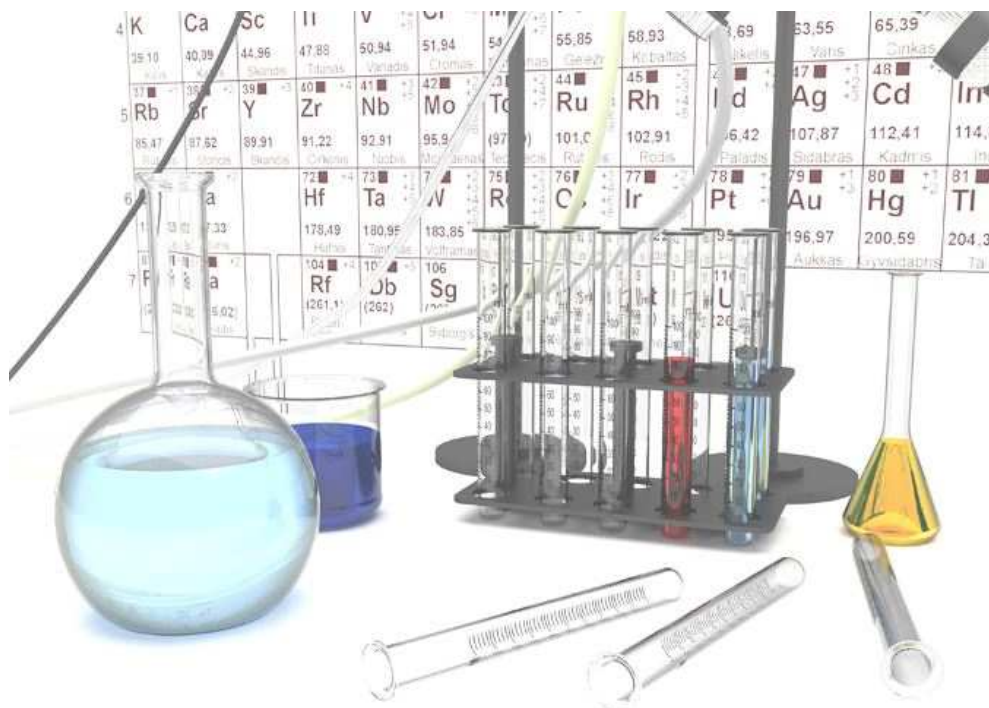




# MAROOCHYDORE SHS



## Metal Extraction Chemistry

including

Worksheet 2.1 : Overview- Metal Extraction Chemistry

Worksheet 2.2 : Metallurgy of Copper

Worksheet 2.3 : Chemistry of Iron and Steel Production

Worksheet 2.4 : Aluminium Extraction

Experiment 2..1 : Production of copper and lead

### Student booklet

Name: \_\_\_\_\_

Teacher: \_\_\_\_\_

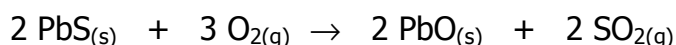
**Worksheet 2.1 OVERVIEW – METAL EXTRACTION CHEMISTRY**

A solid element or compound which occurs naturally in the Earth's crust is called a mineral. A mineral which contains a high enough percentage of a metal for economic extraction is called a metal ore. The most common metal ores are oxides (such as  $\text{Al}_2\text{O}_3$ ) and sulfides (such as  $\text{PbS}$ ). Sulfides are the oldest ores, formed in the Earth's history when there was a lot of sulfur from volcanic activity. Oxides formed later when photosynthesis in plants released large amounts of oxygen into the atmosphere. Metal ore deposits are a finite resource (there are only a certain amount of them) and non-renewable (once used, they are gone and will not be replaced).

