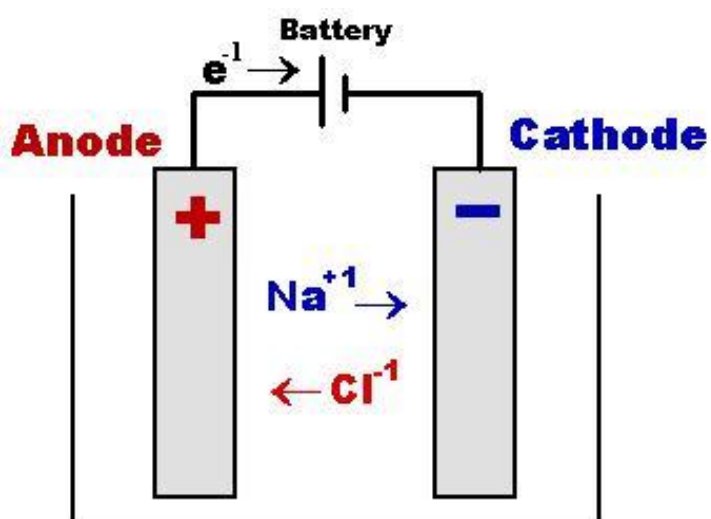
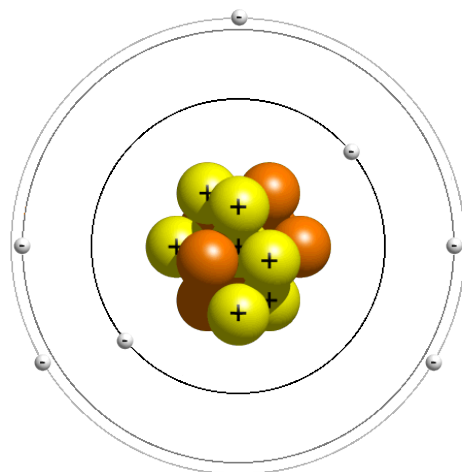


2010 GCSE Chemistry Coursework

Electrolysis



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Introduction to Electrolysis

In chemistry, the production of chemical changes by passing an electric current through a solution or molten salt (the electrolyte), resulting in the migration of ions to the electrodes: positive ions (**cations**) to the negative electrode (**cathode**), and negative ions (**anions**) to the positive electrode (**anode**).

During electrolysis, the ions react with the electrode, either receiving or giving up electrons. The resultant atoms may be liberated as a gas, or deposited as a solid on the electrode, in amounts that are proportional to the amount of current passed, as discovered by English chemist Michael Faraday. For instance, when acidified water is electrolyzed, hydrogen ions (H^+) at the cathode receive electrons to form hydrogen gas; hydroxide ions (OH^-) at the anode give up electrons to form oxygen gas and water.

One application of electrolysis is **electroplating**, in which a solution of a salt, such as silver nitrate (AgNO_3), is used and the object to be plated acts as the negative electrode, thus attracting silver ions (Ag^+). Electrolysis is used in many industrial processes, such as coating metals for vehicles and ships, refining **bauxite** into aluminium, and the chlor-alkali industry, in which **brine** (sodium chloride solution) is electrolyzed to produce **chlorine**, hydrogen, and **sodium hydroxide** (caustic soda); it also forms the basis of a number of electrochemical analytical techniques, such as polarography.

Introduction Continued

Electrolysis formula

$$I = nAve$$

$$I = n \times A \times v \times e$$

I = Current Flow In Electrolysis

n = Number of Ions

A = Surface Area of Electrodes

v = Velocity of Ions

e = Charge of Ions

Ohms law doesn't apply to electrolysis ($V=I \times R$)

By increasing any of the above, overall, the current increase, so current is proportional to nAve.

