Revision Pack Topic 4 Reactions

Topic area	R/A/G
<u>Metal oxides</u> Metals react with oxygen to produce metal oxides. The reactions are oxidation reactions because the metals gain oxygen. Students should be able to explain reduction and oxidation in terms of loss or gain of oxygen.	
<u>Reactivity series</u> When metals react with other substances the metal atoms form positive ions. The reactivity of a metal is related to its tendency to form positive ions. Metals can be arranged in order of their reactivity in a reactivity series. The metals potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper can be put in order of their reactivity from their reactions with water and dilute acids. The non-metals hydrogen and carbon are often included in the reactivity series. A more reactive metal can displace a less reactive metal from a compound. Students should be able to: • recall and describe the reactions, if any, of potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper with water or dilute acids and where appropriate, to place these metals in order of reactivity • explain how the reactivity of metals with water or dilute acids is related to the tendency of the metal to form its positive ion • deduce an order of reactivity of metals based on experimental results.	
<u>Extracting metals</u> Unreactive metals such as gold are found in the Earth as the metal itself but most metals are found as compounds that require chemical reactions to extract the metal. Metals less reactive than carbon can be extracted from their oxides by reduction with carbon. Reduction involves the loss of oxygen. Students should be able to: • interpret or evaluate specific metal extraction processes when given appropriate information • identify the substances which are oxidised or reduced in terms of gain or loss of oxygen.	
<u>Oxidation and reduction</u> **(Higher only)** Oxidation is the loss of electrons and reduction is the gain of	

electrons. Student should be able to: • write ionic equations for displacement reactions • identify in a given reaction, symbol equation or half equation which species are oxidised and which are reduced.	
<u>Reactions of acids</u> Acids react with some metals to produce salts and hydrogen. **(Higher only) Students should be able to:** **• explain in terms of gain or loss of electrons, that these are redox reactions** **• identify which species are oxidised and which are reduced in given chemical equations.** Acids are neutralised by alkalis (eg soluble metal hydroxides) and bases (eg insoluble metal hydroxides and metal oxides) to produce salts and water, and by metal carbonates to produce salts, water and carbon dioxide. The particular salt produced in any reaction between an acid and a base or alkali depends on: • the acid used (hydrochloric acid produces chlorides, nitric acid produces nitrates, sulfuric acid produces sulfates) • the positive ions in the base, alkali or carbonate. Students should be able to: • predict products from given reactants • use the formulae of common ions to deduce the formulae of salts.	
<u>Producing samples of salts</u> Soluble salts can be made from acids by reacting them with solid insoluble substances, such as metals, metal oxides, hydroxides or carbonates. The solid is added to the acid until no more reacts and the excess solid is filtered off to produce a solution of the salt. Salt solutions can be crystallised to produce solid salts. Students should be able to describe how to make pure, dry samples of named soluble salts from information provided. ***Required practical activity 8: preparation of a pure, dry sample of a soluble salt from an insoluble oxide or carbonate***	
<u>The pH scale</u> Acids produce hydrogen ions (H+) in aqueous solutions. Aqueous solutions of alkalis contain hydroxide ions (OH–).	