**TABLE 21.1 Principal Types of Minerals**

Type	Minerals
Native metals	Ag, Au, Bi, Cu, Pd, Pt
Carbonates	$\text{BaCO}_3$ (witherite), $\text{CaCO}_3$ (calcite), $\text{MgCO}_3$ (magnesite), $\text{CaCO}_3 \cdot \text{MgCO}_3$ (dolomite), $\text{PbCO}_3$ (cerussite), $\text{ZnCO}_3$ (smithsonite)
Chromate	$\text{PbCrO}_4$ (crocoite)
Halides	$\text{CaF}_2$ (fluorite), $\text{NaCl}$ (halite), $\text{KCl}$ (sylvite), $\text{Na}_3\text{AlF}_6$ (cryolite)
Hydroxides	$\text{Mg}(\text{OH})_2$ (brucite), $\text{Mg}_3(\text{Si}_4\text{O}_{10})(\text{OH})_2$ (talc)
Oxides	$\text{Al}_2\text{O}_3$ (corundum), $\text{Fe}_2\text{O}_3$ (hematite), $\text{Fe}_3\text{O}_4$ (magnetite), $\text{Cu}_2\text{O}$ (cuprite), $\text{MnO}_2$ (pyrolusite), $\text{SnO}_2$ (cassiterite), $\text{TiO}_2$ (rutile), $\text{ZnO}$ (zincite)
Phosphates	$\text{Ca}_3(\text{PO}_4)_2$ , $\text{Ca}_{10}(\text{PO}_4)_6(\text{OH})_2$ (hydroxyapatite)
Silicates	$\text{Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}$ (beryl), $\text{ZrSiO}_4$ (zircon), $\text{NaAlSi}_3\text{O}_8$ (albite)
Sulfides	$\text{Ag}_2\text{S}$ (argentite), $\text{CdS}$ (greenockite), $\text{Cu}_2\text{S}$ (chalcocite), $\text{HgS}$ (cinnabar), $\text{PbS}$ (galena), $\text{ZnS}$ (sphalerite)
Sulfates	$\text{BaSO}_4$ (barite), $\text{CaSO}_4$ (anhydrite), $\text{PbSO}_4$ (anglesite), $\text{SrSO}_4$ (celestite), $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ (epsomite)

In order to extract the metal, the ore or compound of the metal must undergo a process called reduction (i.e. the positive metal ion gains electrons to form the neutral metal atom). Preliminary reactions may be necessary to convert the ore into a chemical state more suitable for reduction. Often an ore is roasted (heated in air) to drive off volatile impurities and/or convert the sulphide and carbonate forms of the ore to oxides, which more readily undergo reduction. Examples are



Note that in these reactions the metal ion does not undergo reduction.

The reduction itself can be done in several ways and generally speaking the method of extraction used depends on the position of the metal in the reactivity series

The reactivity series of metals is shown with two non-

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



metals, carbon and hydrogen, to help predict which method could be used to extract a metal.

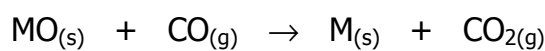
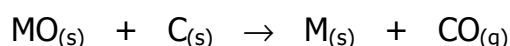
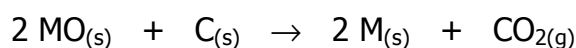
A metal low in the activity series is considered relatively unreactive – it does not lose its valence electrons easily (not undergo oxidation readily).

A metal high in the series is considered reactive – meaning it readily loses its valence electrons (undergoes oxidation readily) to more electronegative elements, such as those metals lower than it in the series.

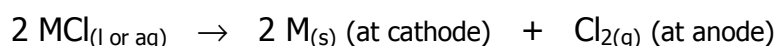
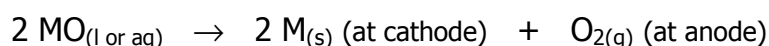
As a general rule any element higher in the series can displace any other lower element. Review the results of the reactivity of metals experiment to confirm this.

Some metals are so unreactive (bottom of the activity series of metals) 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.

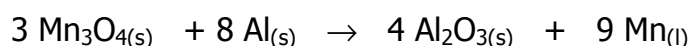
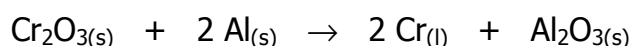
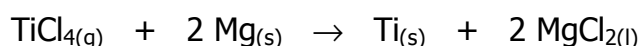
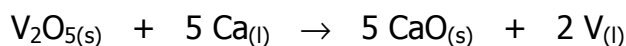
Metals below carbon can be extracted by heating the oxide with carbon. Carbon, or its combustion product carbon monoxide will displace the less reactive metals in a smelter or blast furnace. Carbon acts as a reducing agent and essentially provides electrons for the reduction of the metal ion to metal atom. Examples (M = Metal) of the reactions that occur during smelting...



Metals above 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 e.g. aluminium from molten aluminium oxide. The ore or compound must be molten or dissolved in a solution in an electrolysis cell to allow free movement of ions (electrical current). Examples (M = Metal) include...



