## National 5 Chemistry

### Unit 3:

**Chemistry In Society**

**Student:**

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## Topic 7

**Metal Chemistry**

### Topics

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7.1 Oxidation & Reduction

A Brief History Lesson

John Dalton was the first of the ‘modern’ chemists to put forward the idea of atoms. Much of his work was built on a growing understanding of what happened during chemical reactions - in particular the realisation that substances that were heated in air often got heavy because the atoms in the substance were joining together with atoms of oxygen in the air to form new substances.

In particular, metals were heated and the new compounds formed were called oxides and the reaction was called Oxidation.

Of equal interest, and perhaps of greater importance, was the reverse reaction - turning metal oxides back into metals by removing oxygen. This made the substances become lighter in weight so the name Reduction was used.

So, Oxidation = gai oxygen atoms
Reduction = los oxygen atoms

These definitions remain very useful today. Biochemists, in particular, often find it easier to use these definitions, rather than the more chemical definitions you are going to learn soon. For example, from the last Topic:

Ethanoic acid is normally manufactured from eth in a two-step reaction called oxidation.

\[ \text{乙醇} \rightarrow \text{乙酸} \]

- hydroxyl group is the functional group of an alcohol
- carboxyl group is the functional group of an acid

\[ \text{C}_2\text{H}_5\text{OH} \rightarrow \text{CH}_3\text{COOH} \]
**REDOX Reactions**

**copper oxide** + **carbon**

**lime water**

Oxi and Red usually take place together in the same reaction. For example, previously you have heated copper (II) oxide in a test tube with carbon.

After a few minutes the black copper (II) oxide turned red-brown and the lime water turned cloudy. Copper and carbon dioxide were produced.

The copper(II) oxide has lost oxygen: it has been reduced.

\[ \text{CuO} + \text{C} \rightarrow \text{Cu} + \text{CO}_2 \]

The carbon has gained oxygen: it has been oxidised.

Such a reaction is called a REDOX reaction.

**Modern Definitions**

A more detailed look at what actually happens during Oxi and Red has led to a redefining of these terms, allowing them to be used to describe ANY similar reaction, even when oxygen is not involved.

When magnesium is oxidised to form magnesium oxide the following reaction is taking place.

\[ 2\text{Mg} \, (s) + \text{O}_2 \, (g) \rightarrow 2\text{Mg}^2+\text{O}_2^- \, (s) \]

The magnesium atoms are being changed into magnesium ions by losing electrons. i.e.

\[ \text{Mg} \, (s) \rightarrow \text{Mg}^2+ + \text{e}^- \]

Oxidation means a loss of ele...
When mercury oxide is heated the following reaction takes place.

\[ 2\text{Hg}^{2+}\text{O}_2(\text{s}) \rightarrow 2\text{Hg}(\text{l}) + \text{O}_2(\text{g}) \]

The mercury ions are being changed into mercury atoms by gaining electrons.

REDUCTION means a gain of electrons.

OIL RIG

Oxidation is Loss of electrons
Reduction is Gain of electrons

For the rest of this topic we will examine various aspects of chemistry that involve REDOX reactions.

Much of Topic 3 in S3 was used to establish the characteristic reactions of metals.

- Metal + oxygen → metal oxide
- Metal + water → metal hydroxide (alkali) + hydrogen
- Metal + acid → metal salt + hydrogen

We will quickly revise them, but in the context of REDOX reactions. You will also learn to 'write' ion-electron ½-equations, though, normally these can be extracted from the 'Electrochemical Series' in the Data Booklet.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>lithium</td>
<td>( \text{Li}^+(\text{aq}) + e^- \rightarrow \text{Li}(s) )</td>
</tr>
<tr>
<td>potassium</td>
<td>( \text{K}^+(\text{aq}) + e^- \rightarrow \text{K}(s) )</td>
</tr>
<tr>
<td>calcium</td>
<td>( \text{Ca}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Ca}(s) )</td>
</tr>
<tr>
<td>sodium</td>
<td>( \text{Na}^+(\text{aq}) + e^- \rightarrow \text{Na}(s) )</td>
</tr>
<tr>
<td>magnesium</td>
<td>( \text{Mg}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Mg}(s) )</td>
</tr>
<tr>
<td>aluminium</td>
<td>( \text{Al}^{3+}(\text{aq}) + 3e^- \rightarrow \text{Al}(s) )</td>
</tr>
<tr>
<td>zinc</td>
<td>( \text{Zn}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Zn}(s) )</td>
</tr>
</tbody>
</table>
Q1.

Using your Data Booklet, firstly complete each of the equations below by adding the required number of electrons. Then for each of the equations, decide whether it is a reduction or an oxidation.

a) $\text{Ni}^{2+}_{(aq)} \rightarrow \text{Ni}_{(s)}$

b) $I_2_{(aq)} \rightarrow 2 I^-_{(aq)}$

c) $H_2_{(g)} \rightarrow 2 H^+_{(aq)}$

d) $Fe^{3+}_{(aq)} \rightarrow Fe^{2+}_{(aq)}$

e) $Al_{(s)} \rightarrow Al^{3+}_{(aq)}$

f) $2S_2O_3^{2-}_{(aq)} \rightarrow S_4O_6^{2-}_{(aq)}$

g) $2Cl^-_{(aq)} \rightarrow Cl_2_{(g)}$

h) $Cu^{2+}_{(aq)} \rightarrow Cu_{(s)}$

Q2.

Using your Data Booklet to help you, write ion-electron half-equations for each of the following.

a) reduction of $Al^{3+}_{(aq)}$  
   $Al^{3+}_{(aq)} \rightarrow$

b) oxidation of $Fe^{2+}_{(aq)}$ 
   $Fe^{2+}_{(aq)} \rightarrow$

c) reduction of $Br_2_{(l)}$ 
   $Br_2_{(l)} \rightarrow$

d) oxidation of $SO_3^{2-}_{(aq)}$ 
   $SO_3^{2-}_{(aq)} \rightarrow$
7.2 Metal Reactions

**Metals & OXYGEN**

Various metals were placed in test-tubes as shown, with *pot perm* at the bottom to release *oxy* gas when heated.