Revision Pack Topic 4 Reactions

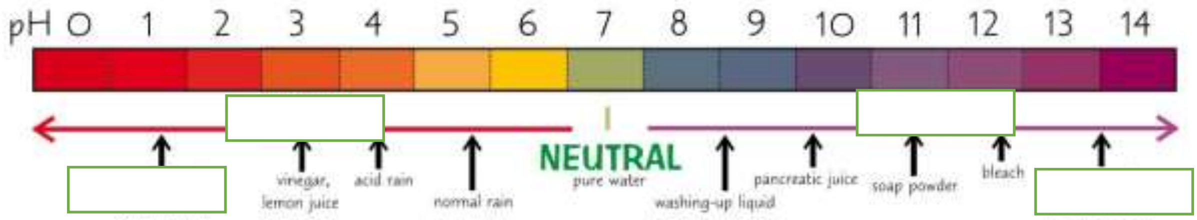
The pH scale, from 0 to 14, is a measure of the acidity or alkalinity of a solution, and can be measured using universal indicator or a pH probe. A solution with pH 7 is neutral. Aqueous solutions of acids have pH values of less than 7 and aqueous solutions of alkalis have pH values greater than 7. In neutralisation reactions between an acid and an alkali, hydrogen ions react with hydroxide ions to produce water. This reaction can be represented by the equation: $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$ Students should be able to: • describe the use of universal indicator or a wide range indicator to measure the approximate pH of a solution • use the pH scale to identify acidic or alkaline solutions. <u>(Higher only)</u> **A strong acid is completely ionised in aqueous solution. Examples of strong acids are hydrochloric, nitric and sulfuric acids. A weak acid is only partially ionised in aqueous solution. Examples of weak acids are ethanoic, citric and carbonic acids.** **For a given concentration of aqueous solutions, the stronger an acid, the lower the pH. As the pH decreases by one unit, the hydrogen ion concentration of the solution increases by a factor of 10.** **Students should be able to:** **• use and explain the terms dilute and concentrated (in terms of amount of substance), and weak and strong (in terms of the degree of ionisation) in relation to acids** **• describe neutrality and relative acidity in terms of the effect on hydrogen ion concentration and the numerical value of pH (whole numbers only).**	
<u>Electrolysis</u> When an ionic compound is melted or dissolved in water, the ions are free to move about within the liquid or solution. These liquids and solutions are able to conduct electricity and are called electrolytes. Passing an electric current through electrolytes causes the ions to move to the electrodes. Positively charged ions move to the negative electrode (the cathode), and negatively charged ions move	

to the positive electrode (the anode). Ions are discharged at the electrodes producing elements. This process is called electrolysis. **(Higher only)** **Students should be able to write half equations for the reactions occurring at the electrodes during electrolysis, and may be required to complete and balance supplied half equations.** When a simple ionic compound (eg lead bromide) is electrolysed in the molten state using inert electrodes, the metal (lead) is produced at the cathode and the non-metal (bromine) is produced at the anode. Students should be able to predict the products of the electrolysis of binary ionic compounds in the molten state.	
<u>Extracting metals by electrolysis</u> Metals can be extracted from molten compounds using electrolysis. Electrolysis is used if the metal is too reactive to be extracted by reduction with carbon or if the metal reacts with carbon. Large amounts of energy are used in the extraction process to melt the compounds and to produce the electrical current. Aluminium is manufactured by the electrolysis of a molten mixture of aluminium oxide and cryolite using carbon as the positive electrode (anode). Students should be able to: • explain why a mixture is used as the electrolyte • explain why the positive electrode must be continually replaced.	
<u>Electrolysis of aqueous solutions</u> The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved. At the negative electrode (cathode), hydrogen is produced if the metal is more reactive than hydrogen. At the positive electrode (anode), oxygen is produced unless the solution contains halide ions when the halogen is produced. This happens because in the aqueous solution water molecules break down producing hydrogen ions and hydroxide ions that are discharged. Students should be able to predict the products of the electrolysis of aqueous solutions containing a single ionic compound. **(Higher only)** **During electrolysis, at the cathode (negative electrode), positively**	

charged ions gain electrons and so the reactions are reductions. At the anode (positive electrode), negatively charged ions lose electrons and so the reactions are oxidations. Reactions at electrodes can be represented by half equations, for example: $2H+ + 2e- \rightarrow H_2$ and $4OH- \rightarrow O_2 + 2H_2O + 4e$ or $4OH- - 4e- \rightarrow O2 + 2H_2O$	

Revision Pack Topic 4 Reactions

The pH scale



Add the following labels to the pH scale: **Acids, alkalis, stomach acid, drain cleaner**

Name the indicator that gives the colours on the pH scale: ________________

State an alternative way of measuring the pH of a solution and describe an advantage of it:

____

____

Use the 2 equations below to explain what happens when equal volumes of acid and alkali are mixed together.

acid + base → salt + water | $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

Key words: Neutralisation, salt, water, OH-, H+, ion

____

____

Give an example of a strong acid and a weak acid. Describe the difference between them.

