# Lesson 10 C4.10 Required practical: Finding the reacting volumes of solutions of acid and alkali (Triple only)

## Connection

Q1. State what colour pure water would turn when universal indicator is added

Q2. Identify the ion in Hydrochloric acid that makes it acidic

Q3. Write an ionic equation for the reaction between sodium hydroxide and hydrochloric acid

## Activation

**LI: Make and record observations and accurate measurements using burettes**

1. https://www.youtube.com/watch?v=saRBT5oZfh8
2. Make a note of the title and the LI
3. Read pages 150-151
4. Give definitions for the key words on page 150
5. Draw fig 4.37 and fig 4.38 page 150-151
6. Carry out titration and plot titration curve

## Demonstration

Attempt questions 1-12
In 15 mins answer as many questions as you can.
Self mark the questions you have done making any necessary corrections in blue pen

Challenge yourself to answer as many as you can:
Green questions to GCSE Level 3
Blue questions to GCSE Level 6
Purple questions to GCSE Level 9

## Consolidation

Complete and self assess the relevant past paper question for this topic - From the C4 DIP file

## Extension

Make a note of one thing you think you understand well and one thing that you would like to ask your teacher

# Answers: C4.10 Required practical: Finding the reacting volumes of solutions of acid and alkali

## Connection

1. **Green**
2. **Hydrogen ion**
3. **$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$**

## Demonstration

1. Neutralisation
2. Pipette
3. Burette
4. Wear safety spectacles Wear gloves. Wear a laboratory coat.
5. 26.633 which should be rounded to 26.6 $cm^3$ (or 26.65 $cm^3$ if the burette is being read to 0.05 $cm^3$ ).
6 She included the rough titre. This is incorrect (it is not within 0.1 $cm^3$ of the other values in any case).
7. The data suggests that the burette was being read to the nearest 0.1 $cm^3$ i.e. 1 decimal place. The average cannot be more accurate than the individual titres. So the average needs to be quoted to 1 decimal place as well.
8. Moles BOH = (25/1000) × 0.12 = 0.003 mol/$dm^3$
9. 0.003 moles in 26.6 $cm^3$ . So moles in 1000 $cm^3$ = 0.03 × (1000/26.6) = 0.113 mol/$dm^3$
10. 63 g/mol
11. 0.113 × 63 = 7.12 g/$dm^3$
12. Moles NaOH in 10.0 $cm^3$ = (10/1000) × 0.25 = 0.0025. Ratio NaOH : HCl = 1:1. Moles HCl in 11.2 $cm^3$ = 0.0025. Moles HCl in 1,000 $cm^3$ = 0.0025 × (1000/11.2) = 0.223 mol/$dm^3$ .

# Lesson 11 C4.11 Strong and weak acids

## Connection

Q1. State an advantage of using a burette over a measuring cylinder

Q2. Describe what the end point of a titration is

Q3. State the type of reaction between an acid and an alkali

## Activation

**LI: Explain weak and strong acids by the degree of ionisation**

1. https://www.youtube.com/watch?v=_gYBbzkqrmE
2. Make a note of the title and the LI
3. Read pages 152-153
4. Give definitions for the key words on page 152
5. Draw fig 4.39 page 152

## Consolidation

Complete and self assess the relevant past paper question for this topic - From the C4 DIP file

## Extension

Make a note of one thing you think you understand well and one thing that you would like to ask your teacher

## Demonstration

Attempt questions 1-4
In 15 mins answer as many questions as you can.
Self mark the questions you have done making any necessary corrections in blue pen

Challenge yourself to answer as many as you can:
Green questions to GCSE Level 3
Blue questions to GCSE Level 6
Purple questions to GCSE Level 9

