Year 9 Chemistry Workbook

The Polesworth School

Chemical Changes

Chemical Changes Mastery Booklet

This booklet covers:

1. Formulae of ionic compounds
2. Reactivity of metals
3. Extraction of metals
4. Acids
5. Ionic equations
6. Half equations
7. Electrolysis

Part 1: Formulae of ionic compounds

Recap questions:

1. Draw an atom of beryllium showing all its electrons
2. Draw an atom of oxygen showing all its electrons
3. In words, explain what occurs in terms of electrons when beryllium reacts with oxygen
4. Draw out the two ions formed in this reaction
5. Repeat the process, but for lithium reacting with oxygen
6. Use the above to explain why the formula for lithium oxide is Li_2O but the formula for beryllium oxide is BeO
7. Challenge: Predict the formulae of:
    a. Lithium fluoride
    b. Lithium nitride
    c. Aluminium nitride
    d. Aluminium chloride
    e. Aluminium oxide
    f. Magnesium nitride

There are many different ways to predict the formulae of ionic compounds. In terms of working out which charge the ion has:

- Group 1: 1+
- Group 2: 2+
- Group 3: 3+
- Group 5: 3-
- Group 6: 2-
- Group 7: 1-
- Transition metals have their charge given by a Roman numeral, e.g. copper (II) is Cu^{2+}
- Some ions can be compounds:

Name	Formula	Example
Ammonium ion	NH_4^+	Ammonium chloride, NH_4Cl
Hydroxide ion	OH^-	Calcium hydroxide. $Ca(OH)_2$
Sulphate ion	SO_4^{2-}	Sodium sulphate, Na_2SO_4
Carbonate ion	CO_3^{2-}	Calcium carbonate, $CaCO_3$
Nitrate ion	NO_3^-	Magnesium nitrate, $Mg(NO_3)_2$

Note that you use brackets to multiply the compound ion.

Practice –

8. Calcium chloride
9. Calcium sulphate
10. Lithium sulphate
11. Lithium nitrate
12. Aluminium bromide
13. Aluminium oxide
14. Aluminium nitrate
15. Aluminium sulphate
16. Ammonium iodide
17. Ammonium sulphate
18. Iron (II) sulphate
19. Iron (III) sulphate
20. Sodium nitrate
21. Barium sulphate (think about what group it is in!)
22. Copper (I) hydroxide
23. Copper (I) sulphate
24. Ammonium oxide
25. Ammonium carbonate
26. Aluminium hydroxide
27. Copper (II) carbonate
28. Beryllium fluoride
29. Beryllium sulphate
30. Vanadium (V) carbonate
31. Manganese (VII) iodide
32. Manganese (VII) sulphate

Part 2: Reactivity of Metals

Recap questions:

33. Describe the structure and bonding in metals. You may use a diagram to help.
34. Magnesium is a metal. Give three likely properties of magnesium.
35. Explain why magnesium conducts electricity.
36. Francium is a metal. It reacts with bromine (Br_2) to form francium bromide (FrBr). Write a word and symbol equation for this reaction.
37. 100g of Fr is used with 50g of bromine. Which is excess and which is limiting?
38. 75g of Fr is reacted with an excess of bromine. How much FrBr is formed?
39. FrBr is an ionic substance. Explain why it only conducts electricity when molten or dissolved in water.
40. Fr is more reactive than Li. Explain why by making reference to its electrons.
41. There are two isotopes of Francium, Fr-223 and Fr-221. Making reference to their numbers of protons, neutrons and electrons, discuss the similarities and differences between these two isotopes
42. A sample of Francium is found to have an abundance of 81% Fr-223 and 19% Fr-221. Calculate the average mass of the atoms in this sample.
43. Challenge: A similar sample is found to have a mass of 222.32. It contains 70% Fr-223, 22% Fr-221 and one other isotope. Calculate the mass of the other isotope.

Potassium
Sodium
Lithium
Calcium
Magnesium
(Carbon)
Zinc
Iron
(Hydrogen)
Copper
Gold

A student heats up a piece of magnesium in air. It reacts with oxygen to make magnesium oxide. The student then heats up a piece of copper, but no reaction takes place. This is because different metals can have different **reactivities**: how easy it is for them to react.

44. Which is more reactive, magnesium or copper? Explain your answer
45. Write a word equation for the reaction of magnesium with oxygen.
46. The formula for magnesium oxide is MgO. Write and balance a symbol equation for the reaction in Q2
47. Label all reactants and products in the equations above.

A reaction where oxygen is added is called an **oxidation** reaction. If oxygen is removed, it is called a **reduction** reaction.

48. Draw an atom of magnesium and an atom of oxygen.
49. Use arrows to show what occurs when they react with each other.
50. Draw the ions that are formed as a result of this reaction.
51. Explain why this reaction is called an oxidation reaction
52. In this reaction, the student started with 5g of magnesium. At the end of the reaction, they had 6.2g of magnesium oxide. Why did the mass increase?
53. Magnesium oxide can also be turned into magnesium and oxygen. Write an equation for this reaction.
54. What name is given to this type of reaction?
55. Explain why oxygen has a low melting point.
56. Explain why magnesium has a high melting point.
57. Under what conditions will magnesium oxide conduct electricity?
58. *Challenge: which is more reactive: magnesium or calcium? Explain your answer.*

The student takes the magnesium and puts it in a test tube with some acid in. The student notices that the acid starts to fizz quiet vigorously. The student takes the copper in and notices that it does not fizz at all. The student decides to try it again but with iron, and sees that it fizzes, but only a little bit.

59. Put the three metals tested in order of their reactivity
60. When iron atoms react, they form iron ions with a 3+ charge. What has happened to the iron atoms in terms of electrons?
61. Why is it important that the student uses the same acid for each reaction?
62. What else should the student try to keep the same?
63. In the iron reaction, iron sulphate is formed. Iron sulphate has a formula of $Fe_2(SO_4)_3$. How many atoms are present in iron sulphate and what atoms are they?
64. *Challenge: what is it about some metals that make them more reactive than other metals?*

Scientists use experiments like the one above to put metals into order by their reactivity. We call this the **reactivity series.** Most elements in the series are metals, but carbon and hydrogen are often included as well. In the series to the right, potassium is the most reactive and copper the least.

