# Topic Exploration Pack

# Exploration of the electrolysis of various molten ionic liquids and aqueous ionic solutions

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## Instructions and answers for teachers

These instructions cover the learner activity section which can be found on [page 16](#_Learner_Activity). This Topic Exploration Pack supports OCR GCSE (9-1) Gateway Science Chemistry A and the Twenty First Century Science Chemistry B qualifications.

**When distributing the activity section to the learners either as a printed copy or as a Word file you will need to remove the teacher instructions section.**

### Mapping to specification level (Learning outcomes)

**GCSE (9–1) Gateway Science Chemistry A/Combined Science A**

C3.4a recall that metals (or hydrogen) are formed at the cathode and non-metals are formed at the anode in electrolysis using inert electrodes

C3.4b predict the products of electrolysis of binary ionic compounds in the molten state

C3.4c describe competing reactions in the electrolysis of aqueous solutions of ionic compounds in terms of the different species present

C3.4d describe electrolysis in terms of the ions present and reactions at the electrodes

C3.4e describe the technique of electrolysis using inert and non-inert electrodes

**GCSE (9–1) Twenty First Century Science Chemistry B/Combined Science B**

C3.2.5 explain why electrolysis is used to extract some metals from their ores

C3.3.1 describe electrolysis in terms of the ions present and reactions at the electrodes

C3.3.3 recall that metals (or hydrogen) are formed at the cathode and non-metals are formed at the anode in electrolysis using inert electrodes

### Introduction

In the learning of the electrolysis of aluminium, studies have shown that learners who are aware of the ‘big picture’ understand and ‘get to grips’ with the smaller concepts easier as they comprehend the reasons.

Electrolysis is the splitting of compounds by using electricity. It is used in industry to extract and purify metals as well as non-metals.

### The electrolysis of aluminium

This would make an ideal research project or homework in preparation for the taught material.

The following YouTube video as produced by the RSC can be used with the five questions shown below.

<http://www.youtube.com/watch?v=WaSwimvCGA8>

1. Why is cryolite added?

To make the material conduct electricity.

1. What are the names of the positive and the negative electrodes?

anode and cathode

1. Which ions move to which electrode?

The positive ions (cations) move to the cathode and the negative ions (anions) move to the anode

1. How long does each carbon electrode last for?

28 days

1. What happens to the remaining carbon when the electrode has worn out?

It is crushed and formed into another electrode.

Key terms: anode (‘+ve’ electrode), cathode (‘-ve’ electrode), cation (‘+ve’ ion) and anion (‘-ve’ ion)

### Part 1

### Introduction

**Objective**

To recall that, in electrolysis using inert electrodes, metals (or hydrogen) are formed at the cathode and non-metals are formed at the anode. It is assumed that learners have a full understanding of balancing equations, ionic bonding and the workings of a simple electrical circuit.

**Research**

Studies have shown that learners have issues understanding that ions dissolve as neutral atoms. This leads to misconceptions with regards to what carries the charge. They believe that the solution carries electrons through it. This then leads on to the learners forming their own conceptions that the positively charged ions lose electrons and the negatively charged ions gain electrons from the solution.

*Journal of Chemical education 74,7, 1997, 819-823*

### Activities

### Activity 1: The electrolysis of distilled water and tap water using a Hoffman Voltameter

This activity is designed to get the learners to understand that the electrical charge is carried by the ions dissolved in the solution rather than electrons given by the anode and the cathode to allow electrolysis to happen.

<http://www.nuffieldfoundation.org/practical-physics/electrolysis-water>

This experiment should be carried out with distilled water and tap water.

Learners will see that the tap water will electrolyse and the distilled water will not.

*At this point, learners should be given the opportunity to discuss why this might be the case.*

The tap water contains dissolved ions which are necessary for electrolysis to take place as they carry the charge through the solution. Electrons cannot be being transferred into the solution from the electrodes. The students should then be given commercially available water bottles which will have the ion content on the label showing that ions are dissolved in the water. This overcomes the misconception ideas regarding how the current flows and that the two electrodes do not push electrons into the solution to allow a current to flow.

### Activity 2: Ions in solutions of bottled water

The data in **Table 1** was obtained from the website:

<http://www.cleanairpurewater.com/best_bottled_water.html>

With this task, learners could be asked to choose which water they would use from the table, and why. This could be set as a homework task.

