** to titanium metal (overall O loss, oxide => metal)

Environmental Impact and Economics of Metal and other Mineral Extraction

- One of the problems of metal or mineral extraction is balancing ecological, environmental, economic, social advantages.
- It doesn't matter whether you are mining and processing iron ore or limestone, many of the advantages and disadvantages are common to these operations.
- Examples of advantages of a country exploiting its own mineral resources:
 - Valuable revenue if the mineral or its products are exported.
 - Jobs for people, especially in poor countries or areas of high unemployment in developed countries.
 - Wages earned go into the local economy.
 - Increase in local facilities promoted eg
 - transport systems, roads and recreational and health social facilities.
 - ?
- Examples of disadvantages of a country exploiting its own mineral resources and reduction of its social and environmental impact:
 - Dust from mining or processing can be reduced by air filter and precipitation systems.
 - Noise from process operation or transport of raw materials and products.
 - Difficult to deal with, sound-proofing often not practical, but operations can be reduced for unsociable hours eg evening and night!
 - Pollution can be reduced by cleaning the 'waste' or 'used' air or water of toxic or acidic materials eg
 - carbon monoxide from the blast furnace extraction of iron

- sulphur dioxide gas from copper extraction of its sulphide ore
- Mining operations will disfigure the landscape BUT it can be re-claimed and 'landscaped' in an attempt to restore the original flora and fauna.
- ?

4. How can metals be made more useful?

Extraction details Aluminium can be made more resistant to corrosion by a process called anodising. Iron can be made more useful by mixing it with other substances to make various types of steel. Many metals can be given a coating of a different metal to protect them or to improve their appearance.

- **Aluminium** is a reactive metal but it is resistant to corrosion. This is because aluminium reacts in air to form a layer of aluminium oxide which then protects the aluminium from further attack.
 - This is why it appears to be less reactive than its position in the reactivity series of metals would predict.
- For some **uses of aluminium** it is desirable to increase artificially the thickness of the protective oxide layer in a process is called **anodising**.
 - This involves removing the oxide layer by treating the aluminium sheet with sodium hydroxide solution.
 - The aluminium is then placed in dilute sulphuric acid and is made the positive electrode (anode) used in the electrolysis of the acid.
 - Oxygen forms on the surface of the aluminium and reacts with the aluminium metal to form a thicker protective oxide layer.
- **Aluminium can be alloyed to make 'Duralumin'** by **adding copper** (and smaller amounts of magnesium, silicon and iron), **to make a stronger alloy** used in aircraft components (low density = 'lighter!'), greenhouse and window frames (good anti-corrosion properties), overhead power lines (quite a good conductor and 'light'), but steel strands are included to make the 'line' stronger and poorly electrical conducting ceramic materials are used to insulate the wires from the pylons and the ground.
- The properties of **iron** can be altered by adding small quantities of other metals or carbon to make steel. **Steels are alloys** since they are mixtures of iron with other metals or with non-metals like carbon or silicon.
- **Making Steel:**
 - **(1)** Molten iron from the **blast furnace** is mixed with recycled scrap iron
 - **(2)** Then pure oxygen is passed into the mixture and the non-metal impurities such as silicon or phosphorus are then converted into acidic oxides (**oxidation process**) ..
 - eg $\text{Si} + \text{O}_2 \Rightarrow \text{SiO}_2$, or $4\text{P} + 5\text{O}_2 \Rightarrow \text{P}_4\text{O}_{10}$
 - **(3)** Calcium carbonate (a base) is then added to remove the acidic oxide impurities (in an **acid-base reaction**). The salts produced by this reaction form a slag which can be tapped off separately.
 - eg $\text{CaCO}_3 + \text{SiO}_2 \Rightarrow \text{CaSiO}_3 + \text{CO}_2$ (calcium silicate slag)
 - Reactions **(1)-(3)** produce pure iron.
 - Calculated quantities of carbon and/or other metallic elements such as titanium, manganese or chromium are then added to make a wide range of steels with particular properties.
 - Because of the high temperatures **the mixture is stirred by bubbling in unreactive argon gas!**
 - **Economics of recycling** scrap steel or iron: Most steel consists of >25% recycled