More reactive metals can displace less reactive ones, taking their place in a compound. For example, potassium reacts with sodium chloride to make potassium chloride and sodium. It has taken the place of the sodium in the chlorine.

Potassium + sodium chloride → sodium + potassium chloride

However, this would not occur in reverse:

Potassium chloride + sodium → no reaction

This is because sodium is *less reactive* than potassium so cannot displace it.

65. For the element pairs below, state which is more reactive and which is less reactive
    - a. Calcium and lithium
    - b. Gold and copper
    - c. Sodium and iron
    - d. Zinc and copper
    - e. Copper and zinc
    - f. Iron and zinc
    - g. Iron and calcium
    - h. Sodium and lithium
66. For each reaction below, state whether or not it would occur.
    - i. Magnesium oxide + calcium
    - j. Iron chloride + zinc
    - k. Copper bromide + gold
    - l. Zinc chloride + potassium
    - m. Iron sulphate + copper
    - n. Iron + lithium sulphate
    - o. Magnesium + iron oxide
67. Write a word equation for the reaction of potassium with sodium chloride
68. Explain why this is a displacement reaction
69. When zinc reacts with iron bromide it forms zinc bromide and iron. Write a word equation for this reaction
70. Identify any elements and compounds in the reaction
71. How does this prove that zinc is more reactive than iron?
72. When copper is added to lithium chloride no reaction takes place. Explain why this is the case.
73. The reaction between calcium chloride and lithium is shown below. Copy the equation into your book and balance it.
    $CaCl_2$ + Li → LiCl + Ca
74. In the reaction above, calcium chloride contains calcium ions. What charge do calcium ions have?
75. Use your periodic tables to calculate the number of protons, neutrons and electrons in calcium, chlorine and lithium.
76. Give the charges and relative masses of protons, neutrons and electrons.
77. A student wishes to know how much gas will be produced when metal X, metal Y and metal Z react with water. The student collects and measures the amount of gas produced in one minute. Why is it important to always use one minute?
78. The student notices that Y produces the most gas and X the least amount of gas. Write a reactivity series for the three metals.
79. Which metal would react easiest with oxygen?

80. Metal X has three electrons in its outer shell. What charge will it form when it becomes an alloy?

Part 3: Extracting metals

We need metals for all sorts of uses, from electrical wiring to construction. Most metals are found as **ores**. An ore is a rock that contains enough of a metal compound in it to be worth extracting. Some metals are so unreactive that they can be found as elements in the Earth's crust.

The way we extract the metal depends on its reactivity

Metals more reactive than carbon need electrolysis. Metals less reactive than carbon can be reduced by it.

Metal	**Extraction method**
Potassium	Electrolysis: using electricity to split up the compound
Sodium	
Lithium	
Calcium	
Magnesium	
(carbon)	
Zinc	Reduction with carbon: where carbon can remove the oxygen from a metal oxide
Iron	
Copper	
Gold	Mined from the Earth's crust

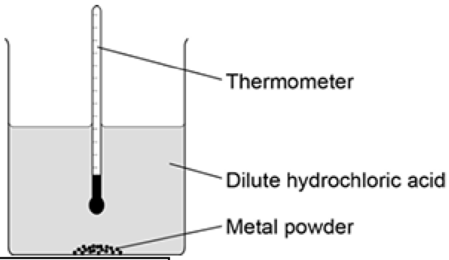
An example of a reduction with carbon is:

Iron oxide + carbon → iron + carbon dioxide

81. Why is the reaction above an example of a reduction?
82. A scientist finds a rock with a very small amount of aluminium oxide in it. Explain why this rock cannot be called an Ore
83. Francium is more reactive than lithium. How can francium be extracted from francium oxide?
84. Aluminium is more reactive than zinc, but less reactive than magnesium. What more information would you need before you could say how aluminium should be extracted from aluminium oxide?
85. Lithium. Iron and zinc are all placed in acid. Which one would react the most? Explain your answer.
86. Zinc can be produced from zinc oxide by reduction with carbon. Write a word equation for this reaction
87. Explain why gold is found naturally in the Earth's crust.
88. Silver is also found naturally in the Earth's crust, but is more reactive than gold. Where would it go in the reactivity series.
89. Lithium reacts with copper oxide to form lithium oxide and copper. Write a word equation for this reaction.
90. The symbol equation is:
    $Li + CuO \rightarrow Li_2O + Cu$
    Copy the equation into your book and balance it.
91. Explain why lithium has been oxidised and copper oxide has been reduced.
92. Iron oxide reacts with sodium to form sodium oxide and iron.
93. Write a word equation for this reaction.
94. The symbol equation for this reaction is:
    $Fe_2O_3 + Na \rightarrow Na_2O + Fe$
    Copy the equation into your book and balance it.
95. Identify what has been oxidised and what has been reduced.
96. Potassium and zinc oxide are reacted together. Suggest the products of this reaction.
97. What has been oxidised and what has been reduced? Explain your answer.

GCSE Practice Questions

Question 1: A student investigated the reactivity of different metals. The student used the apparatus shown in the figure to the right.



The student used four different metals. The student measured the temperature rise for each metal three times. The student's results are shown in the table below.

Metal	Temperature rise in °C			Mean temperature rise in °C
	Test 1	Test 2	Test 3	
Calcium	17.8	16.9	17.5	
Iron	6.2	6.0	6.1	6.1
Magnesium	12.5	4.2	12.3	12.4
Zinc	7.8	8.0	7.6	7.8

(a) Give **two** variables the student should control so that the investigation is a fair test.
(b) One of the results for magnesium is anomalous. Which result is anomalous? Suggest **one** reason why this anomalous result was obtained.
(c) Calculate the mean temperature rise for calcium.
(e) Aluminium is more reactive than iron and zinc but less reactive than calcium and magnesium. Predict the temperature rise when aluminium is reacted with dilute hydrochloric acid.

Question 2: A student investigated the reactivity of three different metals. This is the method used.

1. Place 1 g of metal powder in a test tube.
2. Add 10 cm^3 of metal sulfate.
3. Wait 1 minute and observe.
4. Repeat using the other metals and metal sulfates.

The student placed a tick in the table below if there was a reaction and a cross if there was no reaction.