**Table 1: the pH and mineral content of spring and mineral waters**

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
|  | **Arrowhead** | **Poland Springs** | **Fiji** | **Calistoga** | **Evian** | **San Pellegrino** | **Trinity Springs** | **Volvic** | **Perrier** |
| **pH** | 6.12-7.6 | 6.02-7.6 | 7.5 | 5.84-6.71 | 7.18 | 7.7 | 9.0 | 7 | 5.46 |
| **calcium** | 6-53 | 3.7-8.2 | 17 | 6.0-8.4 | 78 | 208 |   | 10 | 147.3 |
| **chloride** | .73-8.2 | 1.9-8.8 |   | .74-2.6 | 2.2 | 74.3 | .7 | 8 | 21.5 |
| **bicarbonate** | 42-190 | 7.2-20 |   | 36-44 | 357 | 135.5 |   | 65 | 390 |
| **fluoride** | 0.05-1.2 | .052-.20 |   | Nd |   | .52 | 1.3 |   | 0.12 |
| **magnesium** | 1.5-20 | .76-1.4 | 13 | 1.5-2.9 | 24 | 55.9 |   | 6 | 3.4 |
| **nitrate** | .05-.75 | .13-.75 |   | .06-.57 | 3.8 | .45 |   | 1 | 18 |
| **potassium** | 0.7-4.4 | .59-.74 |   | 1.8-3.8 | .75 | 2.7 |   | 6 | 0.6 |
| **sodium** | 2.1-20 | 2.4-4.7 |   | 3.8-9.5 | 5 | 43.6 | 19.3 | 9 | 9 |
| **sulfates** | .64-32 | .81-5.1 |   | Nd 2.1 | 10 | 549.2 | 6 | 7 | 33 |

Teachers could also show learners that minerals are present in the water by boiling a sample of bottled water and showing that there is a residue left over at the end.

This can be done by carrying out the experiment from the website below.

http://www.rsc.org/learn-chemistry/resource/res00001783/to-find-out-if-tap-water-and-sea-water-contain-dissolved-solidsThe experiment uses seawater but the same principle can be used for bottled water.

### Activity 3: The electrolysis of molten zinc chloride and potassium iodide

In the experiments shown in the websites below, the learners are aiming to identify the products at each electrode and show that different ions travel (or are attracted) to different electrodes (anions are attracted to the anode and cations are attracted to the cathode).

<http://www.nuffieldfoundation.org/practical-chemistry/electrolysis-zinc-chloride>

<http://www.nuffieldfoundation.org/practical-chemistry/reaction-zinc-iodine>

Learners should now have an understanding of the fact that ions carry the charge in a solution and that they dissociate in solution (ie are in their ionic state). In these experiments, the products at each electrode should enable learners to understand that the anions and the cations move to opposing electrodes and are formed there.

The products being produced at each electrode come only from ions that are in solution so learners will get the idea that the positive and negative ions travel to the opposing electrodes. This also includes hydrogen, as a positively charged ion, as shown with the Hoffman Voltammeter experiment.

**Now give out the cards shown and ask the learners to match the cards from Set 1 with the corresponding cards from Set 2.**

The numbers in red show the correct answers.

### Set 1

cathode

1

anode

2

cation

3

anion

4

ion

5

### Set 2

The negatively charged electrode.

1

The positively charged electrode.

2

The positively charged species that moves to the cathode.

4

The negatively charged species that moves to the anode.

3

The current is
carried by these.

3, 4 and 5

The charged species that exists in the molten state or in solution.

This carries the charge.

5

# **Part 2**

### Introduction

**Objective**

To predict the products of electrolysis of binary ionic compounds in the molten state.

**Research**

As has already been discussed, learners struggle with the concept that in the molten state the ions are free to move about. However, it has also been shown that learners also struggle with the diagrammatic representations shown in books, which means they have problems when it comes to understanding half equations.

*Journal Chem Educ 1999, 76, p853*

**The electrolysis of molten zinc chloride**

During the electrolysis of molten zinc chloride, a different reaction happens at each electrode and this can be represented by separate equations for each electrode. These are called half equations.

**Reaction at the cathode**

At the cathode (the negative electrode), the positively charged species (in this case, zinc ions, Zn+) move towards it and collects two electrons (e-) from the electrode. Remember, the electrons come from the electrode, not the solution.

The half equation for this reaction is:

Zn2+(l) + 2e- Zn(l)

**Reaction at the anode**

At the anode (the positive electrode), the negatively charged species (in this case, chloride ions, C*l* –) move towards it and discharges its electron there.