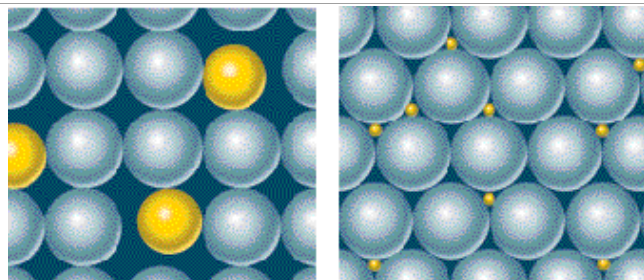
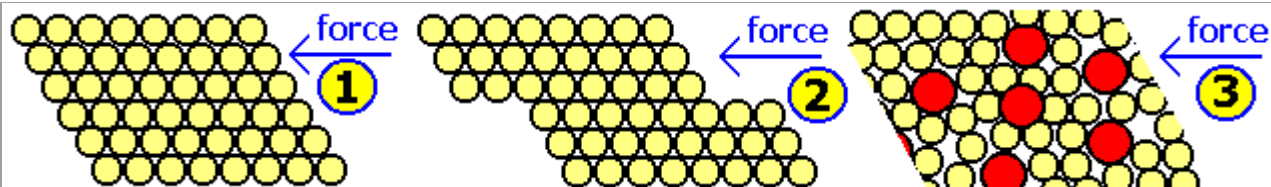
iron/steel and you do have the 'scrap' collection costs and problems with varying steel composition* BUT you save enormously because there is no mining cost or overseas transport costs AND less junk lying around! (NOTE: * some companies send their own scrap to be mixed with the next batch of 'specialised' steel they order, this saves both companies money!)

- **Different steels for different uses:**
 - High % carbon steel is strong but brittle.
 - Low carbon steel or **mild steel** is softer and is easily shaped and pressed eg into a motor car body.
 - **Stainless steel** alloys contain chromium and nickel and are tougher and more resistant to corrosion.
 - **Very strong steels** can be made by alloying the iron with titanium or manganese metal.
- Steel can be **galvanised** by coating in zinc, this is physically done by dipping the object into a bath of molten zinc. On removal and cooling a thin layer of zinc is left on. The zinc chemically bonds to the iron via the free electrons of both metals - its all the same atoms to them! It can also be done by **electroplating**.
- **Steel (and most metals) can be electroplated.**
 - The steel object to be plated is made the negative electrode (cathode) and placed in a solution containing ions of the plating metal.
 - The positive electrode (anode) is made of the pure plating metal (which dissolves and forms the fresh deposit on the negative electrode).
 - Nickel, zinc, copper, silver and gold are examples of plating metals.
 - **If M = Ni, Cu, Zn**
 - At the **positive (+) anode**, the process is an **oxidation**, electron loss, as the metal atoms dissolve to form metal(II) ions.
 - $M_{(s)} \implies M^{2+}_{(aq)} + 2e^{-}$
 - at the **negative (-) cathode**, the process is a **reduction**, electron gain by the attracted metal(II) ions to form neutral metal atoms.
 - $M^{2+}_{(aq)} + 2e^{-} \implies M_{(s)}$
 - **For silver plating it is Ag^{+} , Ag and a single electron change**
 - Any conducting (**usually metal**) object can be **electroplated with copper or silver for aesthetic reasons or steel with zinc or chromium as anti-corrosion protective layer.**
- Many other metals have countless uses eg **zinc**
 - zinc is used to make the outer casing of zinc-carbon-weak acid **batteries**.
 - It is alloyed with copper to make the **useful metal brass** (electrical plug pins). Brass alloy is stronger and more hardwearing than copper AND not as brittle as zinc.