	Zinc	**Copper**	**Magnesium**
Copper sulfate	✓	x	✓
Magnesium sulfate	x	x	x
Zinc sulfate	x	x	✓

(a) What is the dependent variable in the investigation?
(c) The student used measuring instruments to measure some of the variables. Draw **one** line from each variable to the measuring instrument used to measure the variable.

Variable	**Measuring instrument**
	Balance
	Measuring cylinder
Mass of metal powder	
	Ruler
	Burette
Volume of metal sulfate	
	Thermometer
	Test tube

(d) Use the results shown in table above to place zinc, copper and magnesium in order of reactivity.
(e) Suggest **one** reason why the student should **not** use sodium in this investigation.
(f) Out of calcium, gold, lithium and potassium, which metal is found in the Earth as the metal itself?
(g) Iron is found in the Earth as iron oxide (Fe_2O_3). Iron oxide is reduced to produce iron. Balance the equation for the reaction.

$$\_\_\_Fe_2O_3 \quad + \quad \_\_\_C \quad \rightarrow \quad \_\_\_Fe \quad + \quad \_\_\_CO_2$$

(h) Name the element used to reduce iron oxide.
(i) What is meant by reduction?

Question 3: Metals are used in the manufacture of pylons and overhead power cables.
Suggest **one** reason why iron (steel) is used to make pylons.
(b) The table shows some of the properties of two metals.

Metal	**Density in g per cm^3**	**Melting point in °C**	**Percentage(%) relative electrical conductivity**	**Percentage(%) abundance in Earth's crust**
copper	8.92	1083	100	0.007
aluminium	2.70	660	60	8.1

Use the information in the table to suggest why aluminium and **not** copper is used to conduct electricity in overhead power cables.

Part 4: Acids

Acids are substances that have a pH of less than 7. Alkalis have a pH of more than 7. A pH of 7 exactly is called a **neutral solution**.

The three main acids are:

- Hydrochloric acid: $HCl(aq)$
- Nitric acid: $HNO_3(aq)$
- Sulphuric acid: $H_2SO_4(aq)$

Each of these splits up into ions when dissolved in water:

- HCl splits into H^+ and Cl^-
- HNO_3 splits into H^+ and NO_3^-
- H_2SO_4 splits into $2H^+$ and SO_4^{2-}

It is the H^+ which makes them acidic, and the pH scale is based on how much H^+ there is present. The more H^+, the lower the pH.

1. Reactions of acids with metals

When a metal is added to acid, it will react to form a **salt** and hydrogen gas. A salt is a compound where the hydrogen from an acid has been replaced with a metal. For example:

Calcium + sulphuric acid → calcium suphate + hydrogen

$Ca(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2(g)$

Here, the hydrogens from the sulphuric acid have been replaced by calcium.

- Hydrochloric acid makes metal chlorides
- Nitric acid makes metal nitrates
- Sulphuric acid makes metal sulphates

98. Complete the table below. The first one has been done for you

Formula of salt	Name of salt	Original metal	Original acid
$NaCl$	Sodium chloride	Sodium	Hydrochloric acid
$LiCl$			
$CaSO_4$			
$MgSO_4$			
KNO_3			

99. Write a word equation for the reactions between:
    a. Aluminium and hydrochloric acid
    b. Magnesium and hydrochloric acid
    c. Calcium and nitric acid
    d. Beryllium and sulphuric acid
    e. Iron and sulphuric acid
100. Write a word equation for a reaction which forms:
    f. Aluminium sulphate
    g. Calcium chloride
    h. Barium nitrate
    i. Rubidium chloride

Challenge: write symbol equations for every reaction in Q1 and Q2

Part 5: Ionic equations

We will return to acids shortly. Earlier, we looked at how different metals formed different ions. If they are in group 1, they form 1+, if they are in group 2 they form 2+ and in group 3 they form 3+. If they are in the central block, there will be a roman numeral to tell you the charge. For example, Iron (III) is the Fe^{3+} ion.

101. Complete the table below. Some have been done for you.

Atom	Ion	**Atom**	Ion	**Atom**	Ion
Li	Li^{+}	**Be**			Al^{3+}
Na		**Mg**	Mg^{2+}		Fe^{2+}
K			Ca	**Iron (III)**	Fe^{3+}

Remember that the ions from acids are: Cl^{-}, NO_3^{-} and SO_4^{2-}. When the metal ion combines with the ion from an acid to make a salt, the charges must balance. For example:

When magnesium reacts with hydrochloric acid, it forms magnesium chloride.

Magnesium forms Mg^{2+}
Hydrochloric acid forms Cl^{-}
I need two chlorine ions for each magnesium ion so the formula for magnesium chloride is $MgCl_2$.

Worked example 1

What is the formula for magnesium nitrate?

Magnesium forms Mg^{2+}
Nitric acid forms NO_3^{-}
I need two nitrate ions for every magnesium ion so the formula for magnesium nitrate is $Mg(NO_3)_2$.

Worked example 2

What is the formula for lithium sulphate?

Lithium forms Li^{+}
Sulphuric acid forms SO_4^{2-}
I need two lithium ions for every sulphate ion so the formula for lithium sulphate is Li_2SO_4

102. For each of the salts formed in question 2 and question 3 write out the symbol formula.
103. Complete the equations:
    j. $Mg + H_2SO_4 \rightarrow$ _____________ $+ H_2$
    k. $2Li + 2HNO_3 \rightarrow 2$_____________ $+ H_2$
    l. $2Al + 6HNO_3 \rightarrow 2$ _____________ $+ 3H_2$
    m. $2Fe(III) + 6HCl \rightarrow 2$_____________ $+ 3H_2$
104. For the equations below, you will need to balance and complete:
    n. $Mg + HNO_3 \rightarrow$ _______________ $+ H_2$
    o. $Fe(III) + HNO_3 \rightarrow$ _______________ $+ H_2$
    p. $K + H_2SO_4 \rightarrow$ _______________ $+ H_2$
    q. $Li + H_2SO_4 \rightarrow$ _______________ $+ H_2$
105. Write out full word and symbol equations for each of the below:
    r. A reaction between iron (II) and nitric acid
    s. A reaction between barium and sulphuric acid
    t. A reaction between strontium and hydrochloric acid
    u. A reaction between aluminium and sulphuric acid
    v. A reaction between gallium and nitric acid

Part 5: Ionic equations

In each of the reactions above, ions are involved and change charges. We can use an ionic equation to illustrate this.