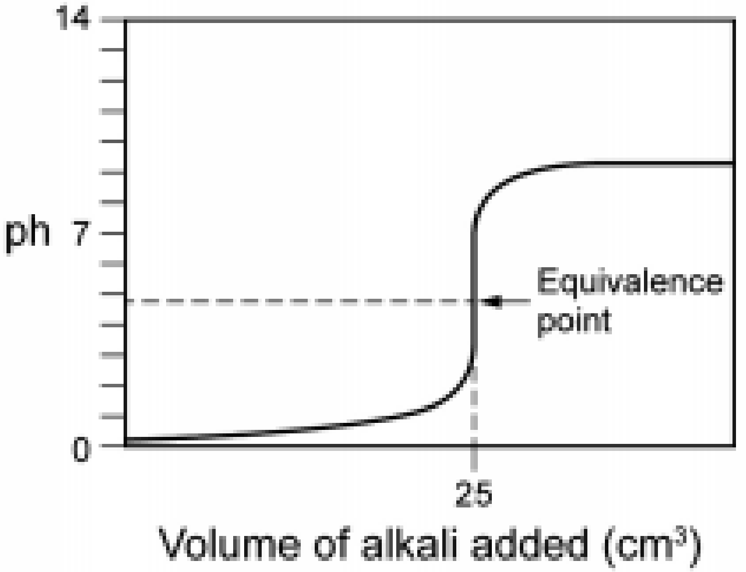
# Answers: C4.11 Strong and weak acids

## Connection

1. **A burette is much more accurate and precise than a measuring cylinder**
2. **The end point of a titration is the point where the indicator changes colour. It also refers to the volume of acid that has been added when the pH of the solution rapidly changes from alkaline to acidic**
3. **Neutralisation**

## Demonstration

1 Hydrogen ions and sulfate ions.
2 The citric acid molecule is not fully ionised. This means that pH is higher than for a strong acid - the concentration of hydrogen ions is lower.
3




4 60 g/dm3

# Lesson 12 C4.12 The process of electrolysis

## Connection

Q1. State an example of a strong acid and a weak acid

Q2. Explain why sulfuric acid is a strong acid

Q3. Explain how a dilute acid can also be a strong acid

## Activation

**LI: Identify reactions at electrodes during electrolysis**

https://www.youtube.com/watch?v=7uIlq_Ofzgw

1. Make a note of the title and the LI
2. Read pages 154-155
3. Give definitions for the key words on page 154
4. Draw fig 4.44 and fig 4.45 page 154

## Consolidation

Complete and self assess the relevant past paper question for this topic - From the C4 DIP file

## Extension

Make a note of one thing you think you understand well and one thing that you would like to ask your teacher

## Demonstration

Attempt questions 1-5
In 15 mins answer as many questions as you can.
Self mark the questions you have done making any necessary corrections in blue pen

Challenge yourself to answer as many as you can:
Green questions to GCSE Level 3
Blue questions to GCSE Level 6
Purple questions to GCSE Level 9

# Answers: C4.12 The process of electrolysis

## Connection

1. **Strong: Hydrochloric, Sulfuric, Nitric**

   **Weak: Ethanoic, Citric, Carbonic**

1. **Because it ionises completely in water**
2. **Neutralisation**

## Demonstration

1. In solid ionic compounds, the ions are fixed in the lattice and cannot move. When molten or aqueous, the ions can move and carry charge. Therefore, electricity can flow.

2. Copper at the cathode. Chlorine at the anode.

3. Aluminium ions, Al3+ are positive and attracted to the negative cathode.

4. $Cu^{2+} + 2e^{-} \rightarrow Cu$ $2Cl^{-} \rightarrow Cl_2 + 2e^{-}$

5. (a) $2Br^{-} \rightarrow Br_2 + 2e^{-}$
   (b) Mg
   (c) Magnesium is above hydrogen in the reactivity series. Magnesium is too reactive to be discharged. So the hydrogen ions in water are discharged.

# Lesson 13 C4.13 Electrolysis of molten ionic compounds

## Connection

Q1. State which electrode is positive and which is negative

Q2. State what a cation is, which electrode it moves to and why.