Most of the metals glowed *bri* and were able to *me* the glass of the test-tube.

*Mag* was the most reactive of the metals used

*General word equation:*

\[
\text{metal } + \text{ oxygen } \rightarrow \text{metal oxide}
\]

**Oxidation:** the metal has been *oxi* because it will be *lo* electrons as the metal *at* change into metal *io*.

eg \[\text{Cu } \rightarrow \text{Cu}^{2+} + 2e^-\]

**Reduction:** the oxygen atoms are *red* because they will be *gai* electrons as the oxygen *at* change into oxygen *io*.

eg \[\text{O } + 2e^- \rightarrow \text{O}_2^-\]

**Redox:** overall, *ele* are transferred from the *me* to the *ox* and an *io* compound is formed.

eg \[\text{Cu } + \text{O}_2 \rightarrow \text{CuO}\]

**Metals & WATER**

Various metals were reacted with water, either in a trough or in a test-tube as shown.

Most metals hardly react at all with water, but the members of the *Alk met* family are very reactive

*Pot* was the most reactive of the metals used
General word equation:

\[
\text{metal} + \text{water} \rightarrow \text{metal hydroxide} + \text{hydrogen}
\]

**Oxidation**: the metal has been *oxi* because it will be *los* electrons as the metal *atro* change into metal *io*.

eg \[ \text{Zn} \rightarrow + \]

**Reduction**: the water molecules are *red* because they will be *gai* electrons as the water *mol* change into hydroxide *io* and hydrogen gas.

eg \[ + \rightarrow + \]

**Redox**: overall, *ele* are *transferred* from the *met* to the *wat* and an *ion* compound is formed.

eg \[ + \rightarrow + \]

---

**Metals & ACID**

Various metals were reacted with acid as shown.

Most metals react with acid, but *cop* is one that does not react at all.

*Mag* was the most reactive of the metals used.

General word equation:

\[
\text{metal} + \text{acid} \rightarrow \text{salt} + \text{hydrogen}
\]

**Oxidation**: the metal has been *oxi* because it will be *los* electrons as the metal *atro* change into metal *io*.

eg \[ \text{Zn} \rightarrow + \]

**Reduction**: the hyd *ions* in the acid are *red* because they will be *gain* electrons and changing into *hydrogen mol*, ie hydrogen gas

eg \[ \text{H}^+ + \rightarrow \text{H}_2 \]

**Redox**: overall, *electrons* are *transferred* from the *metal* to the *hydrogen ionic*.

eg \[ \text{Zn} + \text{H}^+ \rightarrow + \text{H}_2 \]
These reactions were used to place metals into an order of reactivity. This is also known as the Rea Ser.

<table>
<thead>
<tr>
<th>Metal</th>
<th>With Water</th>
<th>With Acid</th>
<th>With Oxygen</th>
</tr>
</thead>
<tbody>
<tr>
<td>potassium</td>
<td>React to produce hydrogen gas</td>
<td></td>
<td></td>
</tr>
<tr>
<td>sodium</td>
<td>too dangerous to add to acid</td>
<td></td>
<td></td>
</tr>
<tr>
<td>lithium</td>
<td>React to produce hydrogen gas</td>
<td></td>
<td></td>
</tr>
<tr>
<td>calcium</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>magnesium</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>aluminium</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>zinc</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>iron</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>tin</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>lead</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>hydrogen</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>copper</td>
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</tr>
<tr>
<td>mercury</td>
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<td></td>
</tr>
<tr>
<td>silver</td>
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<td></td>
<td></td>
</tr>
<tr>
<td>gold</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>platinum</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

NO REACTION

The order of reactivity gives us a measure of how ‘good’ different met are at los electrons during chemical reactions.

Notice that sometimes we start by writing the ion-electron ½-equations and then combine them to produce an overall Redox equation.

For example, a pupil added sodium sulphite to iron (III) nitrate and observed a colour change showing that there had been a reaction. Sodium ions (Na⁺) and nitrate ions (NO₃⁻) are very stable/unreactive so we go looking at the sulphite (SO₃²⁻) and iron (III) (Fe³⁺) ions.

Copying from data book:

\[ \text{SO}_3^{2-} + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} + 2\text{H}^+ + 2\text{e}^- \]

\[ \text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+} \]

Before combining the equations, we make sure that electrons lost = electrons gained

Balancing:

\[ 2 \text{Fe}^{3+} + 2\text{e}^- \rightarrow 2 \text{Fe}^{2+} \]

Now we combine the two equations (electrons should 'cancel' out).

Overall:

\[ 2 \text{Fe}^{3+} + \text{SO}_3^{2-} + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} + 2\text{H}^+ + 2\text{Fe}^{2+} \]

Notice that the overall equation doesn't include the Spectator Ions, sodium ions (Na⁺) and nitrate ions (NO₅⁻)
Displacement Reactions

This is in fact a new type of reaction to you. Displacement is the name given to the reaction in which one metal takes the place of another in a solution. The metal in solution will be in the form of an ionic compound.

When zinc, a silvery grey coloured metal, is added to a blue coloured solution containing copper ions, the following changes occur.

The zinc disappears as the zinc dissolves into the solution.

At the same time the blue colour in the solution disappears and a red brown solid is formed which can be collected in the filter paper.

Oxidation:- the zinc has been oxidised because it will be losing electrons as the metal atoms change into metal ions and dissolve.

eg \[ \text{Zn} \rightarrow \text{Zn}^{2+} + \]

Reduction:- the copper ions in the solution are reduced because they will be gaining electrons and changing into copper atoms which precipitate out of the solution as insoluble solid.

eg \[ \text{Cu}^{2+} + \rightarrow \]

Redox:- overall, electrons are transferred from the zinc to the copper.

eg \[ \text{Zn} + \text{Cu}^{2+} \rightarrow + \]

(NB. this equation represents the reaction without the other ion present in the solution. This ion plays no part in the reaction, it is a spectator ion and will end up with the ion as part of the new solution.)

RULE metals higher in the order of reactivity will be able to displace metals lower in the order

If we consider the H\(^+\) ion as being a ‘metal’ ion then the reaction between metals and acids is just another displacement reaction. This explains why chemists like to include hydrogen in the order of reactivity.
Q1.