I can predict that by increasing the charge of the ions, I can increase current.

$$I = nAve \quad - \text{current flow due to a charge carrier}$$

Variables

- Time
- Current
- Temperature
- Solution
- Molarity/Concentration of solution
- Charge of Ions
- Quantity of solution
- Size of electrodes
- Distance between electrodes
- Surface area of electrodes

My Independent variable will be – Charge of Ions

My Dependent variable will be – Moles Deposited, which is directional proportional to current

I will be using a stopwatch to keep the time the same, by using the same electrodes I can keep the surface area and size the same. All test solutions will be from the same big bottle; so its constant. I will measure the quantity in a measuring cylinder and keep the distance apart the same by using a ruler. Temperature, which directly affects v in $I=nAve$, will be kept as close as possible to the same, I will conduct all tests within a week, at room temperature.

Method

Equipment List

- Silver Nitrate
- Copper Chloride
- Iron Chloride
- 500ml Beaker
- 250ml Beaker
- 100ml Beaker
- Stopwatch
- Copper Electrodes
- Electrode Clamp
- Wires
- Crocodile Clips
- Variable Resistor
- Ammeter
- Power Pack 1V – 13V
- Three Decimal Place Scales
- Measuring Cylinder
- Distilled Water

Safety Precautions

I wore safety glasses to minimise damage to my eyes. I also mopped up any spillages immediately to decrease

the risk of slips and falls. As well as this, I washed my hands thoroughly after each lesson to stop harmful chemicals damaging my skin. Finally, I kept the classroom well ventilated to try and minimise the levels of chlorine gas to lower the risk of irritation to the skin, eyes, and respiratory system.

Method

I chose 3 different electrolytes in to compare and analyse in my experiment. I have chosen Silver Nitrate, Copper Chloride, and Iron Chloride solutions. I have chosen these three because they are easy for me to get hold of, and they have +1, +2 and +3 positive ions.

I will set up the circuit in Figure 1, I will then set the power pack at 8V. I will fill a 100ml beaker with 60ml of electrolyte, measured in a measuring cylinder precise to $\pm 0.5\text{ml}$. Next place the 2 copper electrodes in the Electrode clamp. Then I will turn the power on for 60 seconds. After that time, I will disconnect the power pack, and carefully remove the electrodes from

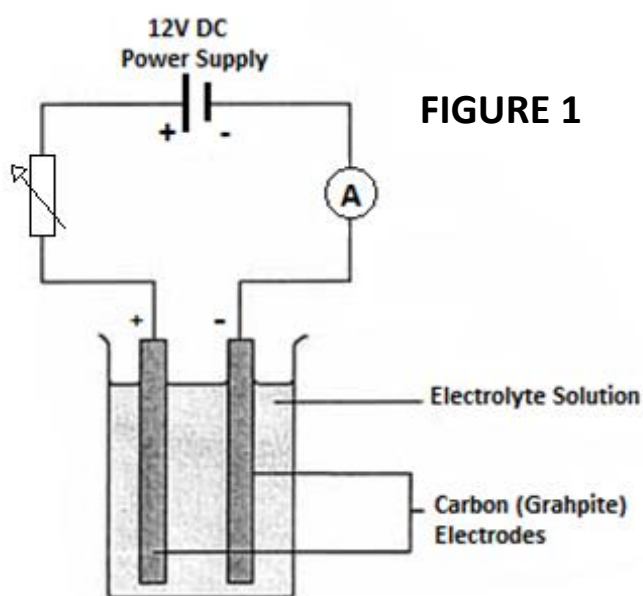


FIGURE 1

the electrolyte. I will weigh the electrodes on the 3dp scale, which is precise to $\pm 0.001\text{g}$, and record the difference. I will do this for +1, +2 and +3 ions, and repeat 3 times. After that, I will find an average for the three attempts, and plot on a Ion Charge vs Moles deposited graph.

Preliminary Tests

For my preliminary experiment, I wanted to use 0.1mol of CuCl. The only solution available was 0.5mol, so I mixed 400ml distilled water with 100ml 0.5mol CuCl to produce 0.1mol Copper Chloride.