Q3. Write a half equation for the discharge of the chloride ion at the anode

## Activation

**LI: Explain how the ions of a molten electrolyte are discharged**

1. https://www.youtube.com/watch?v=87K8QsMl8nc
2. Make a note of the title and the LI
3. Read pages 156-157
4. Give definitions for the key words on page 156
5. Draw fig 4.47 page 156

## Demonstration

Attempt questions 1-4
In 15 mins answer as many questions as you can.
Self mark the questions you have done making any necessary corrections in blue pen

Challenge yourself to answer as many as you can:
Green questions to GCSE Level 3
Blue questions to GCSE Level 6
Purple questions to GCSE Level 9

## Consolidation

Complete and self assess the relevant past paper question for this topic - From the C4 DIP file

## Extension

Make a note of one thing you think you understand well and one thing that you would like to ask your teacher

# Answers: C4.13 Electrolysis of molten ionic compounds

**Connection**

1. **The negative electrode is the cathode**

   **The positive electrode is the anode**

1. **A cation is a positively charged ion, it moves to the cathode as the cathode is negatively charged. It is attracted to the cathode through electrostatic attraction**
2. **$2Cl^- - 2e^- \rightarrow Cl_2$**

**Demonstration**

1. Cathode. Positive ions are attracted to the negative cathode.

2. Anode. Negative ions migrate to the positive anode.

3. Sodium at the cathode since sodium ions are positive and the cathode is negative. Bromine at the anode since bromide ions are negative.

4. Cathode: $Cu^{2+} + 2e^- \rightarrow Cu$
   Anode: $2Br^- \rightarrow Br_2 + 2e^-$

# Lesson 14 C4.14 Using electrolysis to extract metals

## Connection

Q1. Bromine makes a negative ion. To which electrode will it move?

Q2. To which electrode will a copper ion move?

Q3. Write the half equations at the cathode and anode for the electrolysis of Potassium chloride

## Activation

**LI: Explain the process of the electrolysis of aluminium oxide**

1. https://www.youtube.com/watch?v=hOrGNtlN3sg
2. Make a note of the title and the LI
3. Read pages 158-159
4. Give definitions for the key words on page 158
5. Draw fig 4.49 page 159

## Consolidation

Complete and self assess the relevant past paper question for this topic - From the C4 DIP file

## Demonstration

Attempt questions 1-4
In 15 mins answer as many questions as you can.
Self mark the questions you have done making any necessary corrections in blue pen

Challenge yourself to answer as many as you can:
Green questions to GCSE Level 3
Blue questions to GCSE Level 6
Purple questions to GCSE Level 9

## Extension

Make a note of one thing you think you understand well and one thing that you would like to ask your teacher

# Answers: C4.14 Using electrolysis to extract metals

## Connection

1. **Bromide ions will move to the anode**
2. **Copper ions will move to the cathode**
3. **Half equation at cathode:**

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

   **Half equation at anode:**

   $$2Cl^{-} - 2e^{-} \rightarrow Cl_2$$

## Demonstration

1 Sodium is a very reactive metal and above carbon in the reactivity series. Therefore reduction by carbon will not work.

2 The negative oxide ions, $O^{2-}$ are attracted to the positive anode and form oxygen (oxidation since electrons are lost). $Al^{3+}$ ions are attracted to the negative cathode. At the cathode they are reduced (they gain electrons). Molten aluminium is formed.

3 The electrolysis cell has to be heated to high temperatures. Large electrical currents are needed. Both of these require a large amount of energy. This is very costly.

4 $O^{2-}$ ions are attracted to the positive anode. They transfer their electrons to the anode and oxygen forms. The electrons flow around the circuit where they are transferred to $Al^{3+}$ ions at the negative cathode. Aluminium forms.

5 Cathode: $Ca^{2+} + 2e^{-} \rightarrow Ca$ Anode: $2Cl^{-} \rightarrow Cl_2 + 2e^{-}$

6 Half equations: $2K^{+} + 2e^{-} \rightarrow 2K$ and $2F^{-} \rightarrow F_2 + 2e^{-}$ . Overall equation: $2KF \rightarrow 2K + F_2$

# Lesson 15 C4.15 Electrolysis of aqueous solutions

## Connection

Q1. State the name of the ore which aluminium is extracted from

Q2. Explain why electrolysis of aluminium is so expensive

Q3. Why is the purified ore of aluminium oxide mixed with cryolite?