For each reaction, complete the word equation and write an ionic formula equation.

Using your Data Booklet, write the ion-electron $\frac{1}{2}$-equations for the reduction and oxidation reactions.

\( \text{a)} \) word: potassium + water $\rightarrow$ + hydrogen

ionic formulae: (s) + (l) $\rightarrow$ (aq) + (g)

oxidation: (s) $\rightarrow$ (aq) +

reduction: (l) + $\rightarrow$ (aq) + (g)

\( \text{b)} \) word: calcium + water $\rightarrow$ + hydrogen

ionic formulae: (s) + (l) $\rightarrow$ (aq) + (g)

oxidation: (s) $\rightarrow$ (aq) +

reduction: (l) + $\rightarrow$ (aq) + (g)

\( \text{c)} \) word: magnesium + oxygen $\rightarrow$

ionic formulae: (s) + (g) $\rightarrow$ (s)

oxidation: (s) $\rightarrow$ (s) +

reduction: (g) + $\rightarrow$ (s)

\( \text{d)} \) word: + oxygen $\rightarrow$ copper (I) oxide

ionic formulae: (s) + (g) $\rightarrow$ (s)

oxidation: (s) $\rightarrow$ (s) +

reduction: (g) + $\rightarrow$ (s)
Q2.

For each reaction, complete the word equation and write an ionic formula equation.

Using your Data Booklet, write the *ion-electron ½-equations* for the *reduction* and *oxidation* reactions.

| a)   | word: zinc + hydrochloric acid → | ionic formulae: Zn(s) + 2 H^+(aq) → Zn^2+(aq) + H_2(g) |
|      | oxidation: Zn(s) → (aq) + (g) | reduction: (aq) + (g) |

| b)   | word: iron + sulfuric acid → iron (III) + hydrogen sulfate |
|      | ionic formulae: 2 Fe(s) + 3 (H^+)_2SO_4^2−(aq) → (Fe^3+)_2(SO_4^2−)₃(aq) + H_2(g) |
|      | oxidation: 2 Fe(s) → 2 Fe^3+(aq) + 6e− | reduction: 6 H^+(aq) + 6e− → 3 H_2(g) |

| c)   | word: magnesium + copper (II) sulfate → + copper |
|      | ionic formulae: Mg(s) + Cu^{2+}(aq) SO_4^{2−}(aq) → Mg^{2+}(aq) SO_4^{2−}(aq) + Cu(s) |
|      | oxidation: Mg(s) → Mg^{2+}(aq) + 2e− | reduction: Cu^{2+}(aq) + 2e− → Cu(s) |

| d)   | word: copper + silver (I) nitrate → copper (II) nitrate |
|      | ionic formulae: Cu(s) + 2 Ag^+(aq) NO_3^−(aq) → Cu^{2+}(aq)(NO_3^−)₂(aq) + 2 Ag(s) |
|      | oxidation: Cu(s) → Cu^{2+}(aq) + 2e− | reduction: 2 Ag^+(aq) + 2e− → 2 Ag(s) |
7.3 Electrochemistry

A chemical reaction producing electricity is called a Cell. When cells are connected together a Battery is formed.

Redox reactions always involve one substance losing electrons and transferring them to another substance which gains the electrons.

If the two reactants can be kept separate, so the electrons have to travel through wires, then a portable power supply is possible - the battery.

The simplest, and earliest, batteries involved discs of copper and zinc separated by cardboard soaked in salt solution. Whenever two different metals are connected to each other, electrons will flow from the more reactive to the less reactive. In this case, from the zinc to the copper.

Oxidation: the zinc has been oxidised because it will be losing electrons as the metal atoms change into metal ions.

\[ \text{eg } \text{Zn} \rightarrow \text{Zn}^{2+} + 2e^- \]

Reduction: electrons flow to the copper, but copper atoms are not willing to accept extra electrons.

However, even unreactive metals like copper will react with water in the filter paper and will have slowly produced some copper ions. These copper ions in the solution are reduced as they will be forced to gain electrons and change back into copper atoms.

\[ \text{eg } \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \]

Redox: overall, electrons are transferred from the zinc to the copper.

\[ \text{eg } \text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu} \]

The voltage is a measure of how strongly the zinc can push electrons onto copper.
If the \( \text{Zn} \) is replaced with a series of different \( \text{met} \), it is possible to compare how strongly different metals can push \( \text{elec} \) onto \( \text{cop} \).

The \( \text{vol} \) recorded is a measure of the \( \text{stre} \) of the push.

Some typical results are given below.

<table>
<thead>
<tr>
<th>metal connected to copper</th>
<th>Voltage (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>magnesium</td>
<td>1.5</td>
</tr>
<tr>
<td>aluminium</td>
<td>0.9</td>
</tr>
<tr>
<td>zinc</td>
<td>0.9</td>
</tr>
<tr>
<td>iron</td>
<td>0.5</td>
</tr>
<tr>
<td>lead</td>
<td>0.5</td>
</tr>
<tr>
<td>nickel</td>
<td>0.3</td>
</tr>
<tr>
<td>copper</td>
<td>0.0</td>
</tr>
<tr>
<td>silver</td>
<td>-0.2</td>
</tr>
</tbody>
</table>

The order of \( \text{stre} \) for these metals should be familiar to you. They are in the same order as the \( \text{react} \) series.

This is not surprising since \( \text{los} \) electrons in a chemical reaction is likely to be similar to \( \text{pus} \) electrons away round a circuit.

Similar but \textit{not exactly the same}. In your Data Book you will find a more complete list under the heading “\textit{Electrochemical Series}”. The four metals above \( \text{mag} \) are in a different order from the \textit{Reactivity Series}, but after that the order is the same except for \( \text{mer} \) and \( \text{sil} \) which have exchanged positions.

The rule should really be that ‘\textit{electrons will flow from the metal higher in the electrochemical series to the metal lower in the series}’.

Though not exactly the same, the \textit{electro} series can help you to remember the \textit{react} series.