____

____

Reactions of acids

Use the example below to practise writing more equations for the reactions of acids:

Acid + Metal Oxide → Salt + Water

Acid + Metal Hydroxide → Salt + Water

The salt that's produced depends upon the acid and the metal ion in the oxide or hydroxide:

hydrochloric acid	+	copper oxide	→	copper chloride	+	water
$2HCl$	+	CuO	→	$CuCl_2$	+	H_2O

Acid + Metal Carbonate → Salt + Water + Carbon Dioxide

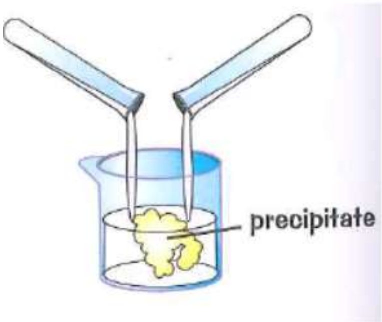
Acid + Metal → Salt + Hydrogen

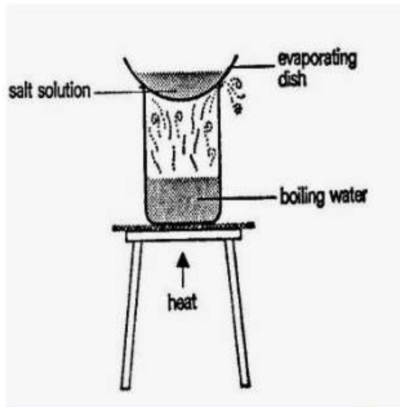
Revision Pack Topic 4 Reactions

Sodium oxide	+	sulphuric acid	→		+	water
Aluminium oxide	+		→	aluminium nitrate	+	water
Potassium hydroxide	+	nitric acid	→		+	
Calcium carbonate	+	nitric acid	→			
	+		→	lead chloride	+	hydrogen

You could collect a sample of the salt produced in the reactions above using the steps below.

Lead Nitrate + Sodium Chloride →





Use the diagrams to write a method for the preparation of a dry sample of lead chloride.

Success criteria

- Key words included
- Step-by-step instructions
- Chemical names of reactants and products given
- Word equation written

Keywords:
Solubility,
Insoluble,
Precipitate,
Barium Sulphate,
Filter paper
Evaporating dish
Soluble

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Revision Pack Topic 4 Reactions

Reactivity of metals

The table below shows the order of the reactivity of metals, and gives information about how the metals react with water and acid.

	Metal	Reaction with water or steam	Reaction with dilute acids
Reactive metals	Potassium Sodium Calcium	React with cold water	Explosive reaction Violent reaction
Fairly reactive metals	Magnesium CARBON Zinc Iron	React with steam	Moderately fast reaction Slow reaction
Unreactive metals	Lead Copper Silver	No reaction with water and steam	No reaction

Explain how the observations in the table show that calcium is more reactive than iron.

____

____

____

____

A student measures the temperature change when 4 different metals react with acid:

Metal	Temperature change (°C)
Gold	3
Magnesium	15
Lead	8
Lithium	11

Explain what the observations show about the reactivity of the 4 metals:

____

____

____

____

____

Revision Pack Topic 4 Reactions

Extraction of metals

An **ore** is a type of rock that contains enough of a __________ compound to make it worthwile extracting the metal from it – e.g. bauxite contains aluminium oxide.

Metals below carbon in the reactivity series can be extracted from their ore by reacting them with carbon, for example:

The extraction of iron from iron oxide:

$2Fe_2O_3$	+	$3C$	$\rightarrow$	$4Fe$	+	$3CO_2$
iron(III) oxide	+	carbon	$\rightarrow$	iron	+	carbon dioxide

Use the information on the previous page to explain why iron can be extracted by reacting iron oxide with carbon, but magnesium cannot be extracted from magnesium oxide in the same way:

____
____
____
____

Oxidation = Gain of Oxygen

E.g. magnesium is oxidised to make magnesium oxide.

$$2Mg + O_2 \rightarrow 2MgO$$

Reduction = Loss of Oxygen

E.g. copper oxide is reduced to copper.

$$2CuO + C \rightarrow 2Cu + CO_2$$

Use this information to identify and the equation above to identify something that is reduced and something that is oxidized in the extraction of iron from iron oxide:

Oxidised: ____________________________________

Reduced: ____________________________________

Explain why it can be said that the extraction of iron takes place in a **redox** reaction.