60ml 0.1mol Copper Chloride Electrolysed at 8V

Initial Current (A)	Time (S)	Cathode Weight (g)		Anode Weight (g)	
		Before	After	Before	After
0.132	30	0.556	0.591	0.613	0.604

Cathode Difference = 0.035

Anode Difference = - 0.007

I found a noticeable difference in the cathode substantial enough to proceed with the experiment. Although, the anode lost weight which is a big problem. This means we will have to change electrodes because copper atoms were becoming positively charged, and then travelling to the cathode. This makes the test unfair, so I am changing my method to incorporate Graphite electrodes because they are very un-reactive.

I will also up the voltage to 12V and the quantity of electrolyte to 100ml. This relates to $I=nAve$, by increasing n (number of ions) and v (velocity of ions), I can increase the current, thus increasing the amount of moles deposited, and widen my range of results to make the test fairer and more accurate.

I also experienced problems when removing the electrode from the electrolyte, bits of the deposited copper fell off which could largely affect my results. Hopefully this will improve with the use of graphite electrodes.

Finally, in the experiment, we will also have to keep the distance in between the electrodes constant to keep the test fair.

Table of Results (Mass Changes)

Silver Nitrate AgNO_3 (+1)

<u>Initial Amps(A)</u>	<u>Cathode Weight Before (g)</u>	<u>Cathode Weight After (g)</u>	<u>Mass Change (g)</u>
0.65	3.473	3.555	0.082*
0.65	3.478	3.511	0.033
0.55	3.477	3.509	0.032

Copper Chloride CuCl (+2)

<u>Initial Amps(A)</u>	<u>Cathode Weight Before (g)</u>	<u>Cathode Weight After (g)</u>	<u>Mass Change (g)</u>
0.65	3.670	3.740	0.070*
0.65	3.680	3.720	0.040
0.65	3.690	3.720	0.030

Iron Chloride FeCl (+3)

<u>Initial Amps(A)</u>	<u>Cathode Weight Before (g)</u>	<u>Cathode Weight After (g)</u>	<u>Mass Change (g)</u>
0.65	3.457	3.567	0.110
0.65	3.560	3.670	0.110
0.65	3.652	3.758	0.106

* -Taken as Outlier

Table of Results (Moles Deposited)

<u>Solution</u>	<u>Test No.</u>	<u>Mass change recorded (g)</u>	<u>Atomic mass</u>	<u>Moles Deposited (6dp)</u>		
Silver Nitrate AgNO ₃ (+1)	1)	0.082*	/ 108	=		0.000759
	2)	0.033	/ 108	=		0.000306
	3)	0.032	/ 108	=		0.000296
Copper Chloride CuCl (+2)	1)	0.070*	/ 63.5	=		0.001102
	2)	0.040	/ 63.5	=		0.000630
	3)	0.030	/ 63.5	=		0.000472
Iron Chloride FeCl (+3)	1)	0.110	/ 56	=		0.001964
	2)	0.110	/ 56	=		0.001964
	3)	0.106	/ 56	=		0.001893