Worked example 3:

For the reaction between magnesium and hydrochloric acid construct a balanced equation and an ionic equation.

We can use the above to write a word equation:

Magnesium + hydrochloric acid → magnesium chloride + hydrogen

$Mg(s) + HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

First the equation must be balanced:

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

Next, we need to split the ionic substances into ions. Remember that in order to have free ions it must:

1. Have an (aq) or (l) symbol
2. Be ionic, so either:
    a. An acid
    b. Or an ionic compound (has a metal and a non-metal)

In our case, HCl and $MgCl_2$ are the only ionic substances which can split up.

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

We can remove the spectator ions (the ions which are unchanged by the reaction):

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

$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

Worked example 4

For the reaction between magnesium and nitric acid construct a balanced equation and an ionic equation. Below is a summarised version of what we did above.

Word equation	Magnesium + nitric acid → magnesium nitrate + hydrogen
Symbol equation	$Mg(s) + HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$
Balanced symbol equation	$Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$
Ionic equation	$Mg(s) + 2H^+(aq) + 2NO_3^-(aq) \rightarrow Mg^{2+}(aq) + 2NO_3^{1-}(aq) + H_2(g)$
Ionic equation without spectator ions	$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

106. For each of the reactions in question 103→105 construct an ionic equation

Part 6: Half equations

In each of the reactions above, the metal loses electrons and the H^+ ions from the acid gain them. We can see this from constructing half equations from ionic equations. This means that the metal is oxidised and the hydrogen ions are reduced.

Worked example 5:

Using the half equation from worked example 3, construct a pair of half equations to show which substance is losing electrons and which is gaining.

$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

In this case, magnesium has gone from no charge to a charge of 2+. This means it must have lost two electrons:

$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

This reaction is oxidation

The hydrogen ions have gained two electrons:

$2H^+(aq) + 2e^- \rightarrow H_2(g)$

This reaction is reduction

107. For each of the ionic equations in question 106 construct half equations.

Acids continued: 5. Metal Hydroxides

An alkali is a metal hydroxide which is soluble (dissolves in water). The hydroxide ion is OH^-.

Worked example 6:

State the formula for sodium hydroxide and for calcium hydroxide:

Sodium forms Na^+ ions
The hydroxide ion is OH^-
These two balance out sot the formula is NaOH

Calcium forms Ca^{2+} ions
The hydroxide ion is OH^-
Two hydroxide ions are needed for each calcium ion so the formula is $Ca(OH)_2$

108. State the formulae for the hydroxides of: lithium, potassium, calcium, aluminium, magnesium, iron (II), iron (III), zinc (II) and beryllium

6. Alkali + acid → salt + water

Whenever an acid reacts with an alkali, a salt and water are formed. This is called a **neutralisation reaction** because you start with an acid but end up with a neutral solution: a solution with a pH of 7. For example, when sodium hydroxide reacts with hydrochloric acid, the reaction below takes place:

Sodium hydroxide + hydrochloric acid → sodium chloride + water

The symbol equation for this is:
$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

The ionic equation is:
$Na^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$

Once you remove the spectator ions you are left with:
$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

Which is the ionic equation for all acid + alkali reactions.

For each of the reactions below, write:

a. A word equation
b. A balanced symbol equation (assume that all hydroxides are dissolved in water)
c. An ionic equation

109. Potassium hydroxide and hydrochloric acid
110. Magnesium hydroxide and hydrochloric acid
111. Sodium hydroxide and nitric acid
112. Sodium hydroxide and sulphuric acid
113. Iron (III) hydroxide and sulphuric acid

7. Acid and base

Bases are any metal hydroxide or oxide which is insoluble in water. Remember that the oxide ion is O^{2-}.

114. Write the formula for the oxides of lithium, potassium, calcium, aluminium, magnesium, iron (II), iron (III), zinc (II) and beryllium

8. Metal oxide + acid → salt + water

The products are exactly the same as for acid + alkali but because the metal oxides are insoluble we do not need to write ionic equations for them.

For each of the reactions below, write a word equation and balanced symbol equation:

115. Potassium oxide and hydrochloric acid

116. Magnesium oxide and hydrochloric acid
117. Sodium oxide and nitric acid
118. Sodium oxide and sulphuric acid
119. Iron (III) oxide and sulphuric acid

9. Metal carbonate + acid → salt + water + carbon dioxide

When a metal carbonate reacts with an acid it forms a salt, water and carbon dioxide. The carbonate ion is CO_3^{2-}. Because carbonates are also insoluble, we do not need to write ionic equations.

State the formula for calcium carbonate and for sodium carbonate:

Calcium forms Ca^{2+} ions
The carbonate ion is CO_3^{2-}
These are balanced out so the formula is $CaCO_3$

Sodium forms Na^{+} ions
The carbonate ion is CO_3^{2-}
Two sodium ions are needed for each carbonate ions so the formula is Na_2CO_3

120. State the formulae for the carbonates of: lithium, potassium, calcium, aluminium, magnesium, iron (II), iron (III), zinc (II) and beryllium

Example of writing the equations for a metal carbonate + acid:

Word equation:
Lithium carbonate + hydrochloric acid → Lithium chloride + water + carbon dioxide

Symbol equation:
$Li_2CO_3(s) + 2HCl(aq) \rightarrow 2LiCl(aq) + H_2O(l) + CO_2(g)$

For each of the reactions below, write a word equation and balanced symbol equation:

121. Potassium carbonate and hydrochloric acid
122. Magnesium carbonate and hydrochloric acid
123. Sodium carbonate and nitric acid
124. Sodium carbonate and sulphuric acid
125. Iron (III) carbonate and sulphuric acid

pH

The pH of a solution tells us how acidic or alkaline it is. Low pH means it is very acidic. pH measures how many H^+ ions there are in solution so more H^+ means more acidic and a lower pH. The pH scale does not work like normal quantities. If we take length for example, a 2m plank is twice the length of a 1m plank, and half the length of a 4m plank. The distance between 1m and 2m is the same as between 2m and 3m.