____
____

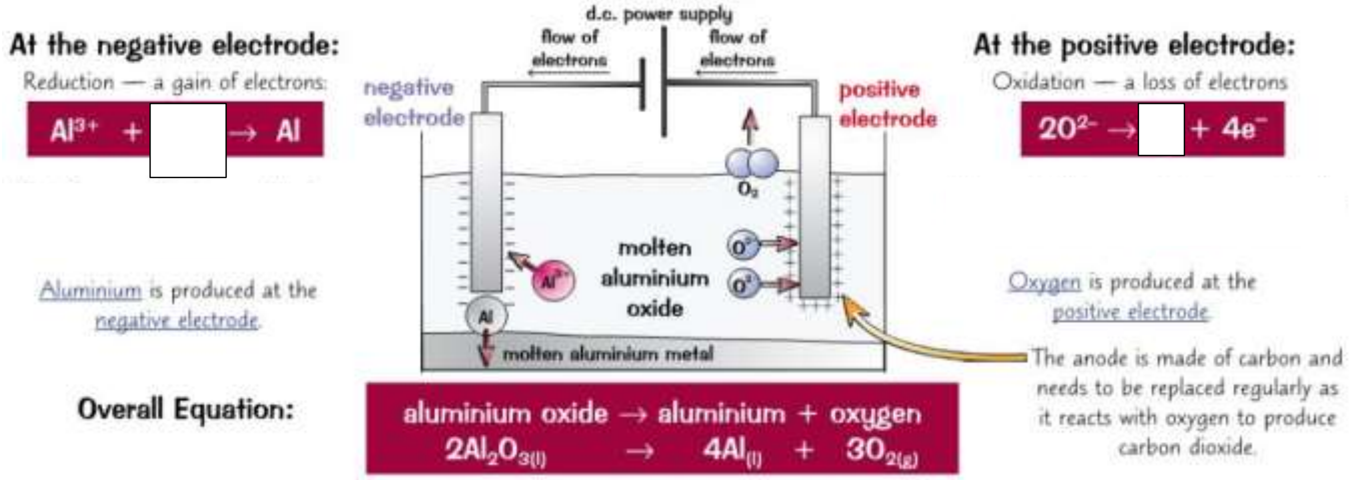
Suggest why gold can be found 'native' (not in a compound)

____
____

Revision Pack Topic 4 Reactions

·Positive
·Anode
·Negative
·Is
·Cathode

Metals that are more reactive than carbon have to be extracted from their ore using **electrolysis**.



Join the key words to their definitions:

Electrolysis	A molten or dissolved ionic compound.
Electrolyte	A positive ion, attracted to the cathode.
Cathode	A compound that is mixed with aluminium oxide to lower its melting point.
Anode	Splitting up an ionic compound using an electric current.
Cryolite	A negative ion, attracted to the anode.
Anion	The negative electrode.
Cation	The positive electron.

Explain why aluminium is produced at the negative electrode, whilst oxygen is produced at the positive electrode.

____
____
____
____

Explain why molten aluminium oxide can conduct electricity, whereas molten silicon dioxide cannot. (Clue – think about the bonding in each compound!)

____
____
____

Revision Pack Topic 4 Reactions

Electrolysis of aqueous solutions

When we carry out electrolysis with ionic compounds dissolved in water, we have to use different rules.

