

# Module 2: Metals

## Supplementary material

### 1. Making alloys

Alloys are made by placing the required masses of the different metals in a suitable container and heating the mixture until it is completely molten: this forms a normal liquid solution. After allowing time for the mixture to become homogeneous right down to the atomic level, it is cooled. It freezes to a homogeneous solid mixture. Melting the mixture and allowing it to cool and solidify is the only way of obtaining a homogeneous mixture of two or more solids, that is to make an alloy.

Fortunately students are no longer required to make an alloy.

### 2. Relating uses to properties (Section 4.3)

*Another example of relating uses to properties is the following.*

#### Copper

Copper has a very high electrical conductivity and is very ductile and corrodes only slowly. These are the properties that make copper extremely suitable for electrical wiring in houses, offices and factories and for the cables that run around our streets. In these situations mass is not the serious factor it is in high voltage cables where the cost of the supporting towers is a dominant consideration.

Copper is easily extruded into tubes or pipes and is sufficiently soft for such pipes to be easily bent. These properties plus its high resistance to corrosion and the ease with which it can be soldered lead to copper being widely used for plumbing in houses.

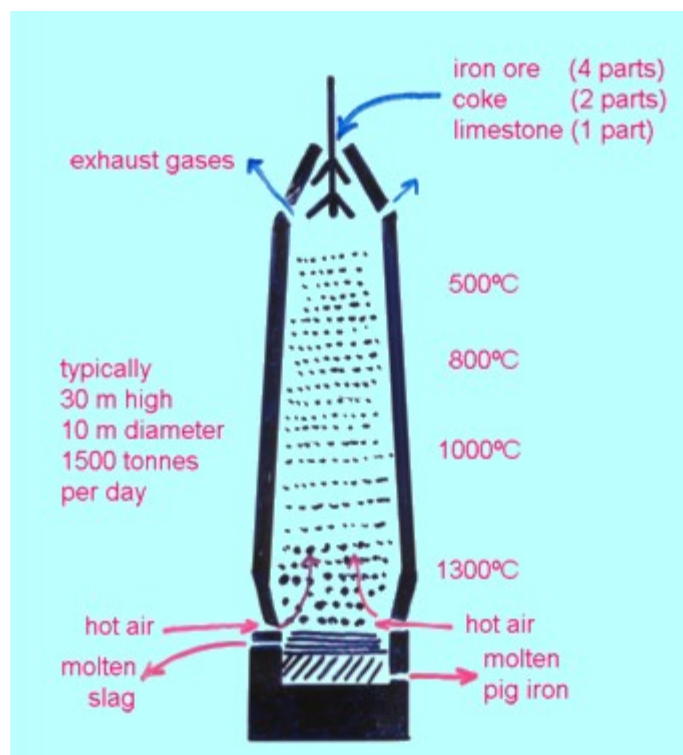
Copper forms several very important alloys, notably brass and bronze. Brass is significantly harder than copper, but still resists corrosion. This is why brass is used for making plumbing fittings such as taps, shower heads and tee junctions.

### 3. Extraction of metals from ores (Section 4.5 to 4.7)

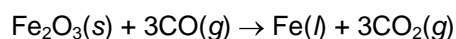
*It seems a pity that only copper is required for the HSC, particularly as its chemistry is quite complex. If you want to look at the extraction of other metals which are perhaps more representative of general methods, here are iron and aluminium.*

#### Extraction of iron

Iron ore (haematite) is mined in the Pilbara region of Western Australia and in the Middleback Range of South Australia. Open cut methods are used. The ore is virtually pure haematite so needs no purification or concentration (no Step 2 of page 110). The ore is reduced to the metal in a Blast Furnace:



A mixture of iron(III) oxide, coke (carbon) and limestone (impure calcium carbonate) is added to the top of a tower-like structure (about 30 m high) with hot air blasted in from the bottom. The carbon is converted to carbon monoxide which reduces the iron(III) oxide to iron:



The limestone decomposes to calcium oxide which combines with silica (from impurities in the ore) to form slag which floats on top of the molten iron that collects at the bottom of the furnace. The product from the blast furnace is called pig iron or cast iron. It contains about 3 to 4% carbon. This is the end of Step 3 in Figure 4.1 (page 110).

