Reaction with water

The metals above aluminium in the reactivity series react with water to produce hydrogen gas and the corresponding metal hydroxide. The general equation for this is:

metal + water → metal hydroxide + hydrogen

For example,

sodium + water → sodium hydroxide + hydrogen

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

Equation showing ions and state symbols:

$$2Na(s) + 2H_2O(l) \rightarrow 2Na^+(aq) + 2OH^-(aq) + H_2(g)$$

Questions

Q2 For the reaction between calcium and water, give:
a) a balanced equation,
b) an equation showing ions and state symbols.

Reaction with dilute acids

All metals above copper in the reactivity series react with dilute acids such as hydrochloric acid and sulphuric acid to produce a salt and hydrogen:

metal + acid → salt + hydrogen

For example,

zinc + hydrochloric acid → zinc chloride + hydrogen

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

Equation showing ions and state symbols:

$$Zn(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Zn^{2+}(aq) + 2Cl^-(aq) + H_2(g)$$

When a metal reacts with an acid it produces bubbles of hydrogen gas. In most cases, the faster the bubbles are produced, the more reactive the metal. One exception is aluminium. It reacts slowly with acids for about the first 20 minutes, after which it reacts quickly. The reason for this is that it is protected by a thin layer of aluminium oxide, which must first be removed by the acid.

Table 2.1 summarises the reactions of metals.

Questions

Q3 For the reaction between magnesium and dilute hydrochloric acid, give:
a) a balanced equation,
b) an equation showing ions and state symbols.

Q4 A, B and C are three metals. Metal A reacts with dilute hydrochloric acid but not with water. Metal B does not react with water or dilute acid. Metal C reacts with water and dilute acid.
a) Place the metals in order of reactivity, with the most reactive first.
b) Using the reactivity series in table 2.1, give all the possible metals that A, B and C could be.

Metal	Oxygen	Reaction with Water	Diluted acid
potassium sodium lithium calcium magnesium	metal + oxygen ↓ metal oxide	metal + water ↓ metal hydroxide + hydrogen	metal + acid ↓ salt + hydrogen
aluminium zinc iron tin lead copper			
mercury silver gold	no reaction	no reaction	no reaction

Table 2.1 Reactions of metals

SECTION 11.3 Extracting metals

Ore	Metal compound present
bauxite	aluminium oxide
haematite	iron oxide
galena	lead sulphide

Table 3.1 Some common ores

How do we get gold? Most of it comes from gold mines where the gold is found as a metal – it is not combined with other elements. Only unreactive metals such as gold and silver are found uncombined. Most metals are obtained from **ores**, which are compounds of metals that occur naturally. Some common ores are shown in table 3.1.

Metals are obtained from their ores by **extraction**. This involves separating the metals from the other elements with which they have combined.

Looking back through history, we can see that the first metals to be discovered were the least reactive ones: gold, silver, copper, etc. This is because they were the easiest to extract. As a rule, the less reactive a metal, the easier it is to extract from its ore.

One of the most common methods of extraction is that used to obtain a metal from its oxide. Oxides of metals below copper in the reactivity series decompose when heated to give the metal and oxygen:

$$\text{metal oxide} \rightarrow \text{metal} + \text{oxygen}$$

For example,

$$\text{silver(I) oxide} \rightarrow \text{silver} + \text{oxygen}$$
$$2Ag_2O \rightarrow 4Ag + O_2$$

Figure 3.1 Gold hair rings found in Northumberland that date from the Bronze Age

Equation showing ions and state symbols:

$$2(Ag^+)_2O^{2-}(s) \rightarrow 4Ag(s) + O_2(g)$$

Questions

Q1 For the effect of heat on mercury(II) oxide,
a) write a balanced equation,
b) write an equation showing ions and state symbols.

The method described above is useful for the extraction of gold, silver and mercury. However, for oxides of more reactive metals, heating alone produces no reaction. Instead, metals below aluminium in the reactivity series can be extracted from their oxides by heating with carbon. The carbon and oxygen join to give carbon dioxide gas:

$$\text{metal oxide} + \text{carbon} \rightarrow \text{metal} + \text{carbon dioxide}$$

For example,

$$\text{zinc oxide} + \text{carbon} \rightarrow \text{zinc} + \text{carbon dioxide}$$
$$2ZnO + C \rightarrow 2Zn + CO_2$$

Figure 3.2 Galena (lead sulphide) crystals

Equation showing ions and state symbols:

$$2Zn^{2+}O^{2-}(s) + C(s) \rightarrow 2Zn(s) + CO_2(g)$$

Table 3.2 summaries the reactions of metal oxides.