\textbf{Electrochemical Cells}

Any two \textit{diff} \( \text{met} \) that are \textit{conn} to each other and placed in a reasonably good conducting liquid, an \textit{electr}, such as salt or acid solution, will be capable of generating a flow of \textit{elec}. Such a system is called a \textit{Cell} in Chemistry.
If you use *met* dipped in any old *io* solution they will only be able to produce a *volt* for a very short *ti* as most metals can only react *slo* to produce a small *num* of ions.

A much better cell is produced if each metal is kept in *sepa* beakers, and surrounded by a solution containing *lar* numbers of its own ions.

The *io* bridge is essential to complete the circuit. Elec *do not* flow through the *bri*, instead *io* are able to *mo* from one beaker to the other and com the circuit. Elec *only* flow in the *wi*, from the metal *los* electrons (*oxi*) to the metal ions *gai* electrons (*red*).

**Try to remember:** the *salt bridge* completes the circuit but only ions move through the bridge and solutions.

Even with *lar* numbers of metal *io* /metal *at* present, cells will ‘run out’ when one or other of the *chem* has been used up.

### Non-Metal Reactions

Any *Redox* reaction can be turned into an *electro* cell - even those involving *sol* of *non-m*.

These reactions are likely to be much less familiar and it is essential that the *Electro Series* in your *Data Book* is used properly.
(Though Na\(^+\) ions are also present, they are unable to be converted into Na atoms in aqueous solution and can be ignored).

Since all equations in the Data Book are written as reductions (gain of electrons), one of the equations must be reversed and rewritten as an oxidation (loss of electrons).

**Reduction:** \( \text{I}_2(s) \rightarrow \text{I}^-(aq) \)

**Oxidation:** the equation involving the \( \text{SO}_3^{2-} \) ion must be reversed:

\[ \text{SO}_3^{2-}(aq) + (l) \rightarrow \text{SO}_4^{2-}(aq) + (aq) + (aq) + \]

Our earlier rule can be modified:

‘electrons will flow from the chemical higher in the electrochemical series
to the chemical lower in the series’.

Or, more useful:

‘equation lower in the electrochemical series goes as written (red)
equation higher in the electrochemical series is reversed (oxi)’.
More complicated cells can present a seemingly bewildering choice of reactions.

You will be able to find six possible reactions in the Data Book!!

In reality, one side usually provides an obvious reaction. In this case, the only possible reaction on the left is:

\[
\text{Oxidation:} \quad \text{SO}_3^{2-} (aq) + (l) \rightarrow (aq) + (aq)
\]

This means that the other reaction must be a Reduction and must be lower in the electrochemical series.

Only one reaction on the right will work:

\[
\text{Reduction:} \quad \text{Fe}^{3+} (aq) + \rightarrow (aq)
\]
Q1.

Four cells were made by joining copper, iron, magnesium and zinc to silver. The four cells produced the following voltages 0.5, 0.9, 1.1 and 2.7 V.

Which of the following will be the voltage of the cell containing silver joined to copper?

(You may wish to use page 10 of the data booklet to help you.)

A 0.5 V
B 0.9 V
C 1.1 V
D 2.7 V

Q2.

Zinc displaces copper from copper(II) sulphate solution. The equation for the reaction is:

\[ \text{Zn}^{(s)} + \text{Cu}^{2+} \text{ (aq)} + \text{SO}_4^{2-} \text{ (aq)} \rightarrow \text{Zn}^{2+} \text{ (aq)} + \text{SO}_4^{2-} \text{ (aq)} + \text{Cu}^{(s)} \]

a) Circle the spectator ion in the above equation.

b) Write the ion-electron equation for the oxidation step in this reaction.

You may wish to use the data booklet to help you.

c) The reaction can also be carried out in a cell.

i) Complete the three labels on the diagram.

ii) What is the purpose of the ion bridge?

Q3.

Electricity can be produced using electrochemical cells.

a) Identify the arrangement which would not produce electricity.

b) Identify the two cells which could be used to compare the reactivity of gold and lead.

A technician set up the following cell.

The reaction taking place at electrode B is:

\[ 2\text{Br}^{-} \text{ (aq)} \rightarrow \text{Br}_2 \text{ (l)} + 2\text{e}^{-} \]

a) On the diagram, clearly mark the path and direction of electron flow.

b) Write the ion-electron equation for the reaction taking place at electrode A.

You may wish to use the data booklet to help you.

c) i) Name the piece of apparatus labelled X.

          
ii) Describe the role of the apparatus labelled X.
7.4 Special Cells

Rechargeable Cells
As previously mentioned, even with large numbers of ions present, cells will ‘run out’ when one or other of the chemicals has been used up. Some reactions are, however, reversible which allows the cell to be rechargeable.

When the power pack is switched on an electric current flows. At the same time the appearance of the lead plates change showing that a chemical reaction has taken place.

Charging: Electrical energy to Chemical energy

Discharging: Chemical energy to Electrical energy

This type of cell is called a lead-acid cell and is used mainly in cars, lorries, motorbikes etc.

Common rechargeable battery types

Nickel–cadmium battery (NiCd)
Created by Waldemar Jungner of Sweden in 1899, it used nickel oxide hydroxide and metallic cadmium as electrodes. Cadmium is a toxic element, and was banned for most uses by the European Union in 2004.

Nickel–metal hydride battery (NiMH)
First commercial types were available in 1989. These are now a common consumer and industrial type. The battery has a hydrogen-absorbing alloy for the negative electrode instead of cadmium.
Lithium-ion battery

The technology behind the lithium-ion battery has not yet fully reached maturity. However, the batteries are the type of choice in many consumer electronics and have one of the best energy-to-mass ratios and a very slow loss of charge when not in use.

Lithium-ion polymer battery

These batteries are light in weight and can be made in any shape desired.

**Fuel Cells**

In a fuel cell, **hydr** and **ox** are converted into **wa**, and in the process, **elec** is generated.

A fuel cell has 3 main parts:- an **anode cat**, a **cath catalyst** with a **poly membrane** in between. The **mem** is made of a special **pol** called Nafion that allows **pos cha ions** to pass through.