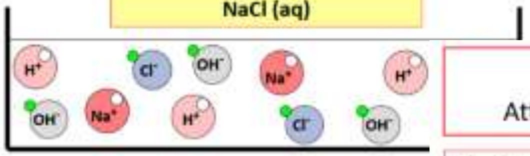
Here, water molecules break up into **HYDROGEN IONS, H^+** and **HYDROXIDE IONS OH^-**

$$H_2O \rightarrow H^+ + OH^-$$

So, in an ionic solution (eg sodium chloride solution), there will be FOUR types of ion present:

TWO from the ionic compound and TWO from the water (H^+ + OH^-)

SODIUM CHLORIDE SOLUTION
NaCl (aq)



+ ANODE Attracts – ions ('Anions')	**- CATHODE** Attracts + ions ('Cations')
If – ions are HALOGENS ie **chloride Cl^-** **bromide Br^-** **iodide I^-** the **HALOGEN** is produced.	If + ions (metals) are MORE REACTIVE than hydrogen **K, Na, Ca, Mg, Zn, Fe** Then **HYDROGEN** is produced
If – ions are NOT HALOGENS Eg sulphate SO_4^{2-}, nitrate NO^{3-} carbonate CO_3^{2-} OXYGEN is produced.	If + ions (metals) are LESS REACTIVE than hydrogen **Cu, Ag, Au** Then the METAL is produced

Complete the table below to show the products at each electrode when the following compounds are dissolved:

(REACTIVITY: K^+ Na^+ Ca^{2+} Mg^{2+} Al^{3+} Zn^{2+} Fe^{3+} H^+ Cu^{2+} Ag^+ Au^{3+})

Compound	State	Ions	Cathode (-)	Anode (+)
potassium chloride	molten	K^+ Cl^-	potassium	chlorine
aluminium oxide	molten			
copper chloride	solution			
sodium bromide	solution			
silver nitrate	solution			
potassium chloride	solution			
zinc sulphate	solution			

(REACTIVITY: K^+ Na^+ Ca^{2+} Mg^{2+} Al^{3+} Zn^{2+} Fe^{3+} H^+ Cu^{2+} Ag^+ Au^{3+})

Revision Pack Topic 4 Reactions

Exam questions

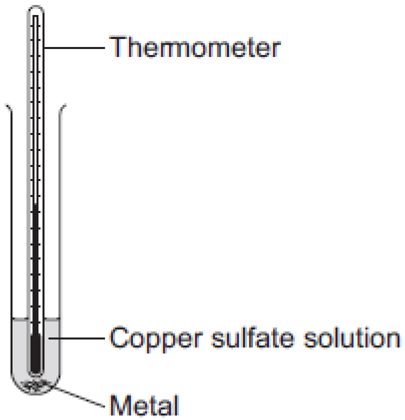
Q1.A student investigated displacement reactions of metals.

The student added different metals to copper sulfate solution and measured the temperature change.

The more reactive the metal is compared with copper, the bigger the temperature change.

The apparatus the student used is shown in **Figure 1**.

Figure 1



(a) State **three** variables that the student must control to make his investigation a fair test.

1...

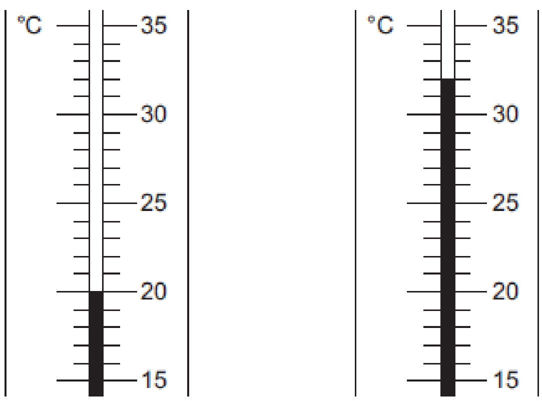
2...

3...

(3)

(b) **Figure 2** shows the thermometer in one experiment before and after the student added a metal to the copper sulfate solution.

Before adding metal **After adding metal**



Use **Figure 2** to complete **Table 1**.

Temperature before adding metal in °C	....
Temperature after adding metal in °C	....
Change in temperature in °C	....