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! Eg. Titanium. Examples include...

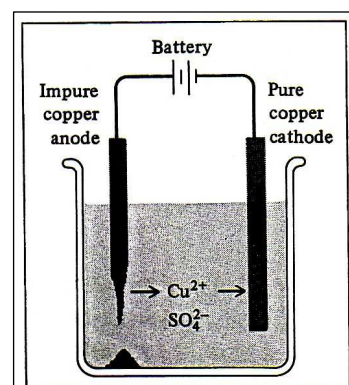


**TABLE 21.2 Reduction Processes for Some Common Metals**

Metal	Reduction Process
Lithium, sodium, potassium, magnesium, calcium Aluminum	Electrolytic reduction of the molten chloride Electrolytic reduction of anhydrous oxide (in molten cryolite)
Chromium, manganese, titanium, vanadium, iron, zinc	Reduction of the metal oxide with a more electropositive metal, or reduction with coke and carbon monoxide
Mercury, silver, platinum, copper, gold	These metals occur in the free (uncombined) state or can be obtained by roasting their sulfides

Questions:

1. Write definitions for: mineral, ore, reactivity series of metals, reactive (metal reactivity).
2. Which metal is found in the largest number of combined states (ores)? Check table 21.1, and list the different ores with their names (common and scientific).
3. Why is reduction an essential part of metal extraction chemistry?
4. What preliminary reactions may be necessary prior to the reduction of a metal ore? List examples of such reactions.
5. List the four main methods used to extract metals from their ores. Beside each write down the main factor that determines if this method is suitable for a particular metal.
6. Copper exists in nature in both the free (elemental) and combined (oxidized) states, whereas aluminium is only ever found in the combined state. Explain
7. Explain how you would obtain Zinc from sphalerite ( $ZnS$ )? Justify your choice of process by providing the reason behind your choice.
8. Consider the electrolytic procedure for purifying copper described in figure 21.1. Suppose that a copper sample to be purified in this manner had the following impurities: Fe, Ag, Zn, Au, Sn, Pt, and Pb. Which of the metals will be oxidised and dissolved into the solution; and which will not be oxidized and fall as sludge to the bottom?
9. Which of the following compounds would require electrolysis to yield the free metals:  $Ag_2S$ ,  $CaCl_2$ ,  $KCl$ ,  $Fe_2O_3$ ,  $Al_2O_3$ , and  $TiCl_4$ ? Which could be smelted in a blast furnace?
10. Outline the reactions that would occur for the production of iron from Iron oxide ( $Fe_2O_3$ ) in a blast furnace (smelting)?
11. Although iron is only about  $2/3^{rd}$  as abundant as aluminium in the earth's crust, mass for mass it costs only about  $1/4$  as much to produce. Why?



**FIGURE 21.1** The electrolytic purification of copper. The more electropositive metals, such as zinc and iron, are oxidized at the anode and enter the solution as  $Fe^{2+}$  and  $Zn^{2+}$ , while copper is transferred from the anode to the cathode via  $Cu^{2+}$  ions. The less electropositive metals, such as gold and silver, are not oxidized at the anode.



## Worksheet 2.2: METALLURGY OF COPPER

Common copper ores include:

Copper pyrite or chalcopyrite ( $\text{CuFeS}_2$ ).

Chalcocite ( $\text{Cu}_2\text{S}$ ) or copper glance.

Malachite green [ $\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$ ]. Azurite blue [ $2\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$ ].

Bornite ( $3\text{Cu}_2\text{S} \cdot \text{Fe}_2\text{S}_3$ ) or peacock ore.

Melaconite ( $\text{CuO}$ ) etc.

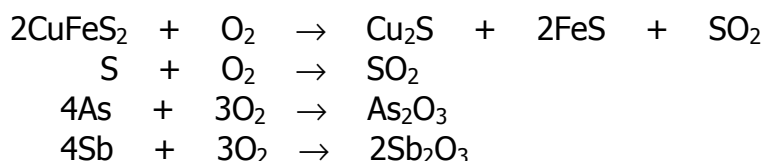
### EXTRACTION OF COPPER FROM SULPHIDE ORE

Large amount of copper are obtained from copper pyrite ( $\text{CuFeS}_2$ ) by smelting. Ores containing 4% or more copper are treated by smelting process. They are first concentrated by froth flotation. Very poor ores are treated by hydro-metallurgical process.