**Hydr** is fed into the cell and flows over the **an catalyst**. When **hydr mol** hit the **an catalyst**, the H₂ molecule separates into two **hyd ions** (that is, two **protons**) and two **ele** by the following:

\[
\text{oxi} : \quad \text{H}_2 \rightarrow +
\]
The electrons flow through the wire toward the other side of the fuel cell while the hydrogen ions pass through the membrane to get to the other side.

Meanwhile, oxygen molecules have been adsorbed onto the surface of the cathode catalyst which produces single oxygen atoms which react with a hydroxyl ion and an electron to form a hydroxyl group:

\[
O + + → \]

Further reaction with another hydroxyl ion and an electron produces water:

\[
OH + + → \]

The widespread use of hydrogen as a clean-burning non-polluting fuel, and an economic system based on it, is referred to as ‘the hydrogen economy’.

If the method used to make the electricity is non-polluting, then the overall process is almost pollution free.

Sulphur dioxide, unburnt hydrocarbons, soot and carbon oxide are all avoided - some nitrogen oxides, however, will still be produced due to the high temperature of the combustion of hydrogen.

Advantages pollution free - made product of combustion is water. Water is a plentiful and cheap raw material can be distributed easily through pipelines or tankers.

* Portable fuel capable of storing energy made by renewable sources

Disadvantages storing as liquid requires high pressure / low temperature. Potentially explosively produces less energy than the electricity used to make it, but *
In Australia flow cells are used to store the energy from solar cells.

a) What is an electrolyte?

b) The reaction taking place at electrode A when the cell is providing electricity is:

\[ \text{Zn} \rightarrow \text{Zn}^{2+} + 2e^- \]

Name the type of chemical reaction taking place at electrode A.

c) On the diagram, clearly mark the path and direction of electron flow.

d) Name the non-metal, that conducts electricity, which could be used as an electrode.

Methanol fuel cells produce electricity by reacting methanol and oxygen gas from the air. A simplified diagram of a methanol fuel cell is shown below.

a) Write an equation that represents the overall redox reaction.

c) Explain why methanol can be described as a "Green Fuel" despite producing carbon dioxide.

The concentration of ethanol in a person’s breath can be determined by measuring the voltage produced in an electrochemical cell.

Different ethanol vapour concentrations produce different voltages as is shown in the graph below.

a) Write a general statement describing the effect of ethanol vapour on the voltage.

b) Use the graph to predict the volume of ethanol needed to produce a voltage of 20 mV.

c) The ion-electron equations for the reduction and oxidation reactions occurring in the cell are shown below.

\[ \text{O}_2 + 4\text{H}^+ + 4e^- \rightarrow 2\text{H}_2\text{O} \]
\[ \text{CH}_3\text{CH}_2\text{OH} + \text{H}_2\text{O} \rightarrow \text{CH}_3\text{COOH} + 4\text{H}^+ + 4e^- \]

Write the overall redox equation for the reaction taking place.

d) Platinum metal acts as a heterogeneous catalyst in this reaction.

What is meant by a heterogeneous catalyst?
### 7.5 Extracting Metals

The extraction of metals usually involves red **duction** of a metal ox **ide** to form the metal.

The reactions are, therefore, **REDOX**.

The different methods used can also be related to the reac **tivities** of the different metals being extracted.

Some of these methods were first met in Topic 3 last year.

#### Extraction Methods

<table>
<thead>
<tr>
<th>Metal</th>
<th>Method of extraction</th>
</tr>
</thead>
<tbody>
<tr>
<td>potassium</td>
<td>electrolysis</td>
</tr>
<tr>
<td>sodium</td>
<td>electrolysis</td>
</tr>
<tr>
<td>magnesium</td>
<td>electrolysis</td>
</tr>
<tr>
<td>aluminium</td>
<td>electrolysis</td>
</tr>
<tr>
<td>zinc</td>
<td>heat oxide with carbon</td>
</tr>
<tr>
<td>iron</td>
<td>heat oxide with carbon</td>
</tr>
<tr>
<td>tin</td>
<td>heat oxide with carbon</td>
</tr>
<tr>
<td>lead</td>
<td>heat oxide with carbon</td>
</tr>
<tr>
<td>copper</td>
<td>heat oxide with carbon</td>
</tr>
<tr>
<td>mercury</td>
<td>heat or found as element</td>
</tr>
<tr>
<td>silver</td>
<td>heat or found as element</td>
</tr>
<tr>
<td>gold</td>
<td>heat or found as element</td>
</tr>
<tr>
<td>platinum</td>
<td>heat or found as element</td>
</tr>
</tbody>
</table>

More reactive metals like sod **ium** and alum **inium** can only be forced to change back into ato **ms** by the use of large amounts of ene **rgy**.

Reactive metals like ir **on** can be made in fur **naces** at high temp **eratures**, but only if oxy **gen** removers like car **bon** are present.

Less reactive metals like mer **cury** can be made by roasting.

Unreactive metals like go **ld** and sil **ver** are found pure in the Earth’s crust.

Suitable ox **re** are all substances eager to join with ox **ides** atoms, or mo **xide** oxygen atoms, to form new compounds and include:

$$\text{Ca} \quad \text{(coke)} \quad \text{C} \quad \rightarrow \quad \text{(g)}$$

$$\text{Hy} \quad \rightarrow \quad \text{(l)}$$

$$\text{Me} \quad \rightarrow \quad \text{(g)} + \text{(l)}$$

$$\text{Carbon mo} \quad \rightarrow \quad \text{(g)}$$

They are more usually referred to as Red **oxides** Agents - cause red.
Iron is made in a **Blast Furnace**. This is a very tall tower lined with fireproof bricks.

Hot air is blasted in at the bottom of the furnace. The coke (carbon) reacts with oxygen, burns, to produce carbon dioxide and heats up the furnace.

As the hot carbon dioxide rises it reacts with hot coke (carbon) to produce the poisonous gas carbon monoxide.

Car monoxide is an excellent reducing agent and it reacts with the iron oxide to produce iron metal.

The molten iron sinks to the bottom of the furnace where it can be drained off.