The half equation (currently incomplete) for this reaction is:

C*l* –(l) C*l* + 1e-

However, the overall electrical charge and species must be balanced. Therefore, this half equation must be multiplied by 2 as two electrons were taken by zinc and a chlorine molecule is Cl2.

Hence the complete half equation for this reaction is:

2C*l* –(l) C*l*2(g) + 2e-

When writing half equations for the anode, it is important that the electron is part of the product and is therefore on the right hand side of the equation.

At this point we should remember that at the anode, the ions lose electrons so they have been oxidised and at the cathode, the ions receive electrons so they have been reduced.

**So a rule is made: any reaction at the cathode is a reduction reaction and any reaction at the anode is an oxidation reaction.**

### Activities

### Activity: Half equation worksheet

Complete the following half equations for the electrolysis of the following molten compounds:

1. NaC*l*

2Na+(l) + 2e- 2Na(l)

2C*l* –(l) C*l*2(g) + 2e-

1. KBr

2K+(l) + 2e- 2K(l)

2Br –(l) Br2(g) + 2e-

1. MgC*l*2

2Mg2+(l) + 4e- 2Mg(l)

4C*l* –(l) 2C*l*2(g) + 4e-

1. Li2O

4Li+(l) + 4e- 4Li(l)

2O2-(l) O2(g) + 4e-

1. A*l*2O3

4A*l*3+(l) + 12e-  4A*l* (l)

6O2-(l)  3O2(g) + 12e-

# **Part 3**

### Introduction

**Objective**

To describe competing reactions in terms of the different species present in the electrolysis of aqueous solutions of ionic solutions.

(Consider using CuSO4 and NaC*l* as practicals.)

**Research**

It has been shown that when another variable such as water is added (aqueous solutions) then learners become confused as to what is happening in the solution and what is happening within the electrolysis cell.

*Chem Educ research practice 2012, 13, 471-483.*

The following ideas first consider what is actually happening when electrolysis is carried out practically. The idea about what is happening in solution is explored by utilising a role play within class.

This is intended to make learners produce evidence which will be contrary to what they believe may happen in molten solutions via the practical. The role play allows learners to physically see and act out what is happening within the aqueous solution before the idea of the reactivity series is introduced.

### Activities

The following practical, taken from the website shown below, will show that different aqueous solutions that are electrolysed produce different products which can be tested.

This therefore provides the evidence to show that aqueous solutions provide different products. The equations and reactions are considered in the resulting role play exercise which will undo any misconceptions that learners may have.

<http://www.nuffieldfoundation.org/practical-chemistry/identifying-products-electrolysis>

### Activity 1: The electrolysis of sodium chloride

When molten sodium chloride is electrolysed, the following reactions happen at the cathode and anode.

Cathode: 2Na+(l) + 2e- 2Na(l)

Anode: 2C*l* –(l) C*l*2(g) + 2e-

What happens when sodium chloride, dissolved in water, is electrolysed?

See the website: <http://www.nuffieldfoundation.org/practical-chemistry/colourful-electrolysis>

During the electrolysis of aqueous sodium chloride, chlorine is made at the anode and hydrogen, instead of sodium, is made at the cathode. This is because it is energetically easier to reduce the water. (Please remember that the sodium ions still move towards the cathode in the solution).

The half equations for each electrode are as follows:

Cathode: 2H2O(l) + 2e- H2(g) + 2OH-(aq)

Anode: 2C*l* –(aq) C*l*2(g) + 2e-

When tested, the solution will turn alkaline which confirms the formation of hydroxide ions as shown in the half equations. The presence of chlorine gas can be detected by moist litmus paper being ‘bleached’ white.

Learners will as the research suggests be confused at this point expecting to see sodium being produced at the cathode when Hydrogen gas is being produced (this can be tested via the squeaky pop method).

**Role Play exercise**

1. Two points can be chosen in the classroom such as two benches, one with a positive sign and one with a negative sign. This represents the electrodes in the system.
2. Create several A4 signs, two with **Na+**, two with **C*l*-**, eight with **H**, four with **O** and one with a picture of a bulb.
3. Choose two pairs of learners, with each pair linking arms.
4. Within each pair, one learner holds the Na+ sign while the other learner holds the *Cl*- sign.
5. These two pairs show an ionically bonded sodium chloride molecule.
6. Another learner can hold the sign of a bulb, showing that the electricity is flowing by holding it up.
7. The learners should then consider what happens when the bond is broken and the Na+ and the *Cl* – ions move to the cathode and anode respectively.