### EXTRACTION OF COPPER BY SMELTING PROCESS

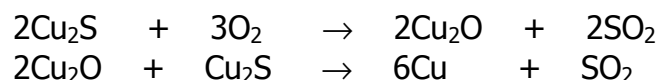
#### ROASTING

The concentrated ore is first roasted in a furnace in the presence of a current of air. Sulphur is oxidized to  $\text{SO}_2$  and impurities of arsenous and antimony are removed as volatile oxides. The following reactions take place.

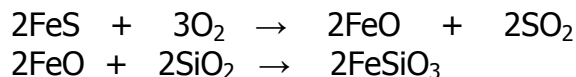


#### SMELTING

The roasted ore is mixed with coke and silica (sand)  $\text{SiO}_2$  and is introduced in to a blast furnace to produce a molten matte (mixture) of copper. The hot air is blasted through the molten matte and converts  $\text{Cu}_2\text{S}$  partly into  $\text{Cu}_2\text{O}$  which reacts with remaining  $\text{Cu}_2\text{S}$  to give molten copper.



$\text{FeS}$  and any  $\text{FeO}$  in the matte is converted to ferrous silicate ( $\text{FeSiO}_3$ ).



$\text{FeSiO}_3$  (slag) floats over the molten matte of copper.

The copper so obtained is called "Blister copper" because, as it solidifies,  $\text{SO}_2$  hidden in it escapes out producing blister on its surface.



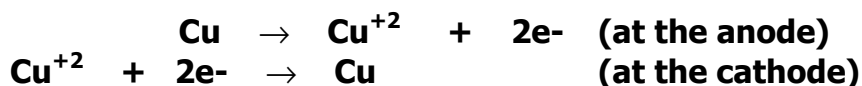
### IMPURITIES IN BLISTER COPPER AND THEIR EFFECTS

Blister copper is 99% pure. It contains impurities, mainly iron but little amount of As, Zn, Pb, Ag and Au may also be present. These impurities adversely affect the electrical as well as mechanical properties of copper. Therefore, they must be removed.

### REFINING OF COPPER

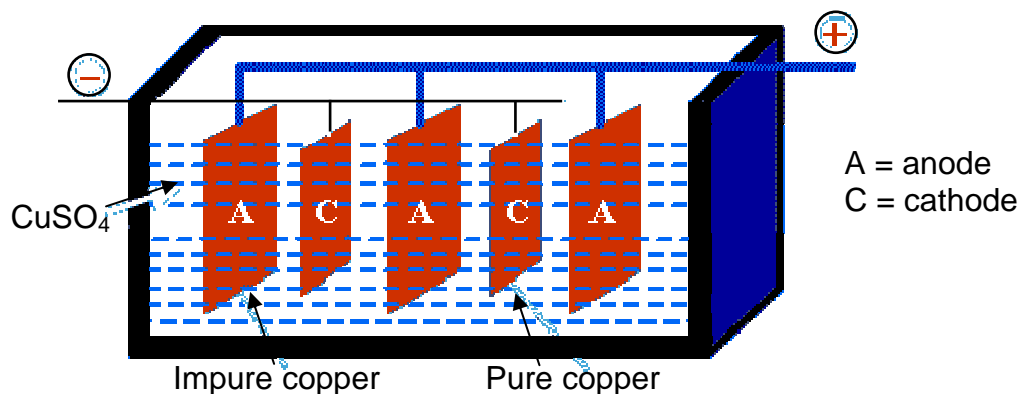
Blister copper is refined by electrolysis. Blocks of blister copper are used as anodes and thin sheets of pure copper act as cathodes. The electrolyte is copper sulphate ( $\text{CuSO}_4$ ) mixed with a little amount of  $\text{H}_2\text{SO}_4$  to increase the electrical conductivity. During electrolysis, pure copper is deposited on the cathode plates and impurities which are soluble and fall to the bottom of the cell as anode mud or sludge.

### ELECTROCHEMICAL CHANGES DURING ELECTROLYSIS



This electrically refined copper is 100% pure.