With pH, the distances are not the same, but increase by 10. If something has a pH of 1, it is not **twice** as acidic as something with a pH of 2, but **ten times** as acidic.

Example: a solution has a pH of 1, and 1000g of acid in it. What mass of acid would be needed for a pH 2?

We know that there will need less acid for a higher pH, and that pH 1 is ten times more acidic than pH 2. Therefore, 1000/10 = 100, so 100g would need to be added.

Example 2: a solution has a pH of 5, and 70g of acid in it. What mass of acid would be required for pH 2?

We know that to go from pH 5 to pH 4 we need to multiply by 10, so we would have 7g acid. To get from pH 4 to 3 multiply by 10 again to 70g, then from pH 3 to pH 2 by 10 again to 700g.

Questions:

126. A substance has 40g of acid in it and is pH 2.
    a. What mass of acid would be required for pH 1?
    b. What mass of acid would be required for pH 3?
    c. What mass of acid would be required for pH 4?
127. A certain solution has 12050 H^+ ions in it and has a pH of 3.
    d. How many H^+ ions would need to be added for a pH of 1?
    e. How many H^+ ions would need to be taken away for a pH of 5?

Strong, weak, dilute, concentrated

All acids have molecules ionise (turn into ions) when they are in water, and all of them release H^+. However, for some acids, not **all** their molecules ionise when in water. These are called **weak** acids, as they do not release as much H^+ as they could. In **strong** acids **all** the molecules ionise when in water, releasing a lot more H^+.

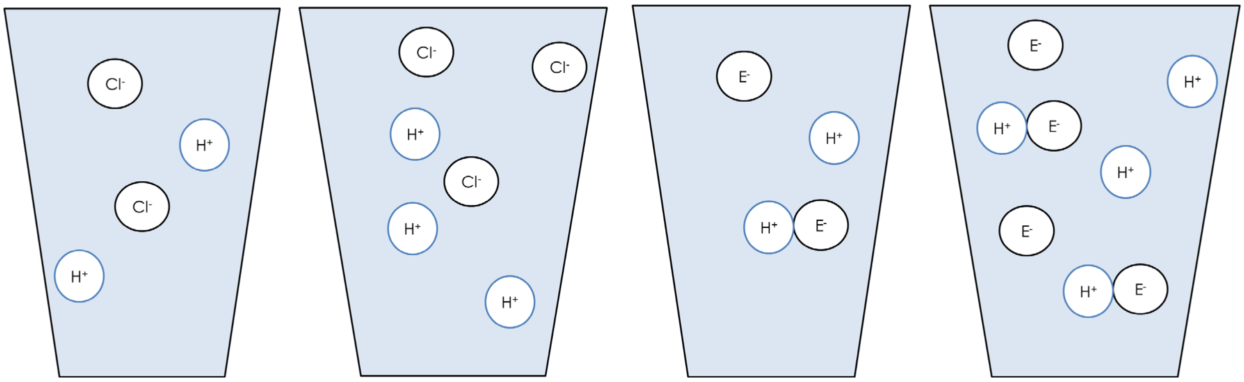
Concentrated or dilute relates simply to how much acid there is in the water. If I add only a little acid to the water, it is a dilute acid. If I add a lot of acid, it is concentrated.

Dilute Strong Acid – there are a few acid molecules but they are all split up

Concentrated Strong Acid – there are lots of acid molecules and they are all split up

Dilute Weak Acid: there are few molecules, and only a few have split up

Concentrated Weak Acid: there are a lot of molecules, but only a few have split up



128. Sulphuric acid is a strong acid. Explain what this means
129. Which ions does sulphuric acid split up into?
130. 2g of sulphuric acid is dissolved in water, and the pH is found to be 2. What mass of acid would be required for a pH of 1?
131. Complete the table below, using the words “high,” “medium” or “low” to represent pH. The first one has been done for you.

	Concentrated	Dilute
Strong acid	*Low*	
Weak acid		

132. Write a word equation for the reaction between sulphuric acid and sodium hydroxide
133. Write a symbol equation for this reaction
134. Write an ionic equation for this reaction
135. 10g of acid is dissolved in water, followed by another 10g. What happens to the pH?
136. In terms of the number of H^+ ions in solution, explain why a strong acid has a lower pH than a weak acid.

Back to metal extraction

We saw that a displacement reaction is one where a more reactive element takes the place of a less reactive one in a compound, for example:

$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

We can write an ionic equation for this:

$Fe(s) + Cu^{2+}(aq) + SO_4^{21}(aq) \rightarrow Fe^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$

Cancel the spectator ions:

$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

We can then write half equations for this too:

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

This shows that iron has been oxidised and copper has been reduced.

137. In a reaction, solid sodium displaces iron from iron (II) sulphate solution:
$Na + FeSO_4 \rightarrow Na_2SO_4 + Fe$
    a. Balance the equation
    b. Add state symbols to the equation
    c. Write an ionic equation for this reaction
    d. Use half equations to show that sodium has been oxidised and copper ions have been reduced
138. Repeat the above process for reactions between:
    e. $Na(s) + LiOH(aq) \rightarrow NaOH(aq) + Li(s)$
    f. $Ca(s) + MgCl_2(aq) \rightarrow CaCl_2(aq) + Mg(s)$
    g. $Ca(s) + Fe_2SO_4(aq) \rightarrow CaSO_4(aq) + Fe(s)$ (the charge on the iron ion in iron sulphate is 3+)
139. Sodium is more reactive than lithium. Explain why.