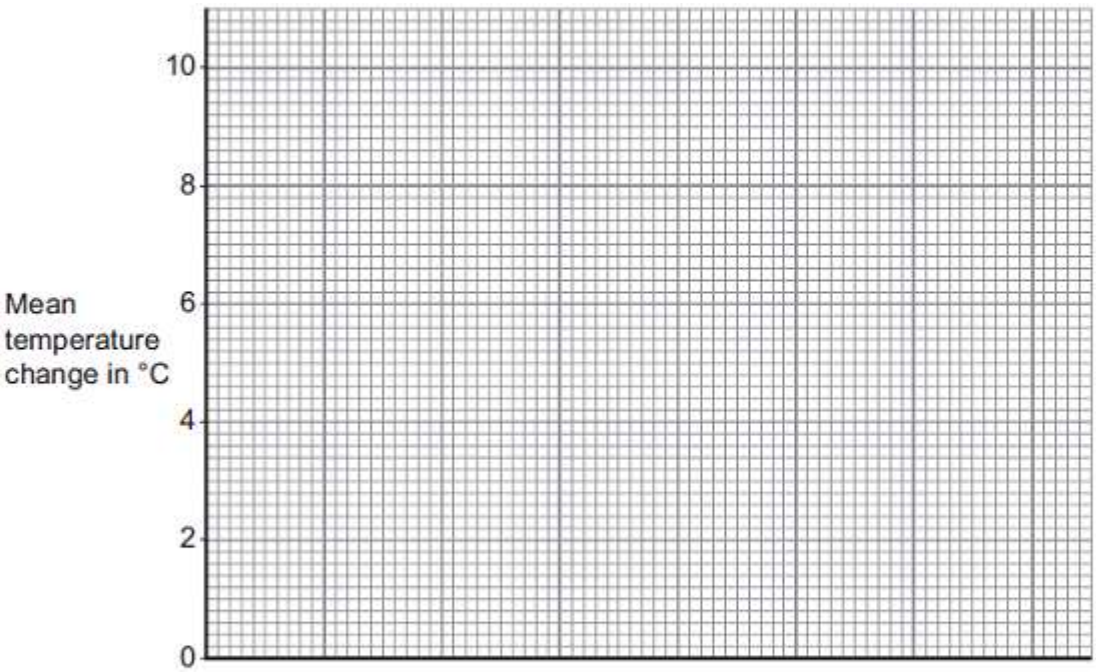
(3)

(c) The student repeated the experiment three times with each metal.

Table 2 shows the mean temperature change for each metal.

Metal	**Mean temperature change in °C**
Cobalt	4.5
Gold	0.0
Magnesium	10.0
Nickel	3.0
Silver	0.0
Tin	1.5

(i) On **Figure 3**, draw a bar chart to show the results.



(ii) Why is a line graph **not** a suitable way of showing the results?

....

....

(1)

(iii) Use the results to work out which metal is the most reactive.

Give a reason for your answer.

Most reactive metal ....

Reason ....

....

(2)

(iv) Explain why there was no temperature change when silver metal was added to the copper sulfate solution.

....

....

....

....

(2)

(v) It is **not** possible to put all six metals in order of reactivity using these results.

Suggest how you could change the experiment to be able to put all six metals into order of reactivity.

....

....

....

....

....

(2)
(Total 16 marks)

Q2. The pH scale is a measure of the acidity or alkalinity of a solution.

(a) Draw one line from each solution to the pH value of the solution.

Revision Pack Topic 4 Reactions

Solution	pH value of the solution
	5
Acid	7
	9
Neutral	11
	13

(2)

(b) Which ion in aqueous solution causes acidity?

H^+ ☐

Na^+ ☐

O^{2-} ☐

OH^- ☐

(1)

(c) When sulfuric acid is added to sodium hydroxide a reaction occurs to produce two products.

The equation is:

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

How many elements are in the formula H_2SO_4?

3 ☐

4 ☐

6 ☐

7 ☐

(1)

(d) What is this type of reaction?

Decomposition ☐

Displacement ☐

Neutralisation ☐

Reduction ☐

(1)

(e) Name the salt produced.

....

(1)

(f) Describe how an indicator can be used to show when all the sodium hydroxide has reacted with sulfuric acid.

....

....

....

....

....

....

(3)
(Total 9 marks)

Revision Pack Topic 4 Reactions

Q3. Explain, in terms of ions and molecules, what happens when any acid reacts with any alkali.

....

....

....

....

....