[www.citycollegiate.com](http://www.citycollegiate.com)



Information adapted from [www.citycollegiate.com/dblock4.htm](http://www.citycollegiate.com/dblock4.htm) – metallurgy of copper



## Questions

1. Copy out each of the reactions listed above for the extraction of copper from chalcopyrite. For each of these reactions assign oxidation numbers and identify which (if any) elements have undergone oxidation or reduction.
2. The smelting process is used for ore bodies over 4% copper mineral. Copper leaching (a hydro-metallurgical process) is used for ores below this (usually copper oxide ores). Discuss the advantages and disadvantages of both processes and why smelting and Leaching are used for the different ores.

Metal Extraction Chemistry



## Worksheet 2.3: Chemistry of IRON AND STEEL extraction

### Extracting iron from iron ore using a Blast Furnace

#### Introduction

The common ores of iron are both iron oxides, and these can be reduced to iron by heating them with carbon in the form of coke. Coke is produced by heating coal in the absence of air.

Coke is cheap and provides both the reducing agent for the reaction and also the heat source.

#### Iron ores

The most commonly used iron ores are **haematite** (US: hematite),  $\text{Fe}_2\text{O}_3$ , and **magnetite**,  $\text{Fe}_3\text{O}_4$ .

#### The Blast Furnace

##### The heat source

The air blown into the bottom of the furnace is heated using the hot waste gases from the top. Heat energy is valuable, and it is important not to waste any.

The coke (essentially impure carbon) burns in the blast of hot air to form carbon dioxide - a strongly exothermic reaction. This reaction is the main source of heat in the furnace.



##### The reduction of the ore

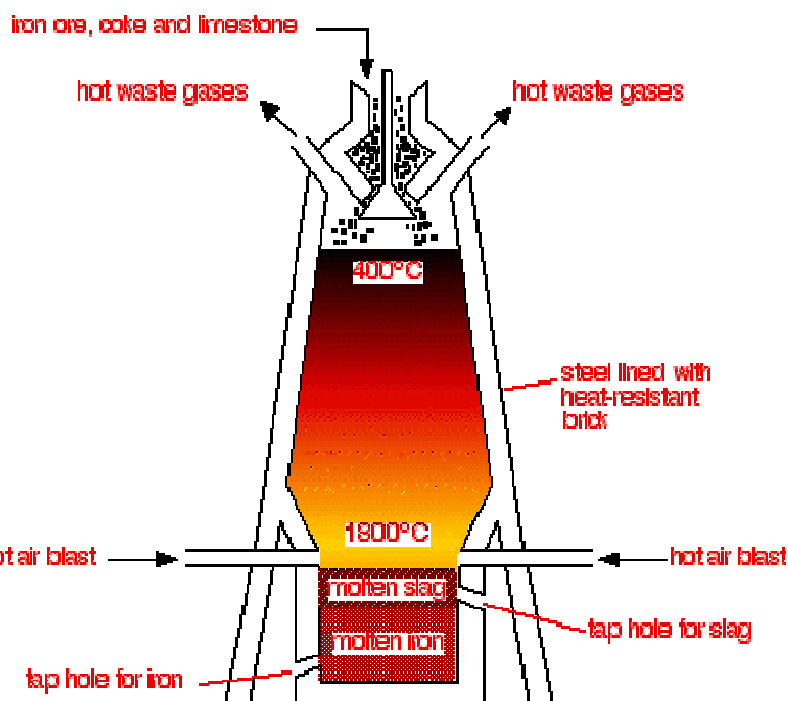
At the high temperature at the bottom of the furnace, carbon dioxide reacts with carbon to produce carbon monoxide.



It is the carbon monoxide which is the main reducing agent in the furnace.



In the hotter parts of the furnace, the carbon itself also acts as a reducing agent. Notice that at these temperatures, the other product of the reaction is carbon monoxide, not carbon dioxide.







The temperature of the furnace is hot enough to melt the iron which trickles down to the bottom where it can be tapped off.

### **The function of the limestone**

Iron ore isn't pure iron oxide - it also contains an assortment of rocky material. This wouldn't melt at the temperature of the furnace, and would eventually clog it up. The limestone is added to convert this into **slag** which melts and runs to the bottom.

The heat of the furnace decomposes the limestone to give calcium oxide.



This is an endothermic reaction, absorbing heat from the furnace. It is therefore important not to add too much limestone because it would otherwise cool the furnace.