There are in fact **three** separate reactions taking place in a blast furnace, though overall the reaction is

\[
\text{iron(III) oxide} + \text{carbon} \rightarrow \text{iron} + \text{carbon dioxide}
\]

**Reaction ①**
\[
\text{carbon} + \text{oxygen} \rightarrow \text{carbon dioxide}
\]

**Reaction ②**
\[
\text{carbon dioxide} + \text{carbon} \rightarrow \text{carbon monoxide}
\]

**Reaction ③**
\[
\text{iron(III) oxide} + \text{carbon monoxide} \rightarrow \text{carbon dioxide} + \text{iron}
\]

Overall, the following things have happened:

Iron ions have been changed back into atoms

\[
\text{C} + e^- \rightarrow (\text{Red})
\]

Carbon atoms have joined with oxygen
Extracting Aluminium

The main requirement in the extraction of alum from alumina (aluminium oxide) is a very very large supply of electricity. In fact, an aluminium smelter was ‘coaxed’ to Invergordon in Scotland by the fact that a hydro-electric power station was specially built for them.

This is an example of Electrolysis - where electricity is used to split apart a compound - as well as being REDOX.

Being a reactive metal, alum cannot be produced from a solution - it is easier to make the water molecules break apart. Unfortunately, alumina has a very high melting point (2318 °C) making it very expensive to try and melt it directly. There is, however, another ionic compound, called cryolite, which melts at a much lower temperature and can then be used to ‘dissolve’ the alumina. Even then, high temperatures are still needed, and the alum forms as a liquid.

The reactions:

\[
\begin{align*}
2O_2^- & \rightarrow O_2 + 4e^- \quad (\text{Oxidation}) \\
3e^- & \rightarrow Al \quad (\text{Reduction})
\end{align*}
\]

Comparing Ores

Rocks that contain significant amounts of metal compounds suitable for extraction are generally called ores.

Iron ores are often red in colour for the same reason that our blood is red - haemoglobin molecules contain iron.

Typical iron compounds include iron (II) sulphide, iron (II) oxide and iron (III) oxide. Percentage compositions (see Calculations Booklet) provide an easy way of comparing iron contents.

FeS

\[
\begin{align*}
1 \times \text{Fe} &= 1 \times 56 = 56 \\
1 \times \text{S} &= 1 \times 32 = 32 \\
\text{formula mass} &= 88 \\
\%	ext{Fe} &= \frac{56}{88} \times 100 = 64\%
\end{align*}
\]

FeO

\[
\begin{align*}
1 \times \text{Fe} &= 1 \times = \text{x Fe} = 2 \times = \\
1 \times \text{O} &= 1 \times = \text{x O} = 3 \times = \\
\text{formula mass} &= \text{formula mass} = \\
\%	ext{Fe} &= \frac{x}{100} = % \quad \text{and} \quad \% \text{O} = \frac{3x}{100} = %
\end{align*}
\]

Fe₂O₃

%Fe = 56/88 x 100 = 64% / x 100 = %

KHS Mar 2014
A metal can be extracted from its ore by heating the ore with carbon but not by heating the ore on its own.

The position of the metal in the reactivity series is most likely to be between (You may wish to use page 10 of the data booklet to help you.)

- A  zinc and magnesium
- B  magnesium and potassium
- C  zinc and copper
- D  copper and gold

Metals can be extracted from metal compounds by heat alone, heating with carbon or by electrolysis.

a) Name the type of chemical reaction which takes place when a metal is extracted from its compound.

b) In an experiment, a solution of copper(II) chloride was electrolysed.

i) Complete the table by adding the charge for each electrode.

<table>
<thead>
<tr>
<th>Observation at electrode</th>
<th>Observation at electrode</th>
</tr>
</thead>
<tbody>
<tr>
<td>bubbles of gas</td>
<td>brown solid formed</td>
</tr>
</tbody>
</table>

ii) How could the gas be identified?

Some metals can be obtained from their metal oxides by heat alone.

Which of the following oxides would produce a metal when heated?

- A  calcium oxide
- B  copper oxide
- C  zinc oxide
- D  silver oxide

Aluminium can be extracted from aluminium oxide by

- A  heating alone
- B  heating with carbon
- C  heating with carbon monoxide
- D  electrolysis

Metals can be extracted from their ores by different methods.

a) Place the following methods in the correct space in the table. You may wish to use the data booklet to help you.

reacting with carbon electrolysis heat alone

<table>
<thead>
<tr>
<th>Metal</th>
<th>Method</th>
</tr>
</thead>
<tbody>
<tr>
<td>mercury</td>
<td></td>
</tr>
<tr>
<td>iron</td>
<td></td>
</tr>
<tr>
<td>magnesium</td>
<td></td>
</tr>
</tbody>
</table>

b) Mercury can be extracted from the ore cinnabar, HgS.

i) Calculate the percentage by mass of mercury in cinnabar.

 ii) Write the formula for the mercury ion in cinnabar

 iii) Write the ion-electron equation for the reduction of the mercury ion in cinnabar.
N5 Knowledge Met in this Section

Definitions

- When a substance reacts by **gaining oxygen** we call it **oxidation**
  \[ \text{e.g. } Mg + O_2 \rightarrow MgO \]
- When a substance reacts by **losing electrons** we call it **oxidation** (OIL)
  \[ \text{e.g. } Mg \rightarrow Mg^{2+} + 2e \]
- When a substance reacts by **losing oxygen** we call it **reduction**
  \[ \text{e.g. } HgO \rightarrow Hg + O_2 \]
- When a substance reacts by **gaining electrons** we call it **reduction** (RIG)
  \[ \text{e.g. } Hg^{2+} + 2e \rightarrow Hg \]

Chemical properties of metals (**Revision of S3**)

- *Metals can be placed in a **reactivity series** by observing their reactions*
- *Potassium, sodium and lithium are stored in oil because they react quickly with oxygen and water vapour in the air.*
- *These metals are too reactive to risk in reactions with acids*
- *Aluminium is slow to react with acids because it has a protective coating of oxide.*

Chemical reactions of metals (**Revision of S3**)