We can also do the same process with the displacement reactions of the halogens. In a reaction between lithium bromide and chlorine, the chlorine displaces the bromine because it is more reactive:

$LiBr(aq) + Cl_2(aq) \rightarrow LiCl(aq) + Br_2$

First, we balance the equation:

$2LiBr(aq) + Cl_2(aq) \rightarrow 2LiCl(aq) + Br_2$

Then, we write an ionic equation (Cl_2 is covalent so does not split into ions)

$2Li^+(aq) + 2Br^-(aq) + Cl_2(aq) \rightarrow 2Li^+(aq) + 2Cl^-(aq) + Br_2(aq)$

Cancel spectator ions:

$2Br^-(aq) + Cl_2(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$

And we can also now write half equations:

$Cl_2(aq) + 2e^- \rightarrow 2Cl^-(aq)$

$2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$

Showing that chlorine has been reduced and bromide ions have been oxidised

140. Repeat the process above for the reaction:
$LiCl(aq) + F_2(aq) \rightarrow LiF(aq) + Cl_2$
141. Repeat the process above for the reaction:
$MgI_2(aq) + Br_2(aq) \rightarrow MgBr_2(aq) + I_2(aq)$
142. Explain why there is no reaction between sodium chloride and bromine
143. In the reaction above between lithium bromide and chlorine, 10g of lithium bromide is used.
    a. How much chlorine will be needed for a complete reaction?
    b. What mass of lithium chloride will be produced?
144. Lithium chloride solution conducts electricity. Explain why.
145. A solution of lithium chloride is made with 20g of lithium chloride in 100cm^3 of water. Calculate its concentration in g/dm^3 (triple only – calculate also in mol/dm^3).
146. Explain why fluorine is more reactive than iodine.

Summary problem

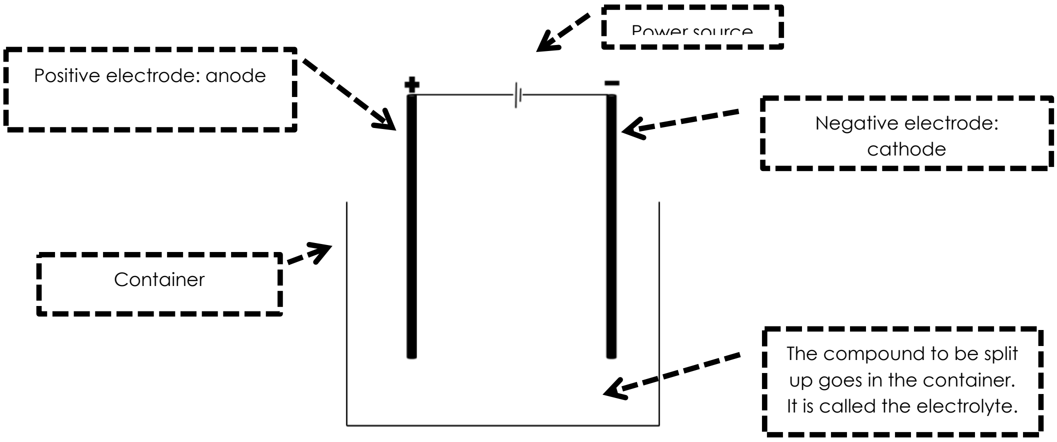
This question is about the extraction of copper, an important metal in a number of settings.

147. Copper is a metal. Describe the structure and bonding of copper.
148. Copper can be alloyed with other elements to make it harder. Explain why an alloy of copper is harder than pure copper
149. Copper conducts electricity. Explain why.
150. Copper is a very good thermal conductor. Explain why.
151. Copper exists in a number of isotopes. What is an isotope?
152. Copper can react with oxygen to form copper oxide (CuO)
    a. Write a word equation for this reaction
    b. Write and balance a symbol equation for this reaction
    c. Identify all elements, compounds, reactants and products in your equation
    d. If 50g of copper is reacted in an excess of oxygen, what mass of copper oxide will be formed?
    e. What is meant by an "excess" of oxygen?

153. Copper be extracted from copper oxide by reacting it with carbon. This is called **reduction**.
   f. What is meant by reduction?
   g. Write a word equation for this reaction
   h. Construct and balance a symbol equation for this reaction
   i. What mass of carbon will be needed to react with 75g of copper?
154. Copper can react with some acids to form salts.
   j. Copper cannot react with some acids because they are weak acids. What is meant by weak acid?
   k. Copper reacts well with concentrated hydrochloric acid. What is meant by concentrated acid?
   l. Write a word equation for the reaction between copper and hydrochloric acid which forms copper (II) chloride and hydrogen
   m. Explain why the mass of this reaction appears to decrease over time
   n. Write a symbol equation for this reaction (including state symbols)
   o. Construct an ionic equation for this reaction
   p. Construct half equations for this reaction
   q. Identify which substance has been oxidised and which has been reduced

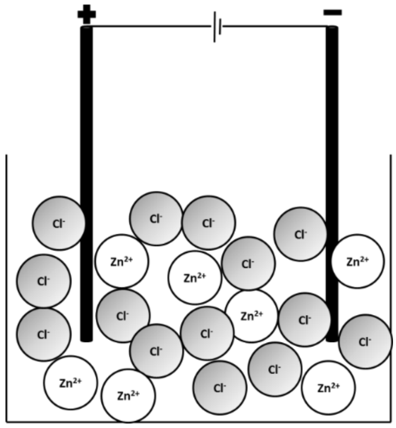
Part 7: Electrolysis

Electrolysis is the process of using electricity to split apart ionic compounds. The compounds can either be molten (liquid) or in a solution. The basic apparatus for electrolysis is always very similar:



A useful mnemonic to remember the names of the electrodes is PANIC: positive anode, negative is cathode.

155. A student is to electrolyse sodium chloride. If the sodium chloride is a solid, what must be done to it before it can be electrolysed?
156. What is the electrolyte in this case?
157. List the apparatus required to electrolyse sodium chloride



Electrolysis of liquids

Ionic compounds need to be molten or in solution for electrolysis to work. This is because the charged particles that make them up (ions) need to be free to move to the electrodes.

The positive ion (always a metal) will travel to the cathode, where it will gain electrons to become an element,

The negative ion (the non-metal) will travel to the anode, where it will lose electrons to become an element.

The example below involves electrolysis of molten zinc chloride, $ZnCl_2$(l). When zinc chloride is melted, the ions which make it up become free to move. The Zn^{2+}(l) will travel to the cathode and the Cl^-(l) will travel to the anode.

Pure zinc metal will be produced at the cathode and chlorine will be produced at the anode.