**Extension**

The model is then extended. This time, the same sodium and chloride ions are used but then a number of groups of three learners linking arms holding an H, an O and another H, depicting a molecule of H2O, are introduced.

This time, as ‘the bulb’ lights, the sodium and the chloride ions move to their electrodes but the water molecule breaks into a hydrogen ion and a hydroxide ion and they both move to their corresponding electrodes. The two chloride ions join to form chlorine gas and the hydrogen ions join to form a hydrogen gas molecule. However, the sodium ions and the hydroxide ions remain as spectator ions. The presence of the OH- ions shows the reason why the solution becomes alkaline.

### Activity 2: Reactivity Series

The previous reaction of sodium chloride in water suggests that some species are easier to remove than others. This can then be related to the reactivity series that the learners are familiar with (Table 2).

**Table 2: positive and negative ions**

|  |  |
| --- | --- |
| **Positive ions** | **Negative ions** |
| K+ | F– |
| Na+ | SO42- |
| Ca2+ | NO3 – |
| Mg2+ | *Cl* – |
| Al3+ | Br – |
| Zn2+ | I – |
| Fe2+ | OH– |
| Sn2+ |  |
| Pb2+ |  |
| H+ |  |
| Cu2+ |  |
| Ag+ |  |

The explanation of the table slightly differs as the reactivity series of metals considers the relative ease of removing electrons thus making it more reactive. However, this table considers the opposite i.e. the difficulty of putting an electron back in so the ease of removal by electrolysis increases down the table.

For example, if a solution contained K+ ions and H+, hydrogen gas would be evolved instead of potassium metal.

When carrying out an electrolysis reaction, the solvent can become part of the reaction, which is why hydrogen gas is given off when electrolysing sodium chloride solution.

This explains the reactions that you have carried out in the previous experiments.

# **Part 4**

### Introduction

In this part, teachers will consider specific reactions such as the electrolysis of copper (II) sulfate and sodium chloride.

This builds on the work carried out in the previous lessons explained by the role play exercise that the learners have carried out. The learners should now be able to predict what happens with these reactions, however, this does change depending upon the electrode used.

**Objective**

To describe electrolysis in terms of the ions present and reactions at the electrodes.

The electrolysis of copper sulfate solution

Please see the website shown for this next experiment:

<http://www.nuffieldfoundation.org/practical-chemistry/electrolysis-copperii-sulfate-solution>

The above experiment can be carried out using both graphite electrodes and copper electrodes to consider the way in which we can extract copper to plate something or purify a sample of copper.

However, learners should be asked to predict what happens based upon the reactivity table  **(Table 2)** and why, when a graphite electrode is used, oxygen is evolved and the colour of the solution disappears, whereas when copper electrodes are used, the mass of one electrode increases and the other one deceases but the solution does not change colour.

The half equations are as follows:

Graphite anode: 2H2O(l) O2(g) + 4H+(aq) +4e-

Copper anode: Cu(s) Cu2+(aq) + 2e-

Graphite or copper cathode: Cu2+(aq) +2e- Cu(s) (for the graphite anode the stoichiometry needs to be doubled to make the charges balance)

The conclusion should be that when both electrodes are copper, there is just a transfer of pure copper from one electrode to the other, with the overall numbers of copper ions remaining the same in solution. This is also the reason why there is no overall change in colour of solution.

In addition, oxygen is given off as the solution is slightly acidic due to the presence of the SO4-2 ions. So another conclusion can be gained from this: if the solution is alkaline, hydrogen gas is given off and if the solution is acidic, oxygen gas is given off when using aqueous solutions.

**Example past paper questions that can be used as an end of topic assessment:**

<http://www.ocr.org.uk/Images/135991-question-paper-unit-b741-02-chemistry-modules-c1-c2-c3-higher-tier.pdf>

– May 2012 Gateway science question 12

<http://www.ocr.org.uk/Images/175767-question-paper-unit-b741-02-chemistry-modules-c1-c2-c3-higher-tier.pdf>

– May 2013 Gateway science question 6b

<http://www.ocr.org.uk/Images/175769-question-paper-unit-b742-02-chemistry-modules-c4-c5-c6-higher-tier.pdf>

– May 2013 Gateway Science question 11

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# Topic Exploration Pack

# Exploration of the electrolysis of various molten ionic liquids and aqueous ionic solutions

## Learner Activity

### Activity 1: The electrolysis of aluminium

This would make an ideal research project or homework in preparation for the taught material.