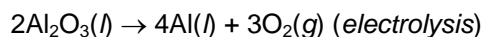
Steel is made by melting cast iron and bubbling oxygen through it to burn off most of the carbon. Alloy steels are made by adding measured amounts of other metals such as chromium, nickel, vanadium or tungsten after the carbon has been burnt off. This is Step 5 of Figure 4.1 (page 110).

### Extraction of aluminium

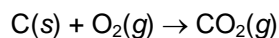
Bauxite, a mixture of aluminium oxide, iron(III) oxide and dirt (photos on pages 3 and 6), is mined by open cut methods at Weipa (Qld), Gove (N.T.) and in the south-west of W.A. The aluminium oxide is separated from the iron(III) oxide and dirt by dissolving it in hot concentrated sodium hydroxide solution. The insoluble iron(III) oxide and dirt are filtered off and discarded as 'red mud' - a major disposal problem for such mines. Cooling the sodium hydroxide solution causes hydrated aluminium oxide to crystallise out. It is filtered off and heated to produce pure alumina (aluminium oxide,  $\text{Al}_2\text{O}_3$ ). This is Step 2 of Figure 4.1. It generally occurs at the mine site.

Aluminium is too reactive a metal (Section 4.11 on page 119) for aluminium oxide to be reduced by carbon or carbon monoxide, so it has to be extracted by electrolysis. As we saw in Section 2.22 on page 57 ionic substances conduct electricity (and therefore undergo electrolysis) when they are molten. Unfortunately aluminium oxide does not melt until the temperature is above  $2000^\circ\text{C}$  which is too high for

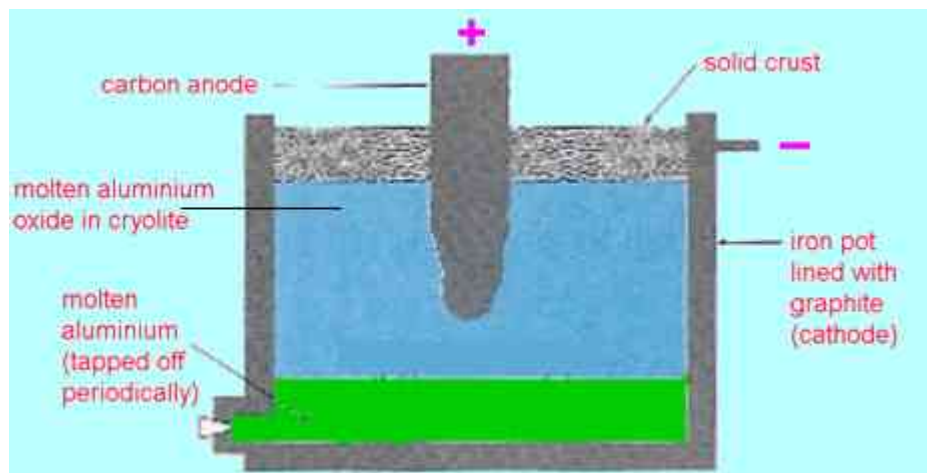
electrolysis to be practical. However a mixture of cryolite,  $K_3AlF_6$ , and aluminium oxide melts at about  $1000^\circ C$ . Electrolysis is possible at this temperature. The overall process can be written as



with the oxygen eating away the graphite (carbon) positive electrode:



This is Step 3 of the extraction process. Diagrammatically:



Virtually pure molten aluminium is removed from the electrolysis cell. While much aluminium is used pure, significant amounts are alloyed with small amounts of other metals for specialised uses.

Iron and aluminium are unusual in that their ores are so rich in the metals that no concentration of the ores (Step 2) is necessary. Copper ores have much lower metal contents and so concentration is needed.

#### 4. Oxidation and reduction and half reactions (Section 4.10)

*Not strictly in the Preliminary Course, but seems odd not to introduce the terms oxidation and reduction particularly when writing 'half equations to represent the electron transfer reactions occurring ...'*

The terms oxidation and reduction were originally introduced in connection with metals. Reduction was used to describe the extraction of a metal from its ore (usually an oxide or sulfide): the ore was *reduced* to the metal. Oxidation was originally used for the corrosion of metals because corrosion generally meant reaction with oxygen. So gain of oxygen was called oxidation. Nowadays these terms have more general use: they are applied to all elements, not just metals, and they involve processes other than just loss or gain of oxygen.