Calcium oxide is a basic oxide and reacts with acidic oxides such as silicon dioxide present in the rock. Calcium oxide reacts with silicon dioxide to give calcium silicate.



The calcium silicate melts and runs down through the furnace to form a layer on top of the molten iron. It can be tapped off from time to time as slag.

Slag is used in road making and as "slag cement" - a final ground slag which can be used in cement, often mixed with Portland cement.

### **Cast iron**

The molten iron from the bottom of the furnace can be used as **cast iron**.

Cast iron is very runny when it is molten and doesn't shrink much when it solidifies. It is therefore ideal for making castings - hence its name. However, it is very impure, containing about 4% of carbon. This carbon makes it very hard, but also very brittle. If you hit it hard, it tends to shatter rather than bend or dent.

Cast iron is used for things like manhole covers, guttering and drainpipes, cylinder blocks in car engines, Aga-type cookers, and very expensive and very heavy cookware.

### **Steel**

Most of the molten iron from a Blast Furnace is used to make one of a number of types of steel. There isn't just one substance called steel - they are a family of alloys of iron with carbon or various metals. More about this later . . .

### **Steel-making: the basic oxygen process**

Impurities in the iron from the Blast Furnace include carbon, sulphur, phosphorus and silicon. These have to be removed.

#### **Removal of sulphur**

Sulphur has to be removed first in a separate process. **Magnesium powder** is blown through the molten iron and the sulphur reacts with it to form magnesium sulphide. This forms a slag on top of the iron and can be removed.





### **Removal of carbon etc**

The still impure molten iron is mixed with scrap iron (from recycling) and **oxygen** is blown on to the mixture. The oxygen reacts with the remaining impurities to form various oxides.

The carbon forms carbon monoxide. Since this is a gas it removes itself from the iron. This carbon monoxide can be cleaned and used as a fuel gas.

Elements like phosphorus and silicon react with the oxygen to form acidic oxides. These are removed using **quicklime (calcium oxide)** which is added to the furnace during the oxygen blow. They react to form compounds such as calcium silicate or calcium phosphate which form a slag on top of the iron.

### **Types of iron and steel**

**Cast iron** has already been mentioned above. This section deals with the types of iron and steel which are produced as a result of the steel-making process.

#### **Wrought iron**

If all the carbon is removed from the iron to give high purity iron, it is known as wrought iron. Wrought iron is quite soft and easily worked and has little structural strength. It was once used to make decorative gates and railings, but these days mild steel is normally used instead.

#### **Mild steel**

Mild steel is iron containing up to about 0.25% of carbon. The presence of the carbon makes the steel stronger and harder than pure iron. The higher the percentage of carbon, the harder the steel becomes.

Mild steel is used for lots of things - nails, wire, car bodies, ship building, girders and bridges amongst others.

#### **High carbon steel**

High carbon steel contains up to about 1.5% of carbon. The presence of the extra carbon makes it very hard, but it also makes it more brittle. High carbon steel is used for cutting tools and masonry nails (nails designed to be driven into concrete blocks or brickwork without bending). You have to be careful with high carbon steel because it tends to fracture rather than bend if you mistreat it.

**Special steels** These are iron alloyed with other metals. For example:

	iron with	special properties	uses include
<b>stainless steel</b>	Cr and Ni	resists corrosion	cutlery, cooking utensils, kitchen sinks, industrial equipment for food and drink processing
<b>titanium steel</b>	titanium	Withstands high temp	gas turbines, spacecraft
<b>manganese steel</b>	manganese	very hard	rock-breaking machinery, some railway track (e.g. points), military helmets

Modified from article sourced: <http://www.chemguide.co.uk/inorganic/extraction/iron.html>  
2005

© Jim Clark



## Questions

1. Copy out each of the reactions listed above for the extraction of Iron from Iron Oxide ore. Do this for both the Haematite ore (as described above) and the Magnetite ore ( $\text{Fe}_3\text{O}_4$ ). For each of these reactions assign oxidation numbers and identify which (if any) elements have undergone oxidation or reduction)

Metal Extraction Chemistry

2. What reaction generates most of the heat necessary for the blast furnace? Is this a redox reaction?
3. List some of the materials which still have to be removed from the molten Iron after it has been extracted? Look carefully at the processes used to remove these impurities. Are these redox reactions?
4. Iron is not one of the most common metals in the earth's crust yet it is one of the cheapest metals to buy. Suggest a reason for this.