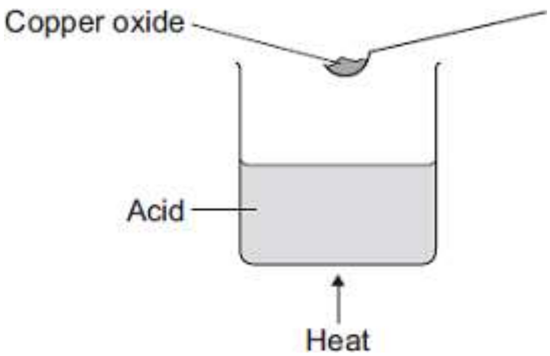
....

(Total 3 marks)

Q4. A student added copper oxide to an acid to make copper sulfate.

The student heated the acid. The student added copper oxide until no more reacted.

(a) The diagram shows the first stage in the experiment.



(i) Complete the word equation.

Copper oxide + ... acid → copper sulfate + water

(1)

(ii) Which **one** of these values could be the pH of the acid?

Draw a ring around the correct answer.

1 **7** **11**

(iii) Why is the acid heated?

....

....
(1)

(b) After the reaction is complete, some solid copper oxide remains.
Why?

....

....
(1)

(c) The student removed the solid copper oxide from the solution.

Suggest what the student should do to the solution to form copper sulfate crystals.

....

....
(1)

(d) The mass of copper sulfate crystals was less than the student expected.

Tick (✓) the **one** statement that explains why the mass of copper sulfate crystals was less than expected.

Statement	Tick (✓)
Some copper sulfate may have been lost during the experiment.	
The student added too much copper oxide.	
The copper sulfate crystals were wet when they were weighed.	

1)
(Total 6 marks)

Q5. A student investigates a potassium salt, **X**.

She finds that salt **X**:

- has a high melting point
- does not conduct electricity when it is solid
- dissolves in water and the solution does conduct electricity.

(a) What is the type of bonding in salt **X**?

Covalent ☐

Giant molecular ☐

Ionic ☐

Metallic ☐

(1)

(b) What is the name given to solutions that conduct electricity?

....
(1)

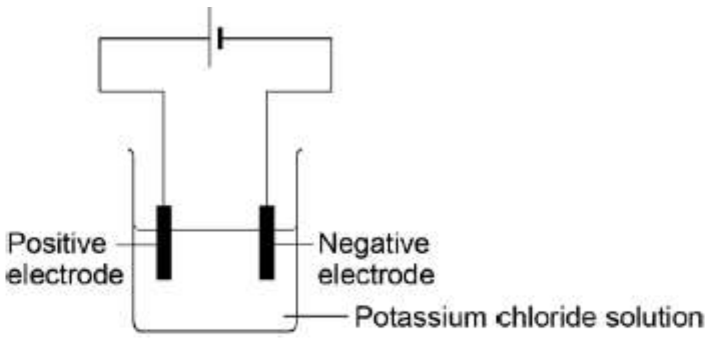
(c) Why does a solution of salt **X** in water conduct electricity?

....

....
(1)

(d) The student electrolyses a solution of potassium chloride.

Figure 1 shows the apparatus she uses.



When the current is switched on, bubbles of hydrogen gas are given off at the negative electrode.

Explain why hydrogen is produced and **not** potassium.

....

....

....

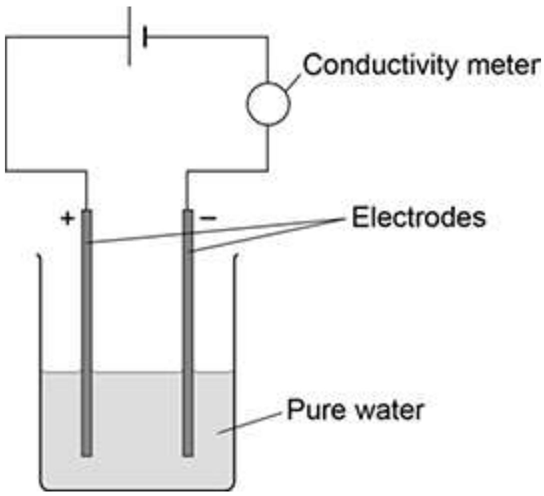
....

....

(2)

(e) The student then compares the relative conductivity of different concentrations of potassium chloride.