## Activation

**LI: Explain the electrolysis of copper sulphate using inert electrodes**

https://www.youtube.com/watch?v=GrgYXk_NCec

1. Make a note of the title and the LI
2. Read pages 160-161
3. Give definitions for the key words on page 160
4. Draw fig 4.50 page 160

## Consolidation

Complete and self assess the relevant past paper question for this topic - From the C4 DIP file

## Extension

Make a note of one thing you think you understand well and one thing that you would like to ask your teacher

## Demonstration

Attempt questions 1-5
In 15 mins answer as many questions as you can.
Self mark the questions you have done making any necessary corrections in blue pen

Challenge yourself to answer as many as you can:
Green questions to GCSE Level 3
Blue questions to GCSE Level 6
Purple questions to GCSE Level 9

# Answers: C4.15 Electrolysis of aqueous solutions

## Connection

1. **Bauxite**
2. **The process requires a great deal of energy which is very expensive**
3. **Cryolite brings down the melting point which means less energy is required to melt it**

## Demonstration

1. Silver would be discharged in preference to hydrogen as it is lower in the reactivity series.

2. Copper would still be discharged at the cathode as it is lower in the reactivity series than sodium. Chlorine would be discharged in preference to oxygen, (provided it was present in sufficient concentration) as chlorine is lower in the reactivity series than hydroxide ions.

3. At the cathode: $2H^+ + 2e^- \rightarrow H_2$
   At the anode : $4OH^- - 4e^- \rightarrow 2H_2O + O_2$

4. At the cathode: $Cu^{2+} + 2e^- \rightarrow Cu$
   At the anode: $2Cl^- - 2e^- \rightarrow Cl_2$

The discharge of copper is a reduction reaction as electrons are gained.
The discharge of chlorine is an oxidation reaction as electrons are lost.

5. $2H^+ + 2e^- \rightarrow H_2$ $4OH^- - 4e^- \rightarrow 2H_2O + O_2$

# Lesson 16 C4.16 Required practical: Investigating what happens when aqueous solutions are electrolysed using inert electrodes

## Connection

Q1. Explain whether copper or hydrogen is discharged first during the electrolysis of aqueous copper sulphate

Q2. Describe what observations you would see during electrolysis of copper sulfate

Q3. Write half equations for the electrode reactions in the electrolysis of copper sulfate?

## Activation

**LI: Plan experiments to make observations and test hypotheses**

https://www.youtube.com/watch?v=tCHE_7QeRUc

1. Make a note of the title and the LI
2. Read pages 162-163
3. Give definitions for the key words on page 162
4. Draw fig 4.51 page 162
5. Carry out required practical

## Consolidation

Complete and self assess the relevant past paper question for this topic - From the C4 DIP file

## Extension

Make a note of one thing you think you understand well and one thing that you would like to ask your teacher

## Demonstration

Attempt questions 1-9
In 15 mins answer as many questions as you can.
Self mark the questions you have done making any necessary corrections in blue pen

Challenge yourself to answer as many as you can:
Green questions to GCSE Level 3
Blue questions to GCSE Level 6
Purple questions to GCSE Level 9

# Answers: C4.16 Req practical Investigating what happens when aqueous solutions are electrolysed using inert electrodes

## Connection

1. **Copper is discharged first as it is less reactive than hydrogen**
2. **The blue colour of the copper sulfate solution would fade. The cathode would be coated in copper. The anode would produce bubbles**
3. **At the cathode:**