## Worksheet 2.4: EXTRACTING ALUMINIUM FROM BAUXITE

### Introduction

Aluminium is too high in the electrochemical series (reactivity series) to extract it from its ore using carbon reduction. The temperatures needed are too high to be economic.

Instead, it is extracted by electrolysis. The ore is first converted into pure aluminium oxide by the Bayer Process, and this is then electrolysed in solution in molten cryolite - another aluminium compound. The aluminium oxide has too high a melting point to electrolyse on its own.

### Aluminium ore

The usual aluminium ore is **bauxite**. Bauxite is essentially an impure aluminium oxide. The major impurities include iron oxides, silicon dioxide and titanium dioxide.

### Purifying the aluminium oxide - the Bayer Process

#### Reaction with sodium hydroxide solution

Crushed bauxite is treated with moderately concentrated sodium hydroxide solution. The concentration, temperature and pressure used depend on the source of the bauxite and exactly what form of aluminium oxide it contains. Temperatures are typically from 140°C to 240°C; pressures can be up to about 35 atmospheres.

High pressures are necessary to keep the water in the sodium hydroxide solution liquid at temperatures above 100°C. The higher the temperature, the higher the pressure needed.

With hot concentrated sodium hydroxide solution, aluminium oxide reacts to give a solution of sodium tetrahydroaluminate.



The impurities in the bauxite remain as solids. For example, the other metal oxides present tend not to react with the sodium hydroxide solution and so remain unchanged. Some of the silicon dioxide reacts, but goes on to form a sodium aluminosilicate which precipitates out.

All of these solids are separated from the sodium tetrahydroaluminate solution by filtration. They form a "red mud" which is just stored in huge lagoons.

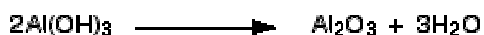
#### Precipitation of aluminium hydroxide

The sodium tetrahydroaluminate solution is cooled, and "seeded" with some previously produced aluminium hydroxide. This provides something for the new aluminium hydroxide to precipitate around.



#### Formation of pure aluminium oxide

Aluminium oxide (sometimes known as alumina) is made by heating the aluminium hydroxide to a temperature of about 1100 - 1200°C.



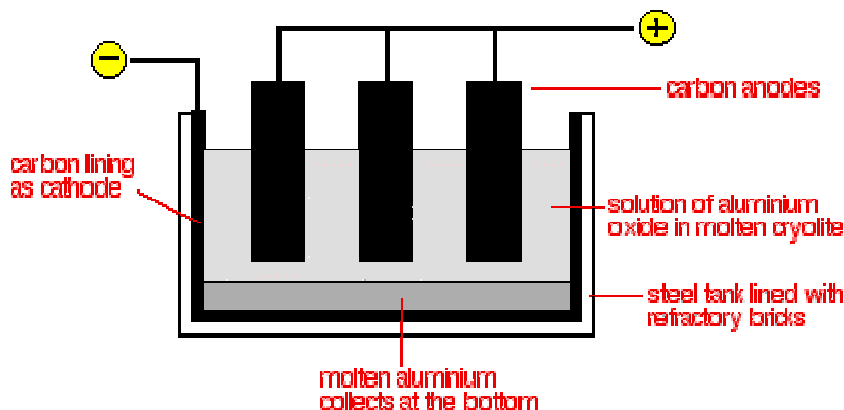


## Conversion of the aluminium oxide into aluminium by electrolysis

The aluminium oxide is electrolysed in solution in molten **cryolite**,  $\text{Na}_3\text{AlF}_6$ . Cryolite is another aluminium ore, but is rare and expensive, and most is now made chemically.

### The electrolysis cell

The diagram shows a *very simplified* version of an electrolysis cell.



Although the carbon lining of the cell is labelled as the cathode, the effective cathode is mainly the molten aluminium that forms on the bottom of the cell.

Molten aluminium is syphoned out of the cell from time to time, and new aluminium oxide added at the top.

The cell operates at a low voltage of about 5 - 6 volts, but at huge currents of 100,000 amps or more. The heating effect of these large currents keeps the cell at a temperature of about 1000°C.