Since the industrial revolution in the eighteenth century, there has been a great demand for iron metal. This was originally used to make bridges, railway lines, factories, ships, etc. Now most iron is converted into steel

Metal	Effect of heating metal oxide: Alone	Effect of heating metal oxide: With carbon or carbon monoxide
potassium sodium lithium calcium magnesium aluminium	no reaction	no reaction
zinc iron tin lead copper		metal oxide + carbon or carbon monoxide ↓ metal + carbon dioxide
mercury silver gold	metal oxide ↓ metal + oxygen	

Table 3.2 Reactions of metal oxides

before being used. Steel is a stronger metal than iron. In 1999 over 550 million tonnes of iron were produced worldwide.

Iron is extracted from iron ore, which is usually iron(III) oxide. The process is carried out in a **blast furnace** (see figure 3.3). This is a huge structure, up to 70 metres high. Coke, which is mainly carbon, is used in the extraction. Iron ore and coke are fed in at the top of the blast furnace. Some of the iron(III) oxide and the carbon in the coke react to give iron and carbon dioxide, but the important reactions are as follows:

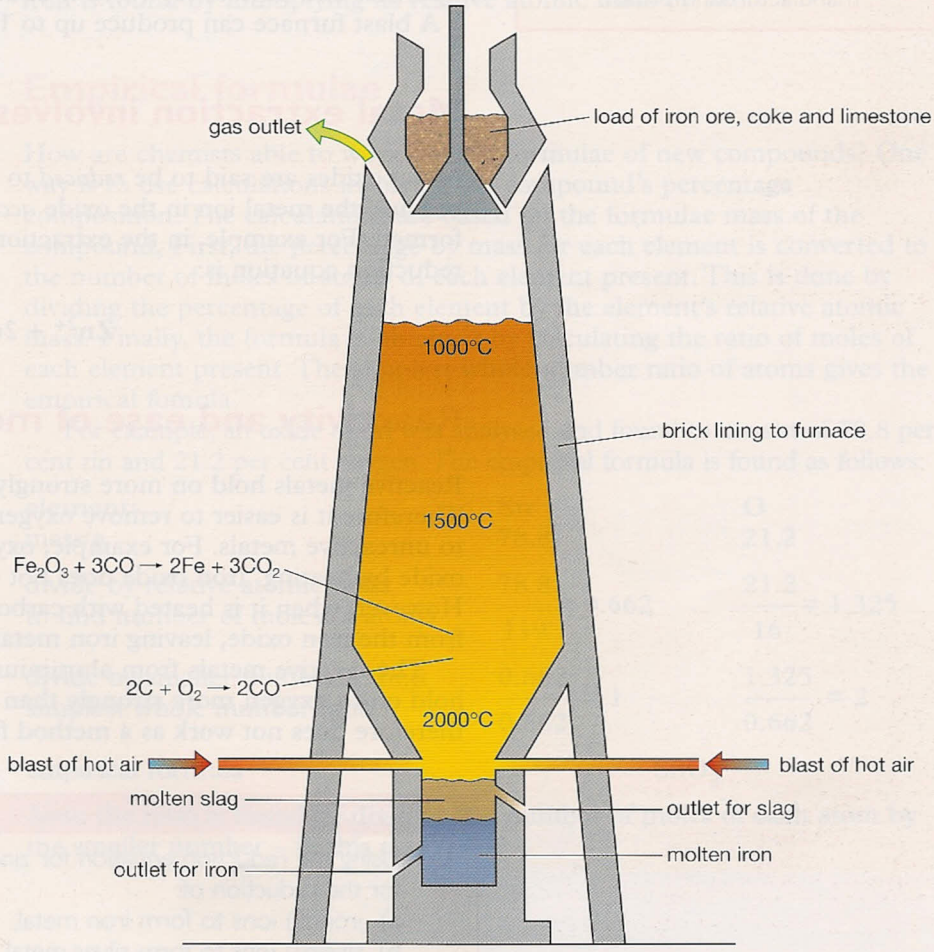

Figure 3.3 Inside a blast furnace

- At the bottom of the furnace, blasts of air are blown in. Carbon in the coke reacts with the oxygen in the air to undergo incomplete combustion. This produces carbon monoxide gas.
- Further up the furnace, carbon monoxide and iron(III) oxide react to produce iron metal and carbon dioxide – the carbon monoxide has removed the oxygen from the iron(III) oxide.