The following YouTube video as produced by the RSC can be used with the five questions shown below. <http://www.youtube.com/watch?v=WaSwimvCGA8>

1. Why is cryolite added?
2. What are the names of the positive and the negative electrodes?
3. Which ions move to which electrode?
4. How long does each carbon electrode last for?

1. What happens to the remaining carbon when the electrode has worn out?

###

### Activity 2: Ions in solutions of bottled water

Choose which water you would use from the table, and explain why.

**Table 1: the pH and mineral content of spring and mineral waters**

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
|  | **Arrowhead** | **Poland Springs** | **Fiji** | **Calistoga** | **Evian** | **San Pellegrino** | **Trinity Springs** | **Volvic** | **Perrier** |
| **pH** | 6.12-7.6 | 6.02-7.6 | 7.5 | 5.84-6.71 | 7.18 | 7.7 | 9.0 | 7 | 5.46 |
| **calcium** | 6-53 | 3.7-8.2 | 17 | 6.0-8.4 | 78 | 208 |   | 10 | 147.3 |
| **chloride** | .73-8.2 | 1.9-8.8 |   | .74-2.6 | 2.2 | 74.3 | .7 | 8 | 21.5 |
| **bicarbonate** | 42-190 | 7.2-20 |   | 36-44 | 357 | 135.5 |   | 65 | 390 |
| **fluoride** | 0.05-1.2 | .052-.20 |   | Nd |   | .52 | 1.3 |   | 0.12 |
| **magnesium** | 1.5-20 | .76-1.4 | 13 | 1.5-2.9 | 24 | 55.9 |   | 6 | 3.4 |
| **nitrate** | .05-.75 | .13-.75 |   | .06-.57 | 3.8 | .45 |   | 1 | 18 |
| **potassium** | 0.7-4.4 | .59-.74 |   | 1.8-3.8 | .75 | 2.7 |   | 6 | 0.6 |
| **sodium** | 2.1-20 | 2.4-4.7 |   | 3.8-9.5 | 5 | 43.6 | 19.3 | 9 | 9 |
| **sulfates** | .64-32 | .81-5.1 |   | Nd 2.1 | 10 | 549.2 | 6 | 7 | 33 |

Preferred water:

Reason for choice:

### Activity 3: The electrolysis of molten zinc chloride and potassium iodide

Match the cards from Set 1 with the associated cards from Set 2.

### Set 1

cathode

anode

cation

anion

ion

### Set 2

The negatively charged electrode.

The positively charged electrode.

The positively charged species that moves to the cathode.

The negatively charged species that moves to the anode.

The current is
carried by these.

The charged species that exists in the molten state or in solution.

This carries the charge.

### **Part 2**

### The electrolysis of molten zinc chloride

During the electrolysis of molten zinc chloride, a different reaction happens at each electrode and this can be represented by separate equations for each electrode. These are called half equations.

**Reaction at the cathode**

At the cathode (the negative electrode), the positively charged species (in this case, zinc ions, Zn+) move towards it and collects two electrons (e-) from the electrode. Remember, the electrons come from the electrode, not the solution.

The half equation for this reaction is:

Zn2+(l) + 2e- Zn(l)

**Reaction at the anode**

At the anode (the positive electrode), the negatively charged species (in this case, chloride ions, C*l* –) move towards it and discharges its electron there.

The half equation (currently incomplete) for this reaction is:

C*l* –(l) C*l*(g) + 1e-

However, the overall electrical charge and species must be balanced. Therefore, this half equation must be multiplied by 2 as two electrons were taken by zinc and a chlorine molecule is C*l*2.

Hence the complete half equation for this reaction is:

2C*l* –(l) C*l*2(g) + 2e-

When writing half equations for the anode, it is important that the electron is part of the product and is therefore on the right hand side of the equation.

At this point we should remember that at the anode, the ions lose electrons so they have been *oxidised* and at the cathode, the ions receive electrons so they have been *reduced*.

**So a rule is made: any reaction at the cathode is a reduction reaction and any reaction at the anode is an oxidation reaction.**

### Activity: Half equation worksheet

Complete the following half equations for the electrolysis of the following molten compounds:

1. NaC*l*
2. KBr
3. MgC*l*2
4. Li2O
5. Al2O3

### **Part 3**

### Activity 1: The electrolysis of sodium chloride

When molten sodium chloride is electrolysed, the following reactions happen at the cathode and anode.

Cathode: 2Na+(l) + 2e- 2Na(l)

Anode: 2C*l* –(l) C*l*2(g) + 2e-

What happens when sodium chloride, dissolved in water, is electrolysed?