## 5. Calculating formulae from experimental data (Section 5.8)

*Calculating molecular formulae from empirical formulae appears not to be required for the HSC any more. However if you still want to do it, then continuing from the bottom of page 141:*

A chemical analysis gives us the empirical formula for a compound. This is the only formula for ionic and covalent lattice compounds. However for molecular compounds we want the molecular formula. This requires some extra information such as an estimate of the molecular weight.

### Formulae for molecular compounds

There are a variety of ways of getting an approximate estimate of the molecular weight of a compound even when the formula is not known. One such way, using volumes of gas or vapour, will be treated in Section 4.12 of the HSC Course book (pages 128–131).

Such experimental measurements show that the molecular weight of ethene is 28 whereas a chemical analysis shows that its empirical formula is  $\text{CH}_2$ . If the empirical formula of ethene is  $\text{CH}_2$ , the molecular formula must be  $(\text{CH}_2)_n$ , where  $n$  is the integer we want to determine. If the molecular formula is  $(\text{CH}_2)_n$ , then the molecular weight is:

$$n(12.0 + 2 \times 1.01) = 14n$$

$$\text{Hence } 14n = 28$$

$$n = 2$$

Therefore the molecular formula is  $\text{C}_2\text{H}_4$ .

We sometimes use the term **empirical formula weight** for the sum of the atomic weights in the empirical formula. The empirical formula weight of ethene is

$$12 + 2 \times 1 = 14. \text{ It follows then that:}$$

$$\text{molecular weight} = n \times (\text{empirical formula weight})$$

where  $n$  is the small integer described above. In summary:

### **To calculate a molecular formula from an approximate molecular weight**

1. Calculate the empirical formula (Box 5.1 on page 141)
2. Calculate the empirical formula weight (in the same way as you would calculate a molecular weight; for example for  $\text{C}_p\text{H}_q\text{O}_r\text{Cl}_s$  it is  $p \times 12.0 + q \times 1.0 + r \times 16.0 + s \times 35.5$ )
3. Calculate  $n$  from *approx molecular weight* =  $n \times (\text{empirical formula weight})$
4. Round off  $n$  to the nearest integer (errors of up to 10% are quite normal)
5. Multiply each of the subscripts in the empirical formula by  $n$  to get the molecular formula

An example will consolidate these ideas.

In an analysis of a sample of glucose it was found that it was 39.0% carbon and 6.8% hydrogen with the balance being oxygen. Calculate the empirical formula. The molecular weight was found to be approximately 175. Determine the molecular formula and calculate an accurate molecular weight for the compound.

Following Box 5.1 on page 141:

Step 1. 39.0 g C is combined with 6.8 g H and with  $(100 - 39.0 - 6.8) = 54.2$  g O.

Step 2.  $\frac{39.0}{12.0}$  mol C is combined with  $\frac{6.8}{1.01}$  mol H and with  $\frac{54.2}{16.0}$  mol O

that is, 3.25 mol C with 6.73 mol H with 3.39 mol O

Step 3.  $\frac{3.25}{3.25}$  mol C with  $\frac{6.73}{3.25}$  mol H with  $\frac{3.39}{3.25}$  mol O

1.00 mol C with 2.07 mol H with 1.04 mol O

Step 4 is not necessary, so

Step 5. Ratio by atoms is C : H : O = 1 : 2 : 1

Hence the *empirical formula* is **CH<sub>2</sub>O**. The molecular formula is therefore (CH<sub>2</sub>O)<sub>n</sub> where *n* is a small integer.

That is Step 1 of the procedure just described; now use Steps 2 to 5 of that procedure to get the molecular formula:

Step 2. empirical formula weight =  $12.0 + 2 \times 1.0 + 16.0$   
= 30.0

Step 3.  $175 = 30n$   
 $n = \frac{175}{30}$   
= 5.83

Step 4. We round *n* to 6

Step 5. The *molecular formula* is C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>.

The accurate molecular weight is  $6 \times 12.0 + 12 \times 1.01 + 6 \times 16.0 = \mathbf{180.1}$