USES OF METALS

METAL	USES	PROPERTIES THAT MAKE IT SUITABLE
Aluminium	a) Structural material for ships, planes, cars, saucepans. b) Overhead electricity cables	a) strong but light; oxide layer prevents corrosion. b) light but good conductor
Zinc	a) Coating iron to give galvanized iron b) To make alloys e.g brass (Zn/Cu) and bronze (Zn/Sn/Cu).	a) Reactive- gives sacrificial protection to iron; does not corrode easily. b) Modifies the properties of other elements.
Iron	Structural material for all industries (in the form of steel)	Strong and cheap; properties can be made suitable by alloying.
Lead	a) Car batteries. b) Solder (Pb/Sn) alloys	a) Design of battery makes recharging possible. b) low melting point.
Copper	a) Electric cables b) Pipes c) Alloys d) Coins (Cu/Ni)	a) very good conductors b) Very ductile, does not corrode easily c) d) A traditional metal for coins
Tin	Coating steel cans or tins.	Un reactive and non- toxic. Protects the steel from rusting
Nickel	Electroplating steel	Resist corrosion, shiny and attractive to look at.

Note on Alloy Structure



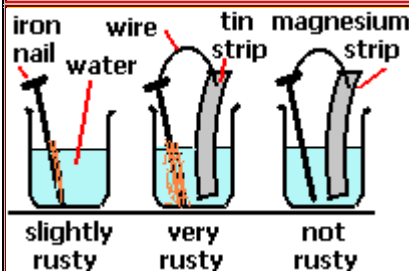
(a)

(b)

- Shows the regular arrangement of the atoms in a metal crystal and the white spaces show where the free electrons are (yellow circles actually positive metal ions).
- Shows what happens when the metal is stressed by a strong force. The layers of atoms can slide over each other and the bonding is maintained as the mobile electrons keep in contact with atoms, so the metal remains intact BUT a different shape.
- Shows an alloy mixture. It is NOT a compound but a physical mixing of a metal plus at least one other material (shown by red circle, it can be another metal eg Ni, a non-metal eg C or a compound of carbon or manganese, and it can be bigger or smaller than iron atoms). Many alloys are produced to give a stronger metal. The presence of the other atoms (smaller or bigger) disrupts the symmetry of the layers and reduces the 'slip ability' of one layer next to another. The result is a stronger harder less malleable metal.

ALLOY	COMPOSITION %	SPECIAL PROPERTIES	USES
Stain less steel	Fe = 74% Cr = 18 % Ni = 8 %	Resist corrosion	Car parts, kitchen sinks, cutlery
Cupronickle	Cu = 75% Ni = 25%	Hard wearing, attractive silver color	Silver coins
Manganese steel	Fe = 85 % Mn = 13.8 % C = 1.2	Very hard	Springs
Brass	Cu = 70% Zn = 30 %	Harder then Copper, does not corrode	Musical instruments, taps
Bronze	Cu = 90 % Sn = 10 %	Harder then brass, does not corrode.	Statues, ornaments.
Magnalium	Al = 90 % Mg = 10 %	Light but strong.	Aeroplanes bodies
Solder	Pb = 50 % Sn = 50 %	Low mwltng point but form a strong solid	Joining wires and pipes.

METAL CORROSION and the RUSTING of IRON



- **Iron (or steel)** corrodes more quickly than most other transition metals and readily does so in the presence of **both oxygen** (in air) and **water** to form an iron oxide. You can do simple experiments to show that BOTH oxygen and water are needed. Put an iron nail into (1) boiled water in a sealed tube; (2) a tube of air and a drying agent; (3) an open test tube with water. Rusting appears overnight with (3) only.
- **Rusting is speeded up** in the presence of **salt** or acid solutions because of an **increased concentration of ions**. Corrosion is a redox process involving redox electron transfer and ion movement. The rusting

metal behaves like a simple cell and more ions enable the current, and hence the electron transfer, to occur more readily.