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

   **At the anode:**

   $$4OH^- - 4e^- \rightarrow 2H_2O + O_2$$

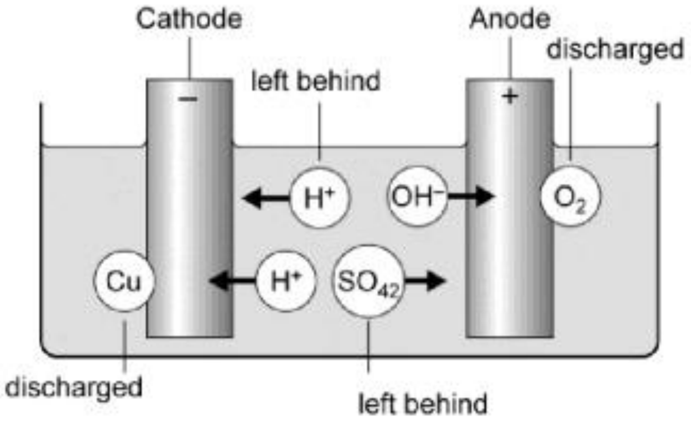
## Demonstration

1. The cathode (if the metal is less reactive than $H^+$ ) since positive metal ions are attracted to the negative cathode.
2. Hydrogen. Insert a lighted splint into the gas. A popping sound indicates the presence of hydrogen gas.
3. No they don't. The order of increasing reactivity for the metals is: Silver, copper, zinc, sodium The less reactive the metal the easier it is to deposit on the cathode. The products formed (going down the column) were: Hydrogen, copper, hydrogen, silver.
4. The reactivity series.
5. The order of increasing reactivity for the metals is: silver, copper, $H^+$, zinc, sodium
   The products formed (going down the column) were: hydrogen, copper, silver. Copper and silver are less reactive than $H^+$ so are discharged. Sodium and zinc are more reactive than $H^+$ so $H^+$ is discharged as $H_2$

Answers: <u>C4.16 Req practical Investigating what happens when aqueous solutions are electrolysed using inert electrodes (continued)</u>

# Demonstration

6.




7. Zinc Sulfate: Oxygen. Copper sulfate: Oxygen. Sodium chloride: Chlorine. Silver nitrate: Oxygen.

8. The order of increasing reactivity for the metals is: Silver, copper, $H^+$ , zinc, sodium So $Cu^{2+}$ and $Ag^+$ more readily accept electrons (and are reduced) than $H^+$ ions. Sodium and zinc will not accept electrons in preference to $H^+$ since they are more reactive than $H^+$ . The overall conclusion is that if the metal is less reactive than $H^+$ , it will deposit on the cathode.

9. The overall conclusion is that if the metal is less reactive than $H^+$ , it will deposit on the cathode. This contradicts the original hypothesis. However, the evidence supports the alternative hypothesis "Does the ability to deposit on an electrode link to the reactivity of the metal ion in solution". This only applies to inert electrodes. If other electrodes were used (e.g. metal electrodes) then more experiments would have to be carried out.

# Lesson 17 C4.17 Key concept: Electron transfer, oxidation and reduction

## Connection

Q1. State what OIL RIG stands for

Q2. State which electrode the copper will deposit onto in the electrolysis of copper sulfate

Q3. State what gas is given off at the anode and how to test for it

## Activation

**LI: Explain why atoms gain or lose electrons**

https://www.youtube.com/watch?v=jyvcVjrZnJA

1. Make a note of the title and the LO
2. Read pages 164-165
3. Give definitions for the key words on page 164
4. Draw fig 4.58 page 165

## Consolidation

Complete and self assess the relevant past paper question for this topic - From the C4 DIP file

## Demonstration

Attempt questions 1-6
In 15 mins answer as many questions as you can.
Self mark the questions you have done making any necessary corrections in blue pen

Challenge yourself to answer as many as you can:
Green questions to GCSE Level 3
Blue questions to GCSE Level 6
Purple questions to GCSE Level 9

## Extension

Make a note of one thing you think you understand well and one thing that you would like to ask your teacher

# Answers: C4.17 Key concept: Electron transfer, oxidation and reduction

## Connection

1. **Oxidation Is Loss, Reduction Is Gain**
2. **The cathode**
3. **Oxygen gas is given off and to test for it, it should relight a glowing splint**

## Demonstration

1. Sodium loses one electron and transfers it to chlorine which gains one electron. Both end up with the stable electron arrangement of a noble gas. There is a force of attraction between the two oppositely charged ions produced.

2. They do not need to lose or gain electrons from their outer shell.

3. (a) Zinc is oxidised since oxygen is added. In terms of electrons, zinc loses 2 electrons and is oxidised to $Zn^{2+}$. Oxygen gains electrons and is reduced to oxide ions, $O^{2-}$.
   (b) Iron(II) oxide is reduced to iron since oxygen is removed.
   $Fe^{2+}$ gains 2 electrons which is reduction.

4. (a) Electrons are lost by the hydroxide ions ( $OH^-$ ).
   This is oxidation ( $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ ).
   (b) Silver gains one electron and is reduced ($Ag^+ + e^- \rightarrow Ag$).

5. Both calcium and magnesium have two outer shell electrons. However, calcium has more shells and the outer shell is further from the nucleus. So the attraction between the outer electrons and the nucleus is less in calcium.