### The electrode reactions

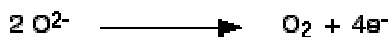
These are very complicated - in fact one source I've looked at says that they aren't fully understood. For chemistry purposes at this level, they are always simplified (to the point of being wrong! - see comment below).

This is the simplification:

Aluminium is released at the cathode. Aluminium ions are reduced by gaining 3 electrons.



Oxygen is produced initially at the anode.



However, at the temperature of the cell, the carbon anodes burn in this oxygen to give carbon dioxide and carbon monoxide.

Continual replacement of the anodes is a major expense.



## Questions

1. Copy out each of the reactions listed above for the purification of Aluminium Oxide (Alumina). For each of these reactions assign oxidation numbers and identify which (if any) elements have undergone oxidation or reduction.
2. Write a combined redox reaction for the extraction of Aluminium from alumina. Make sure the reaction is balanced in terms of electrons gained and lost.
3. Find the following sentence in the text "However, at the temperature of the cell, the carbon anodes burn in this oxygen to give carbon dioxide and carbon monoxide". Write a balanced redox reaction for this reaction. Does this reaction contribute to the extraction of aluminium or make it more difficult?
4. List the parts of the alumina purification and aluminium extraction that appear very energy intensive and expensive to perform. Explain, in terms of its standard electrode potential (-1.66 volts) AND its position in the Activity series, why the procedure for aluminium production is more expensive and complicated than most metals.

**EXPERIMENT 2.1****Production of Copper from its Ore: A Redox Reaction:**

Copper is often found in the form of a black ore, copper oxide, or a brassy-colored ore, copper sulphide. In both cases the metal has to be separated from its compound.

The demonstration in this experiment shows how copper can be extracted from its copper oxide ore. To obtain pure copper the oxygen has to be removed, using a process called reduction. The reducing agent used here is charcoal, a form of pure carbon. The reaction which occurs is similar to that used in the smelting of copper oxide ores in blast furnaces. The final product of the Redox reactions is carbon dioxide gas and copper metal.

Completion of the reaction can be visually assessed by the appearance of distinct copper metal. This can most easily be seen by examining the products using a x20 stereo microscope.

**Materials needed:**

- 1 g CuO
- 1 g charcoal
- Small amount (<0.5 gm) PbO
- Tripod
- Clay triangle
- Crucible with cover
- Bunsen burner
- Spatula
- Filter paper
- Watch glass
- Balance
- Tongs
- Matches
- Stereo microscope

Metal Extraction Chemistry

**Procedure:****Copper Oxide Smelting**

Measure 1g of CuO onto a piece of filter paper. On a second piece of filter paper, measure out 1g of charcoal. Transfer both chemicals into the crucible. Keep the filter paper that held the CuO, but discard the other piece. Use the spatula to stir and thoroughly mix the two chemicals in the crucible. Place the cover on the crucible, and set it in the clay triangle on the tripod. Heat strongly for 10-15 minutes. Then turn off the burner and allow the crucible to cool for 10 minutes (keeping the lid on). (See below for Lead smelting to complete during this time)

Use tongs to carefully remove the crucible cover. (Caution: It may still be hot!) Tip the contents of the crucible onto the watch glass. Use the Spatula to spread out the mixture, which will have the characteristic red colour of copper metal, but to see the metal properly examine under stereo microscope. Record your observations, especially the presence of any copper metal.

**Lead Oxide Smelting**

Light a match and let it burn approximately half way. Dip the charred end of the match into water and then into the PbO powder. Place the match head in the burner flame (Blue flame) and watch for the formation of tiny drops of shiny grey molten lead. Allow the match to cool, and observe the mixture under a stereo microscope. Record your observations, especially the presence of any Pb Metal.

**Questions:**

1. Describe the appearance of the copper metal produced? Explain why the copper did not form larger pieces of pure metal?
2. Describe the appearance of the Lead metal produced? Explain why Lead did form the shiny droplets of pure metal.
3. What is the oxidation number of copper in CuO, of Lead in PbO?
4. What is the oxidation number of the copper in the crucible after the completion of the reaction? And of the Lead on the matchstick?





5. Write the chemical reactions for both of these smelting/reduction processes. For each reaction nominate oxidation numbers for each chemical species, list the atoms which have undergone reductions, the atoms which have undergone oxidation, and the oxidizing and reducing agents.

6. Why did these reactions require a significant amount of heat in order to occur?