- Rusting is **overall ...** $\text{Fe}_{(s)} + \text{O}_{2(g)} + \text{H}_2\text{O}_{(l)} \Rightarrow \text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}_{(s)}$ ie rust is a hydrated iron(III) oxide (the equation is not meant to be balanced and the amount of water x is variable, from dry to soggy!).
 - The reaction proceeds via iron(II) hydroxide $\text{Fe}(\text{OH})_2$ which is the oxidised further to the Fe_2O_3
 - **Rusting is an oxidation** because it involves iron gaining oxygen ($\text{Fe} \Rightarrow \text{Fe}_2\text{O}_3$) or iron atoms losing electrons ($\text{Fe} - 3e^- \Rightarrow \text{Fe}^{3+}$).
 - See [more examples of oxidation and reduction](#) below.
- The rusting of iron is a major problem in its use as a structural material.
- Iron and steel (alloy of iron) are most easily protected by **paint** which provides a barrier between the metal and air/water. Moving parts on machines can be protected by a **water repellent oil or grease** layer.
- This 'rusting' corrosion can be prevented by connecting iron to a more reactive metal (e.g. zinc or magnesium). This is referred to as **sacrificial protection or sacrificial corrosion**, because the more reactive protecting metal is preferentially oxidised away, leaving the protected metal intact. The picture illustrates what might be seen after a few days.* Iron or steel can also be protected by mixing in other metals (e.g. **chromium**) to make non-rusting alloys called stainless steel. The chromium, like aluminium, forms a protective oxide layer.
- * Theoretically, any iron ions formed by oxidation would be reduced by electrons from the oxidation of the more reactive 'sacrificed' metal.
- Coating iron or steel with a thin zinc layer is called '**galvanising**'. The layer is produced by electrolytic deposition by making the iron/steel the negative cathode or by dipping the iron/steel object in molten zinc ([more details](#)). The zinc preferentially corrodes or oxidises to form a zinc oxide layer that doesn't flake off like iron oxide rust does. Also, if the surface is scratched, the exposed zinc again corrodes before the iron and continues to protect it.
- Steel tin cans are protected by relatively unreacted tin and works well as long as the thin tin layer is complete. **HOWEVER**, if a less reactive metal is connected to the iron, it then the iron rusts preferentially (try scratching a 'tin' can and leave out in the rain and note the corrosion by the scratch!)

Methods of Prevention of Rusting of Iron

Covering with Paint

Covering with Grease or Oil

Covering with Chromium (Chrome Plating)

Covering with Tin (Tin plating)

Covering with Zinc Metal (Galvanising)

Using Blocks of Zinc Metal

Making Stainless Steel

Using Bocks of Magnesium Metal

- **Aluminium** does not oxidise (corrode) as quickly as its reactivity would suggest. Once a **thin oxide layer of Al_2O_3** has formed on the surface, it **forms a barrier to oxygen and water** and so prevents further corrosion of the aluminium.
- Aluminium is a useful structural metal. It can be made harder, stronger and stiffer by mixing it with small amounts of other metals (e.g magnesium) to make alloys.
- **Copper and Lead** are both used in roofing situations because neither is very reactive and the compounds formed do not flake away as easily as rust does from iron. Lead corrodes to a white lead oxide or carbonate and **copper corrodes to form a basic green carbonate** (combination of the hydroxide $\text{Cu}(\text{OH})_2$ and carbonate CuCO_3 eg seen as green roof on buildings).
- **Both metals have been used for piping** but these days lead is considered too toxic and copper is usually used as the stronger, but equally unreactive alloy with zinc, **brass**. Now of course, most piping is flowing in the plastic direction which doesn't corrode at all!
- **The Group 1 Alkali Metals** rapidly corrode in air and need to be stored under oil.

- Apart from their structural weakness they would hardly be used for any other purpose!

DONE

Mega Lecture