6. Magnesium has a greater nuclear charge. Therefore, the attraction for the outer shell electrons is greater than for sodium. So more energy is required to remove magnesium’s electrons. Also, to achieve full shell stability, magnesium has to lose 2 electrons. This takes much more energy than losing 1 electron as in sodium.

# C4 - Revision

## Connection

Q1. If oxidation occurs at the anode. What happens at the cathode?

Q2. Why is it so expensive to extract metals from their ores?

Q3. Why will magnesium metal displace copper from copper sulfate solution, but not sodium from sodium sulfate solution?

## Activation

**LI: Create a topic summary sheet**

1. Fold an A3 sheet so it is divided into 8 sections
2. Look back over you lessons and group them into 8 main headings
3. Summarise the key points into each section, use keywords and diagrams and symbols rather than sentences

## Demonstration

Test yourself by working with the person sitting next to you by talking though each box on your summary sheet and seeing how many key facts you can remember

## Consolidation

Look though the relevant past paper questions for this topic - From the B7 DIP file – see if you can complete any additional questions

## Extension

Make a list of anything that you would like to ask your teacher to go over again

# Answers: C4 - Revision

## Connection

**1** Reduction
**2** It requires huge amounts of expensive energy to extract metals from their ore
**3** Magnesium is more reactive than copper and so displaces it, but it is less reactive than sodium and so there is no reaction

# Research on Electrolysis Leads to a More Environmentally Friendly Way of Producing Cement

In a recent study by researchers from MIT, electrolysis can now be used to produce cement. In their study published in Proceedings of the National Academy of Sciences of the United States of America, the scientists report that this new production technique would reflect a decrease in the emissions associated with cement production. And, in addition to that, the gas streams that would be produced may be used in other processes.

**Cement Industry Emissions Today**

In a statistical study in 2015 according to Carbon Brief, the cement industry is responsible for about eight percent of the total carbon dioxide emissions worldwide. That is 2.8 billion tons of carbon dioxide per year, emitted from the cement industry alone.

**MIT's Proposed Process**

Electrolysis is the process by which a substance seperates at a chemical level, due to an electric current that is allowed to pass through a solution. From the knowledge that the electrolysis of solutions that have pH values near neutral causes the formation of a pH gradient.

Upon electrolysis, the pH gradient formed in the research had an acid at the anode and a base at the cathode. Ground up limestone-or calcite, the same compound found in chalk-was then added to the acid to produce lime water precipitate. When heated with silica, the precipitate then decomposes to lime and combines with the silica to form alite, which is a mineral that is found abundant in Portland cement.

**Reusable Byproducts**

The researchers made use of an electrochemical decarbonation reactor, which produced a mixed stream of oxygen gas and carbon dioxide at the anode and hydrogen gas at the cathode. The mixed stream of oxygen gas and carbon dioxide could be used as an oxygen-enhanced fuel in the kiln. This would result to more efficient burning of fossil fuels, compared to the use of air, which naturally contains nitrogen. This is because oxygen-enhanced combustion would not require an increase in temperature for any nitrogen. At the same time, there would be no harmful nitrous oxide emissions. Alternatively, the carbon dioxide can be captured directly, without using expensive units like an amine scrubber. At the cathode, the hydrogen gas produced could be used as green fuel to power some equipment at the plant. This is especially important since a lot of researches have proven that the use hydrogen gas as a source of energy is very advantageous.

**Research Implications**

In addition to the environmentally friendly results from the research, the process can be seen to work using renewable electricity instead of fossil fuels to generate heat. The process could be scaled up according to necessity, after further development, the researchers say.

**C4: Questions**

1a. State what the process of electrolysis does
 b. what percentage of carbon dioxide emissions is the cement industry responsible for?
 c. In the research, what gas was produced at the cathode?

2a. Explain, with references to charges, why the gas in your answer to 1c was produced at the cathode
 b. Explain why the use of the decarbonation reactor results in more efficient burning of fossil fuels
 c. Explain why there would also be a decrease in nitrous oxide emissions

3a. Choose three advantages of using electrolysis to produce cement
 b. Evaluate the impact on the environment of using electrolysis to produce cement
 c. The process produces hydrogen gas. Is this a good thing? Explain why.