158. In the electrolysis of zinc chloride, what is the electrolyte?
159. In the electrolysis of each of the molten compounds below, state which elements will be produced:
   a. Zinc iodide
   b. Lithium bromide
   c. Iron fluoride
   d. Sodium oxide
   e. Potassium chloride
160. For each of the compounds in question 5, state at which electrode each element will be produced.
161. Deduce the formulae of each of the compounds in question 5

162. Why can electrolysis not be performed on covalent substances?
163. Why can electrolysis not be performed on metals?

Redox and half equations recap

In previous units, we have learnt that:

- Oxidation is Loss of Electrons
- Reduction is Gain of Electrons

In the electrolysis of zinc chloride, $Zn^{2+}(l)$ gains two electrons to form Zn(s). This can be represented by a half equation:

$Zn^{2+}(l) + 2e^- \rightarrow Zn(s)$

This occurs at the cathode. At the anode, $Cl^-(l)$ turns into $Cl_2(g)$. This can also be represented by a half equation:

$Cl^-(l) \rightarrow Cl_2(g) + e^-$

This is not balanced as we need two Cl^- ions in order to form Cl_2. Each of those ions will lose one electron so two electrons are lost overall:

$2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

Example 2: electrolysis of NaCl(l)

NaCl(l) will split up into $Na^+(l)$ and $Cl^-(l)$. $Na^+(l)$ will travel to the cathode where it will be reduced to form Na(s). $Cl^-(l)$ will travel to the anode where it will be oxidised to form $Cl_2(g)$.

$Na^+(l) + e^- \rightarrow Na(s)$
$2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

Ideally, there should be the same number of electrons in each half equation so we multiply the first equation by 2:

$2Na^+(l) + 2e^- \rightarrow 2Na(s)$
$2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

164. Copy the table below into your exercise book. Add 7 empty rows.

Formula	**Positive ion**	**Negative ion**	**Element formed at cathode**	**Element formed at anode**	**Half equation at cathode**	**Half equation at anode**
$ZnCl_2$	$Zn^{2+}(l)$	$Cl^-(l)$	Zn(s)	$Cl_2(g)$	$Zn^{2+}(l) + 2e^- \rightarrow Zn(s)$	$2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

165. Complete the table for the electrolysis of the compounds:
    f. NaCl
    g. NaBr
    h. KI
    i. $CaCl_2$
    j. $AlBr_3$
    k. Na_2O (to form oxygen gas)
    l. Al_2O_3
166. Define an ore
167. Why is gold found naturally in the Earth's crust?
168. Why is electrolysis not necessary to extract iron from iron oxide?
169. What properties would you expect iron oxide to have?
170. When iron (III) oxide is reacted with carbon, iron and carbon dioxide are produced. Write a word equation for this reaction
171. A student has a sample of calcium chloride from which they want to extract pure calcium. Why is electrolysis of calcium oxide required to extract pure metal calcium?
172. *Challenge: write a balanced symbol equation for the reaction in 16. Remember that you will have to deduce the formula of iron oxide (it isn't FeO!)*
173. *Challenge: deduce the formula of vanadium (V) oxide and give half equations for its electrolysis*

Electrolysis of aluminium oxide

Because aluminium is more reactive than carbon, it must be extracted using electrolysis. Electrolysis requires a lot of energy so scientists have to find ways to minimise the energy use

Aluminium oxide's melting point is very high so we mix it with a substance called cryolite which brings down the melting point.

At the cathode, $Al^{3+}(l)$ is reduced to Al(s): $Al^{3+}(l) + 3e^- \rightarrow Al(l)$

Al is formed as a liquid as the temperatures used are so hot.

At the anode, $O^{2-}(l)$ is oxidised to $O_2(g)$: $2O^{2-}(l) \rightarrow O_2(g) + 4e^-$

The electrodes are made of carbon. When the oxygen gas is produced, it reacts with the carbon to make carbon dioxide:

$C(s) + O_2(g) \rightarrow CO_2(g)$

This means that gradually the anode wears away over time and needs to be replaced. This is another cost to consider in the production of aluminium.

174. Explain why the formula of aluminium oxide is Al_2O_3 and not AlO_3
175. Balance the equation:
 $Al + O_2 \rightarrow Al_2O_3$
176. In terms of electrons, explain how aluminium reacts with oxygen to form aluminium oxide.
177. In terms of its bonding and structure, explain why aluminium oxide has a high melting point.
178. Explain why aluminium oxide needs to be molten before it can be electrolysed.
179. In the electrolysis of aluminium oxide, the electrodes are made of graphite. Explain how graphite can conduct electricity.
180. Why must the anodes be regularly replaced?
181. Sometimes, the anodes react with oxygen to form carbon monoxide (CO). Write a balanced symbol equation for this reaction.
182. Is electrolysis an exothermic or endothermic change?
183. Explain your answer.

Electrolysis of solutions

If an ionic compound is soluble, then we can electrolyse its solution as the ions become free to travel through the water. If we take NaCl as an example, when it is dissolved in water we obtain $Na^+(aq)$ and $Cl^-(aq)$.

We would expect to therefore obtain pure Na at the cathode. However, the element produced actually depends on the reactivity of the elements involved.

- If the metal is more reactive than hydrogen, hydrogen gas will be produced at the cathode
- Unless the non-metal is a halogen, oxygen gas will be produced

This is because water can also be electrolysed, breaking down to form hydrogen and oxygen. The process by which water is electrolysed works as follows:

First, water breaks apart into ions: $H_2O(l) \rightarrow H^+(aq) + OH^-(aq)$

The hydrogen ions are attracted to the cathode where they gain electrons and form hydrogen gas:

$2H^+(aq) + 2e^- \rightarrow H_2(g)$

The OH^- ions are attracted to the anode where they lose electrons and form oxygen:

$4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$

184. For each of the below, state which elements are formed at the anode and at the cathode
 m. Copper sulphate
 n. Silver nitrate
 o. Tin chloride
 p. Zinc fluoride
 q. Zinc sulphate
 r. Calcium nitrate
 s. Potassium chloride
185. Explain why potassium can only be extracted from potassium nitrate if it is molten, not if it is dissolved.
186. Sodium chloride is dissolved in water.
 t. Which ions are present when the compound is dissolved? (hint – there are 2)
 u. As soon as the electrolysis starts, which ions are present? (hint – there are 4)
 v. Which element will be formed at the anode?
 w. Which element will be formed at the cathode?
 x. Give a half equation for the reaction at the anode and at the cathode
 y. The process is repeated but with copper (II) sulphate. Give half equations for the reactions at the anode and the cathode

GCSE Questions

Q1. This question is about halogens and their compounds.