- *Metal + oxygen → metal oxide*
- *Metal + water → metal hydroxide (alkali) + hydrogen*
- *Metal + hydrochloric acid → metal chloride + hydrogen*
- A **displacement** reaction is when a more reactive metal can take the place of a less reactive metal.
- The more reactive metal dissolves into the solution, the less reactive metal is forced out of the solution.
- The more reactive metal, doing the displacement, loses electrons and forms ions
  \[ \text{e.g. } Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e \]
  \[ \text{oxidation} = \text{loss of electrons} \]
- The ions of the less reactive metal, being displaced, gain electrons and form atoms
  \[ \text{e.g. } Cu^{2+}_{(aq)} + 2e \rightarrow Cu_{(s)} \]
  \[ \text{reduction} = \text{gain of electrons} \]
- Displacement reactions are redox processes since they involve both loss and gain of electrons
  \[ \text{e.g. } Cu^{2+}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)} \]
**Batteries and Cells**

- A **chemical cell** converts chemical energy into electrical energy
- A **battery** is two or more cells joined together
- Chemicals are used up when cells produce electricity
- Some cells are **rechargeable**, e.g. a ‘lead-acid’ car battery
- In a cell, one substance loses electrons whilst another substance gains electrons
- An electrolyte is often needed to complete the circuit in a cell
- Compared to mains electricity, batteries are safer and portable but more expensive and make greater use of resources such as metals etc.
- Electricity can be produced in a cell by connecting two different metals in solutions of their own ions

![Diagram of a battery showing magnesium and zinc separated by an electrolyte](image)

- In the cell above:
  
  \[
  \begin{align*}
  Mg_{(s)} & \rightarrow Mg^{2+}_{(aq)} + 2e^- \\[0.5cm]
  Zn^{2+}_{(aq)} + 2e^- & \rightarrow Zn_{(s)}
  \end{align*}
  \]
  (oxidation) (reduction)

- All cells of this type must have an **ion bridge** to complete the circuit

**Metals and the Electrochemical series (ECS)**

- The **electrochemical series** is very similar to the reactivity series but lists metals in order of their ability to push electrons round a circuit
- A cell can be two **different** metals connected by wires and an electrolyte to complete the circuit
- The further apart two metals are in the **ECS**, the greater is the cell voltage
- In a cell, electrons flow **from the metal higher** in the **ECS** to the one lower down
- Electrons only travel through the wires connecting the two metals
Cells without metals

- Carbon rods can be used to make contact with chemicals dissolved in solutions

![Diagram of cells with carbon rods and solutions]

- In the cell above:
  \[ H_2O(l) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^- \] (oxidation)
  \[ I_2(aq) + 2e^- \rightarrow 2I^-(aq) \] (reduction)

- Electrons flow from the substance higher in the ECS to the one lower down

Extracting Metals (*Revision of S3)

- *Only unreactive metals like gold and silver are found in the Earth’s crust
- *An ore is a naturally occurring compound of a metal
- *More reactive metals must be extracted from their ores
- *Many ores are oxides
- *The higher a metal is in the reactivity series, the more stable are its compounds
- Iron is extracted from iron ore in a blast furnace in which carbon monoxide reduces iron oxide to iron:
  \[ \text{iron oxide} + \text{carbon monoxide} \rightarrow \text{iron} + \text{carbon dioxide} \]

Reactivity and ease of metal extraction

- Reactive metals hold on to oxygen more strongly than less reactive metals
- Heating alone is sufficient to release the metal from oxides of unreactive metals
- Heating with oxygen removers, like carbon or carbon monoxide, releases the metal from the oxides of moderately reactive metals
- Oxygen removers form stronger bonds with oxygen than these metals do
- Only electricity can help release the metal from oxides of reactive metals
- The extraction of a metal from its ore is an example of reduction because the most important process taking place is metal ions gaining electrons to form metal atoms
CONsolidation Questions

Q1. Int2

In the cell shown electrons flow through

A the solution from copper to tin
B the solution from tin to copper
C the wires from copper to tin
D the wires from tin to copper.

Q2. Int2

When a metal element reacts to form a compound the metal is

A displaced
B oxidised
C precipitated
D reduced

Q3. Int2

Which of the following metals can be obtained from its ore by heating with carbon monoxide?

(You may wish to use the data booklet to help you.)

A aluminium
B calcium
C magnesium
D nickel

Q4. Int2

Which metal will displace zinc from a solution of zinc sulphate?

A iron
B magnesium
C silver
D tin

Q5. Int2

Which of the following metals, when linked to zinc, would give the highest cell voltage?

(You may wish to use the data booklet to help you.)

A copper
B iron
C magnesium
D tin

Q6. Int2

In which of the following test tubes will a reaction occur?

A

B

C

D
Q1. Which line in the table is correct for the above cell?

<table>
<thead>
<tr>
<th>Zinc electrode</th>
<th>Copper electrode</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>mass increases</td>
</tr>
<tr>
<td>B</td>
<td>mass increases</td>
</tr>
<tr>
<td>C</td>
<td>mass decreases</td>
</tr>
<tr>
<td>D</td>
<td>mass decreases</td>
</tr>
</tbody>
</table>

Q2. Cells can be made in which both metals and non-metals are used.

a) The ion-electron equation for the reaction taking place at the carbon electrode is:

\[ I_{(aq)} + 2e^- \rightarrow 2I^-_{(aq)} \]

On the diagram clearly mark the path and direction of electron flow.

b) With the help of your data book, write the ion-electron equation for the reaction at the other electrode.

c) What property of carbon makes it suitable for use as an electrode?

Q3. Which of the following metals would react with zinc chloride solution?

You may wish to use page 10 of the data booklet to help you.

A. copper
B. gold
C. iron
D. magnesium

Q4. Which of the following shows the metals in order of increasing reactivity?