# DART C4: Electrolysis - Answers

1a. From the text :
“Electrolysis is the process by which a substance separates at a chemical level, due to an electric current that is allowed to pass through a solution. “
1b. From the text :
“the cement industry is responsible for about eight percent of the total carbon dioxide emissions worldwide. “
1c. Hydrogen gas

2a. Hydrogen gas was produced at the cathode. Hydrogen ions in solution have a charge of +1. These positive ions(cations) will be attracted to the negative cathode
2b. The decarbonation reaction releases oxygen in concentrations that are higher than air. This would mean a more efficient burning of fossil fuels.
2c. There would be no nitrous oxide emission due to the fact that the fossil fuel is not being burned in air, which contains nitrogen

3a. Any three from the following:
- more efficient burning of fossil fuels
- no harmful nitrous oxide emissions
- the hydrogen gas produced could be used as green fuel to power some equipment at the plant
- the carbon dioxide can be captured directly, without using expensive units like an amine scrubber

3b. An advantage of using electrolysis to produce cement is the possibility of releasing less carbon dioxide into the atmosphere. This is because carbon dioxide can be directly removed from the process at the anode. Also the production of oxygen leads to the more efficient burning of fossil fuels meaning less CO2 produced per unit fossil fuel.
Another advantage would be that the process can be done using renewable electricity instead of fossil fuels. This would further reduce the impact on the environment.
Conversely, whilst the long term effects of this process may be environmentally friendly, the initial cost of the design and development of this may be high, both financially as well as environmentally.
3c. producing hydrogen gas can be good as it can be used as a fuel for combustion instead of fossil fuels. This means that fewer greenhouse gases will be released as hydrogen burning only releases water vapour as a product.

| Attainment Band : | C4 Chemical Changes AQA) |
|---|---|
| | **Knowledge and Understanding** |
| **Yellow Plus/ Yellow** | Identify metal oxides as bases or alkalis.<br>Deduce an order of reactivity of metals based on experimental results.<br>Interpret or evaluate information on specific metal extraction processes.<br>Write ionic equations for displacement reactions.<br>Use half equations to describe oxidation and reduction.<br>Derive a formula for a salt from its ions.<br>Investigate pH changes when a strong acid neutralises a strong alkali.<br>Explain the terms dilute and concentrated as the amounts of substances dissolved.<br>Write half equations for the electrode reactions and complete and balance half equations.<br>Predict the products of the electrolysis of binary ionic compounds in the molten state.<br>Explain which non-metals are formed at the anode in preference.<br>Predict the products of the electrolysis of aqueous solutions containing a single ionic compound.<br>Represent reactions at electrodes by half equations. |
| **Blue** | Explain reduction and oxidation by loss or gain of oxygen.<br>Explain how the reactivity is related to the tendency of the metal to form its positive ion.<br>Explain how extraction methods depend on metal reactivity.<br>Use the reactivity series to predict displacement reactions.<br>Write full balanced symbol equations for making salts.<br>Explain how to name a salt.<br>Use the pH scale to identify acidic or alkaline solutions.<br>Describe neutralisation through the effect on hydrogen ions and pH.<br>Identify reactions at electrodes during electrolysis.<br>Describe the products of molten binary electrolytes.<br>Explain the process of the electrolysis of aluminium oxide.<br>Explain which metals (or hydrogen) are formed at the cathode in preference.<br>Predict the products of the electrolysis of aqueous solutions. |
| **Green** | Identify that metals react with oxygen to form metal oxides.<br>Describe the reactions, if any, of metals with water or dilute acids to place these metals in order of reactivity.<br>Identify substances oxidised or reduced by gain or loss of oxygen.<br>Use experimental results of displacement reactions to confirm the reactivity series.<br>Describe how to make salts from metals and acids.<br>Describe how to make pure, dry samples of soluble salts.<br>Describe the use of universal indicator to measure pH.<br>Explain weak and strong acids by the degree of ionisation.<br>Describe the use of an electrolyte, anode and cathode in an electrolysis.<br>Explain why electrolytes need to be molten to conduct electricity.<br>Explain why some metals need to be extracted by electrolysis.<br>Use apparatus to electrolyse aqueous solutions in the laboratory.<br>Explain the electrolysis of copper sulfate using inert electrodes. |
| **White** | Some elements of the above have been achieved |