How is carbon monoxide made in a blast furnace?

The oxygen in the blast of hot air reacts with the carbon in the coke to make carbon dioxide (CO_2).

$$C + O_2 \rightarrow CO_2$$

The carbon dioxide then reacts with carbon in the coke to make carbon monoxide.

$$CO_2 + C \rightarrow 2CO$$

Taking these two equations together gives the following overall equation for the reaction.

$$2C + O_2 \rightarrow 2CO$$

The blast furnace is hot enough to melt the newly formed iron that falls to the bottom. On top floats a liquid slag which is formed by reactions between impurities and limestone, which is added with the iron ore and coke.

A blast furnace can produce up to 10 000 tonnes of iron a day.

Questions

Q2 Blast furnaces are used to make iron for the steel industry. Find out at least two places where steel is manufactured in Britain.

Metal extraction involves reduction

Metal oxides are said to be *reduced* to the metal during extraction. This is because the metal ion in the oxide *accepts electrons* when the metal is formed. For example, in the extraction of zinc from zinc oxide, the reduction equation is:

$$Zn^{2+} + 2e^- \rightarrow Zn$$

Reactivity and ease of metal extraction

Reactive metals hold on more strongly to oxygen than less reactive metals. Therefore it is easier to remove oxygen from compounds where it is joined to unreactive metals. For example, oxygen can be removed from mercury oxide by heating. Iron oxide does not change in this way with heat alone. However, when it is heated with carbon, the carbon removes the oxygen from the iron oxide, leaving iron metal behind.

The reactive metals from aluminium upwards in the reactive series hold on to oxygen more strongly than carbon does. Heating with carbon therefore does not work as a method for extracting these metals.

Questions

Q3 Using the reduction equation for zinc to help you, write similar equations for the reduction of:

a) iron(III) ions to form iron metal,

b) silver(I) ions to form silver metal.

SECTION 11.4 Calculations

Questions

Q1 Calculate the percentage by mass of iron in the ore magnetite, formula Fe_3O_4.

Q2 Calculate the percentage by mass of aluminium in the compound alumina. This can be assumed to be pure aluminium oxide, Al_2O_3.

Percentage composition

There are two important forms of iron ore. One contains a higher percentage of iron than the other. This means that, in a blast furnace, the ores will give different amounts of iron metal.

Chemists often need to know the percentage of each element in a compound. This is called the **percentage composition** of the compound.

For example, haematite consists of a compound of the elements iron and oxygen. The name of the compound is iron(III) oxide. The percentage of each can be calculated as follows:

$$\text{percentage mass of a given element in a compound} = \frac{\text{mass of element in formula}}{\text{formula mass}} \times 100$$

$$\text{formula mass of haematite } (Fe_2O_3) = (2 \times 56) + (3 \times 16) = 160$$

$$\text{percentage of iron} = \frac{\text{mass of iron in formula}}{\text{formula mass}} \times 100 = \frac{(2 \times 56)}{160} \times 100 = 70\%$$

Note: because there are two iron atoms in the formula, the mass of the iron is found by multiplying its relative atomic mass by two.

Empirical formulae

How are chemists able to work out the formulae of new compounds? One way is to use calculations involving the compound's percentage composition. The calculations are based on the formulae mass of the compound, First, the percentage by mass for each element is converted to the number of moles of atoms of each element present. This is done by dividing the percentage of each element by the element's relative atomic mass. Finally, the formula is obtained by calculating the ratio of moles of each element present. The simplest whole number ratio of atoms gives the empirical fomula.