A. X Y Z
B. Y X X
C. Z X Y
D. Z Y X

Q5. A student carried out some experiments with four metals and their oxides. The results are shown in the table.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Reaction with cold water</th>
<th>Reaction with dilute acid</th>
<th>Effect of heat on metal oxide</th>
</tr>
</thead>
<tbody>
<tr>
<td>W</td>
<td>no reaction</td>
<td>no reaction</td>
<td>no reaction</td>
</tr>
<tr>
<td>X</td>
<td>no reaction</td>
<td>gas produced</td>
<td>no reaction</td>
</tr>
<tr>
<td>Y</td>
<td>gas produced</td>
<td>gas produced</td>
<td>no reaction</td>
</tr>
<tr>
<td>Z</td>
<td>no reaction</td>
<td>no reaction</td>
<td>metal produced</td>
</tr>
</tbody>
</table>

a) Place the four metals in order of reactivity (most reactive first).

b) Name the gas produced when metal Y reacts with cold water.

c) Suggest names for metals Y and Z.

metal Y ___________ metal Z ___________
CONSOLIDATION QUESTIONS

Q1.

Titanium metal is used to make dental braces.

Titanium is extracted from its ore in the Kroll process. One step in this process involves the displacement of titanium chloride by sodium metal.

The equation is shown.

\[ 4\text{Na} + \text{TiCl}_4 \rightarrow 4\text{NaCl} + \text{Ti} \]

a) What does this method of extraction tell you about the reactivity of titanium metal compared to sodium metal?

b) During the displacement, sodium atoms, Na, form sodium ions, \(\text{Na}^+\).

Write the ion-electron equation for this change.

c) The displacement reaction is carried out in an atmosphere of the noble gas, argon.

Suggest why an argon atmosphere is used.

d) The formula of titanium chloride is \(\text{TiCl}_4\). Use this to work out the charge on the titanium ion and then write the ion-electron equation for the formation of titanium atoms.

Q2.

Aluminium is extracted from the ore bauxite.

a) Circle the correct phrase to complete the sentence.

Aluminium is extracted from its ore

\[ \{\text{by heating with carbon}\} \]
\[ \{\text{by heating alone}\} \]
\[ \{\text{by electrolysis}\} \]

b) Aluminium can be mixed with other metals to make a magnet.

What term is used to describe a mixture of metals?

c) The composition of a 250 g magnet is shown.

<table>
<thead>
<tr>
<th>Metal</th>
<th>aluminium</th>
<th>nickel</th>
<th>cobalt</th>
<th>copper</th>
<th>titanium</th>
<th>iron</th>
</tr>
</thead>
<tbody>
<tr>
<td>% by mass</td>
<td>10</td>
<td>25</td>
<td>20</td>
<td>4</td>
<td>1</td>
<td>49</td>
</tr>
</tbody>
</table>

i) Calculate the mass, in grams, of aluminium in the magnet.

Show your working clearly.

\[ \text{mass of aluminium} = \text{total mass} \times \text{percentage of aluminium} \]

\[ \text{mass of aluminium} = 250 \, \text{g} \times 0.10 \]

\[ \text{mass of aluminium} = 25 \, \text{g} \]

\[ \text{mass of aluminium} = 25 \, \text{g} \]

ii) Using your answer to c) i), calculate the number of moles of aluminium in the magnet.

Show your working clearly.

\[ \text{number of moles of aluminium} = \frac{\text{mass of aluminium}}{\text{molar mass of aluminium}} \]

\[ \text{number of moles of aluminium} = \frac{25 \, \text{g}}{27 \, \text{g/mol}} \]

\[ \text{number of moles of aluminium} = 0.926 \, \text{mol} \]

Q3.

The ion-electron equation for the oxidation and reduction steps in the reaction between magnesium and silver(I) ions are:

\[ \text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^- \]

\[ \text{Ag}^+ + \text{e}^- \rightarrow \text{Ag} \]

The overall redox equation is

A \[ \text{Mg} + 2\text{Ag}^+ \rightarrow \text{Mg}^{2+} + 2\text{Ag} \]

B \[ \text{Mg} + \text{Ag}^+ \rightarrow \text{Mg}^{2+} + \text{Ag} \]

C \[ \text{Mg} + \text{Ag}^+ + \text{e}^- \rightarrow \text{Mg}^{2+} + \text{Ag} + 2\text{e}^- \]

D \[ \text{Mg} + 2\text{Ag} \rightarrow \text{Mg}^{2+} + 2\text{Ag}^+ \]
A student’s report is shown for the “Reaction of metals with oxygen”.

Title: Reactions of Metals with Oxygen  
Date: 15/11/12

Aim: The apparatus required to carry out the experiment was collected and assembled as shown.

Procedure: The reaction was carried out as shown.

Results: The reaction was successful and the products were collected and assembled as shown.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Observations</th>
</tr>
</thead>
<tbody>
<tr>
<td>zinc</td>
<td>moderately fast reaction</td>
</tr>
<tr>
<td>magnesium</td>
<td></td>
</tr>
<tr>
<td>copper</td>
<td></td>
</tr>
</tbody>
</table>

a) State the aim of the experiment.

b) Why is potassium permanganate used in this experiment?

c) Complete the table to show the observations for magnesium and copper.

d) For safety reasons this experiment would not be carried out with potassium metal. Suggest a reason for this.

The diagram shows how an object can be coated with silver.

The following reactions take place at the electrodes.

Negative electrode: $\text{Ag}^{+} (aq) + \text{e}^- \rightarrow \text{Ag} (s)$

Positive electrode: $\text{Ag} (s) \rightarrow \text{Ag}^+ (aq) + \text{e}^-$

A) Ions flow through the solution.
B) Silver ions move towards the silver electrode.
C) The process is an example of galvanising.
D) The mass of the silver electrode decreases.
E) Reduction occurs at the silver electrode.

Identify the two correct statements.

Q3.

Galena is an ore containing lead sulphide, PbS.

a) What is the charge on this lead ion?

b) Calculate the percentage by mass of lead in galena, PbS.

Most metals have to be extracted from their ores.

c) Name the metal extracted in a Blast furnace.

d) Place the following metals in the correct space in the table.

You may wish to use the data booklet to help you.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Method of extraction</th>
</tr>
</thead>
<tbody>
<tr>
<td>copper, mercury, aluminium</td>
<td>using heat alone</td>
</tr>
<tr>
<td></td>
<td>electrolysis of molten ore</td>
</tr>
<tr>
<td></td>
<td>heating with carbon</td>
</tr>
</tbody>
</table>