The table below shows the boiling points and properties of some of the elements in Group 7 of the periodic table.

Element	**Boiling point in °C**	**Colour in aqueous solution**
Fluorine	−188	colourless
Chlorine	−35	pale green
Bromine	X	orange
Iodine	184	brown

(a) Why does iodine have a higher boiling point than chlorine?

(b) Predict the boiling point of bromine.

(d) A redox reaction takes place when aqueous chlorine is added to potassium iodide solution. What is the ionic equation for the reaction of chlorine with potassium iodide?

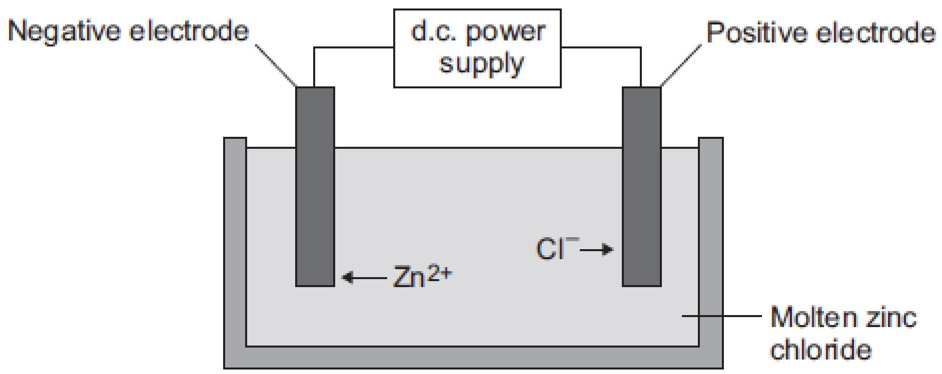
(e) Why does potassium iodide solution conduct electricity?

(f) What are the products of electrolysing potassium iodide solution?

Q2. This question is about zinc.

Figure 1 shows the electrolysis of molten zinc chloride.

Figure 1



(a) Zinc chloride is an ionic substance.

Complete the sentence.

When zinc chloride is molten, it will conduct ....**(1)**

(b) Zinc ions move towards the negative electrode where they gain electrons to produce zinc.

(i) Name the product formed at the positive electrode. .... **(1)**

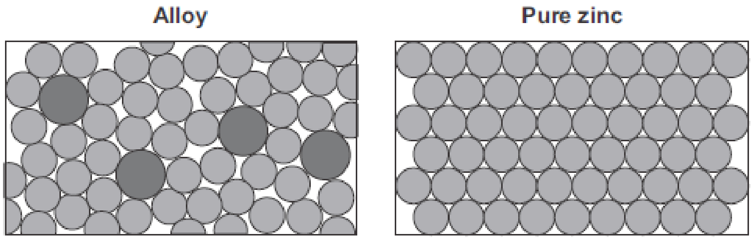
(ii) Explain why zinc ions move towards the negative electrode. **(2)**

(iii) What type of reaction occurs when the zinc ions gain electrons? **(1)**

(c) Zinc is mixed with copper to make an alloy.

(i) **Figure 2** shows the particles in the alloy and in pure zinc.

Figure 2



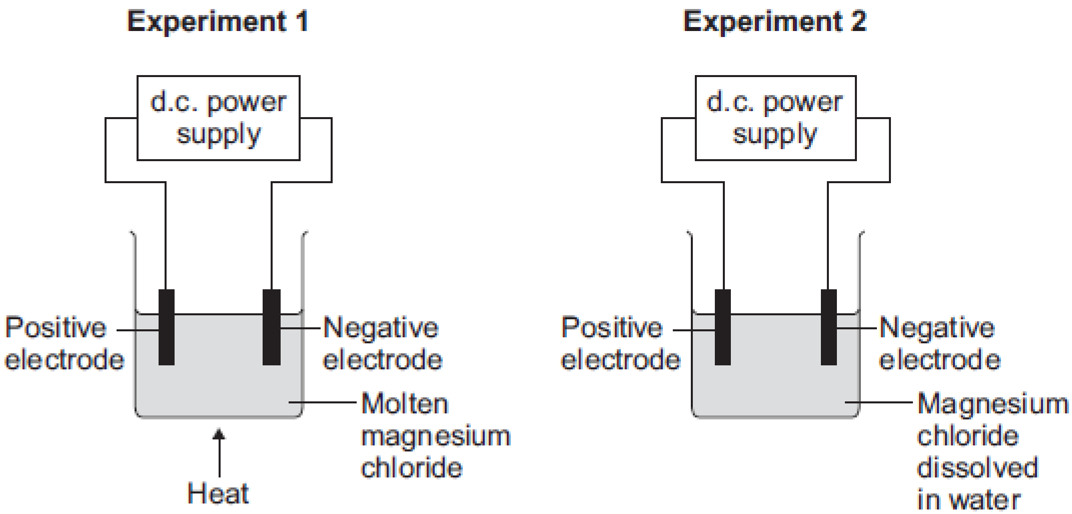
Use **Figure 2** to explain why the alloy is harder than pure zinc. **(2)**

Q3. This question is about magnesium and magnesium chloride.

(a) Magnesium chloride contains magnesium ions (Mg^{2+}) and chloride ions (Cl^-). Describe, in terms of electrons, what happens when a magnesium atom reacts with chlorine atoms to produce magnesium chloride. **(4)**

(b) Magnesium chloride can be electrolysed.

The diagram below shows two experiments for electrolysing magnesium chloride.



(i) Explain why magnesium chloride must be molten or dissolved in water to be electrolysed. **(2)**

(ii) Explain how magnesium is produced at the negative electrode in **Experiment 1**. **(3)**

(iii) In **Experiment 2** a gas is produced at the negative electrode. Name the gas produced at the negative electrode. **(1)**

(iv) Suggest why magnesium is **not** produced at the negative electrode in **Experiment 2**. **(1)**

(v) Complete and balance the half equation for the reaction at the positive electrode.

.... Cl^- $\rightarrow$ Cl_2 + ... **(1)**

(c) Magnesium is a metal. Explain why metals can be bent and shaped. **(2)**