For example, an oxide of tin was analysed and found to consist of 78.8 per cent tin and 21.2 per cent oxygen. The empirical formula is found as follows:

element:	Sn	O
mass/g	78.8	21.2
divide by relative atomic mass to find number of moles of atoms	$\frac{78.8}{119} = 0.662$	$\frac{21.2}{16} = 1.325$
divide by smaller number to find simplest whole number ratio	$\frac{0.662}{0.662} = 1$	$\frac{1.325}{0.662} = 2$
empirical formula	SnO_2	

Note: the ratio is found by dividing the number of moles of each atom by the smaller number – in this case 0.662.

Questions

Q3 An oxide of copper was analysed and found to contain 88.8 per cent copper and 11.2 per cent oxygen. Calculate the empirical formula for this oxide.

Q4 0.1225 g of magnesium were found to give 0.2025 g of magnesium oxide on complete combustion. The mass of oxygen joining with the magnesium is therefore (0.2025 – 0.1225) = 0.8000 g. Use this information to work out the empirical formula for magnesium oxide.

Q5 4.78 g of an oxide of lead on reduction gave 4.14 g of lead metal. Use this information to work out the empirical formula for the oxide.

Empirical formulae can also be found using the masses of the elements involved, not just the percentages.

For example, the formula of magnesium oxide can be worked out by heating a known mass of magnesium in a crucible until it is completely converted to magnesium oxide, which is then weighed (see figure 4.1). Subtraction gives the mass of oxygen which has combined with the magnesium.

In an experiment 2.80 g of iron combined with 1.20 g of oxygen to make an oxide of iron.

The empirical formula is found as follows:

	Fe	O
	2.80 g	1.20 g
divide by relative atomic mass to find number of moles of atoms	$\frac{2.80}{56} = 0.05$	$\frac{1.20}{16} = 0.075$
divide by the smaller number to find the simplest whole number ratio	$\frac{0.05}{0.05} = 1$	$\frac{0.075}{0.05} = 1.5$
multiplying by 2 to change 1.5 into a whole number	$1 \times 2 = 2$	$1.5 \times 2 = 3$

This gives the empirical formula of Fe_2O_3

Finding the molecular formula

The empirical formula is not always the same as the molecular formula for a covalent compound. However, given the empirical formula and the formula mass it is possible to work out the molecular formula.

For example, the hydrocarbon benzene has the empirical formula CH and a formula mass of 78. The molecular formula of benzene must be C_nH_n, where n is a whole number.

formula mass of $C_nH_n = (12 \times n) + (1 \times n) = 78$

$$13n = 78$$

$$n = 6$$

Thus the molecular formula of benzene is C_6H_6.

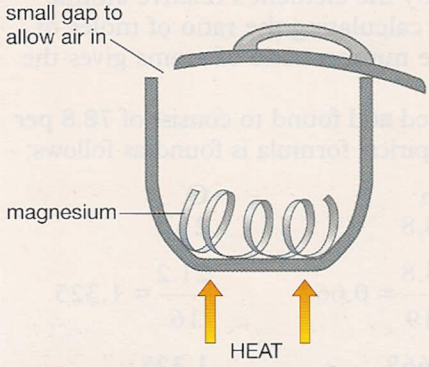

Figure 4.1 Making magnesium by combustion

Calculations based on balanced equations

Balanced equations can be used to make predictions about the masses of substances either reacting or being produced. In order to do this, every formula is taken to represent one mole of the substance concerned.

For example, what mass of hydrogen would be produced when 6 g of magnesium react completely with dilute hydrochloric acid?

balanced equation:	$Mg + 2HCl \rightarrow MgCl_2 + H_2$		
relate moles of required substances:	1 mole	⟷	1 mole
replace moles by formula mass in grams:	24 g	⟷	2 g
use simple proportion:	6 g	⟷	$\frac{2 \times 6}{24}$
			= 0.5 g

Questions

Q6 The gas cyanogen has the empirical formula CN and a formula mass of 52. What is its molecular formula?

Questions

Q7 Calculate the mass of tin that would be produced if 7.55 g of tin(IV) oxide were reduced by hydrogen. The equation for the reaction taking place is:

$$SnO_2 + 2H_2 \rightarrow Sn + 2H_2O$$