Figure 2 shows the apparatus she uses.



This is the method used.

1. Add potassium chloride solution to the water one drop at a time.
2. Stir the mixture.
3. Record the reading on the conductivity meter.

The table below shows the student's results.

Number of drops of potassium chloride solution	Relative conductivity of solution
0	0
1	90
2	180
3	270
4	360
5	450
6	540

When there is no potassium chloride in the beaker no electrical charge flows.

Suggest why pure water does **not** conduct electricity.

....

....

....

....

(2)

(f) Describe the relationship shown in the table above.

....

....

....

....

(2)
(Total 9 marks)

Revision Pack Topic 4 Reactions

M1.(a) any **three** from:

- concentration of (salt) solution
- volume of (salt) solution
    ignore amount of solution
- **initial** temperature (of the solution)
    ignore room temperature
- surface area / form of metal
- moles of metal
    allow mass / amount
    ignore time
    ignore size of tube

3

(b) 20

1

32

1

12

allow ecf

1

(c) (i) four bars of correct height

tolerance is + / - half square
*3 correct for **1** mark*

2

bars labelled

1

(ii) *one variable* is non-continuous / categoric

accept qualitative or discrete
accept no values between the metals

1

(iii) magnesium

1

because biggest temperature change

accept gives out most energy
ignore rate of reaction
dependent on first mark

1

(iv) does not react / silver cannot displace copper

1

because silver not more reactive (than copper) **or** silver below copper in reactivity series

*do **not** accept silver is less reactive than copper sulfate*

1

(v) replace the copper sulfate

could be implied

1

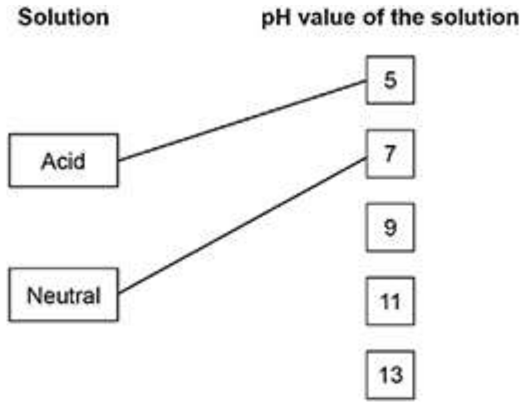
with any compound of a named metal less reactive than copper

allow students to score even if use an insoluble salt

1

[16]

2.(a)



extra lines from solution negate the mark

2

(b) H^+

1

(c) 3

1

(d) Neutralisation

1

(e) sodium sulfate

1

(f) Add indicator to sodium hydroxide solution

allow add indicator to sulfuric acid

1

Add sulfuric acid (gradually)

allow add sodium hydroxide solution (gradually)

1

allow pH probe

until indicator just changes (colour)

or until universal indicator turns green or shows pH7

1

[9]

M3. hydrogen ions (from acid) or protons / H^+

1

react with hydroxide ions (from alkali) / OH^-

1

to produce water

*$H^+ + OH^- \longrightarrow H_2O$ gains all **3** marks*
ignore state symbols
*molecules of hydrogen <u>ions</u> and molecules of hydroxide <u>ions</u> produce water = **2** marks*
*if they fail to get any of the above marks they can get **1** mark for neutralisation / product neutral*

1

[3]

M4.(a) (i) sulfuric

1

(ii) 1

1

(iii) to speed up the reaction

1

(b) because copper oxide in excess

allow copper oxide unreacted

or

because acid all used up / neutralised

1

(c) evaporation

allow heating
allow cooling
allow leave (to evaporate)
*do **not** accept freezing*

or

crystallisation

1

(d) Some copper sulfate may have been lost during the experiment

1

[6]

M5.(a) Ionic

1

(b) electrolyte

1

(c) because the ions are free to flow 1

(d) because potassium is higher in the reactivity series than hydrogen 1

so it is less easily discharged than hydrogen 1

(e) because water is covalent / molecular / contains molecules 1

so there are no free electrons to move **or** does not have an overall electrical charge 1

(f) conductivity of the solution increases with concentration 1

in a linear relationship **or** directly proportional 1

[9]