* - Taken as Outlier

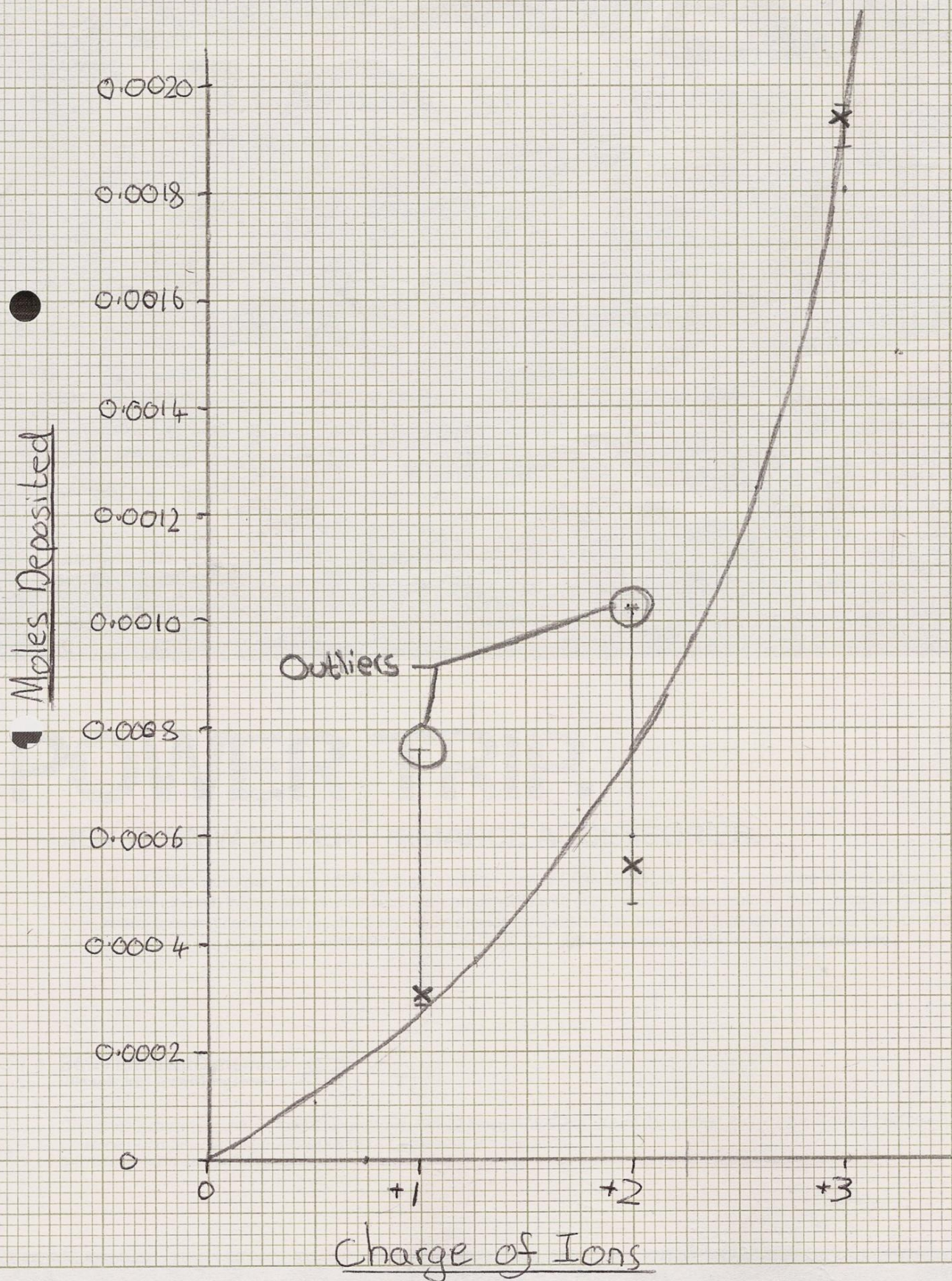
Average Moles Deposited (6dp) :

Silver Nitrate (+1 Ion Charge) = **0.000301**

Copper Chloride (+2 Ion Charge) = **0.000551**

Iron Chloride (+3 Ion Charge) = **0.001940**

Moles Deposited vs Ion Charge



Analysis

My graph shows, with an increase of ion charge, we see an increase in moles deposited. My graph shows an exponential increase, but it's not an exponential graph as it would never be able to go through zero; but no ion charge = no moles deposited.

I am confident of the trend because of scientific evidence and my prediction, although I am not confident of the shape because I have large error bars.

By increasing "e" in $I = nAve$, I have increased the moles deposited.

Gradient Analysis

$$+1 \quad \frac{dy}{dx} = \frac{0.0005}{1.5} = 0.00033 \text{ (5dp)}$$

$$+2 \quad \frac{dy}{dx} = \frac{0.0015}{2.5} = 0.00060 \text{ (5dp)}$$

$$+3 \quad \frac{dy}{dx} = \frac{0.0018}{0.75} = 0.00240 \text{ (5dp)}$$