7.5 - Electrolysis

Electrolysis and electrolytic cells make up the second branch of electrochemistry.

We can think of these two things as opposite of electrochemical cells.

Electrolytic cells convert electrical energy into chemical energy (the electricity is needed to cause a non spontaneous reaction to occur)

The E° value is negative instead of positive.

Electrolysis of Molten NaCl

Electrolysis can be used to break up ionic compounds into its component elements.

For example,

$2NaCl_{(aq)} \rightarrow 2Na_{(s)} + Cl_{2\,(g)}$

Note that this is a redox reaction. Therefore, the two half reactions can be used to find the total voltage:

		E°
reduction	$2Na^+_{(aq)} + 2e^- \rightarrow Na_{(s)}$	-2.71 V
oxidation	$Cl^-_{(aq)} \rightarrow Cl_{2\,(g)} + 2\,e^-$	-1.36 V

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net voltage required - 4.07 V

This negative voltage tells us that the overall reaction will **not** be spontaneous and it will take 4.07 V to cause the reaction to occur.

7.5 - Electrolysis

Some key differences between electrolytic and electrochemical set-ups:

1. There is no salt bridge separating the two half reactions.
2. A source of current is **needed**.
3. The anode is positively charged and the cathode is negatively charged, but the anode is still the site of **oxidation** and the cathode is still the site of **reduction**.

Here is a diagram, with explanations, of an electrolytic cell.

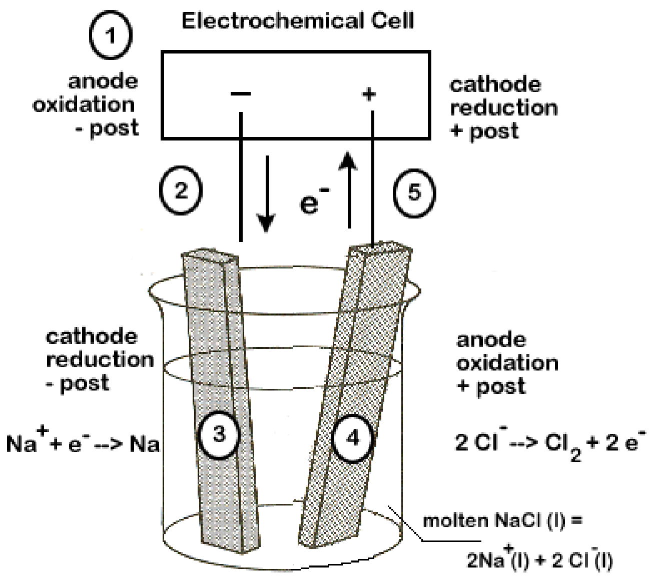
1. Electrons are "produced" in the battery at the anode, the site of oxidation.

2. The electrons leave the electrochemical cell through the external circuit.

3. These negative electrons create a negative electrode in the electrolytic cell which attracts the positive Na^+ ions in the electrolyte. Na^+ ions combine with the free electrons and become reduced ($2Na^+ + 2e^- \rightarrow Na$)

4. Meanwhile the negative Cl^- become attracted to the positive electrode of the electrolytic cell. At this electrode chlorine is oxidized, releasing electrons ($Cl^- \rightarrow Cl_2 + 2\ e^-$

5. These electrons travel through the external circuit, returning to the electrochemical cell.



Electrolytic Cell

$NaCl\ (l) \rightarrow 2\ Na(l) + Cl_2(g)$

Electroplating

https://www.youtube.com/watch?v=FnJ0V7B7nKo

In electroplating, we will see that the cathodes not only **just** carry a charge, but they actively participate in the reaction.

Electroplating is when a thin layer of a desired metal is used to coat (or **plate**) another object. The purpose is to protect against corrosion or improve appearance.

For example, forks made from inexpensive metal are often coated with silver.

The requirements for electroplating:

1. An electrolytic solution which contains ions of the plating metal. $AgNO_3$ will produce sufficient Ag^+ ions.
2. A source of current (a battery).
3. Two electrodes. The first is the object we are plating (the fork), while the second must be the plating metal (silver).

Here are the half reactions:

$Ag^+ + e^- \rightarrow Ag$	cathode	reduction
$Ag \rightarrow Ag^+ + e^-$	anode	oxidation

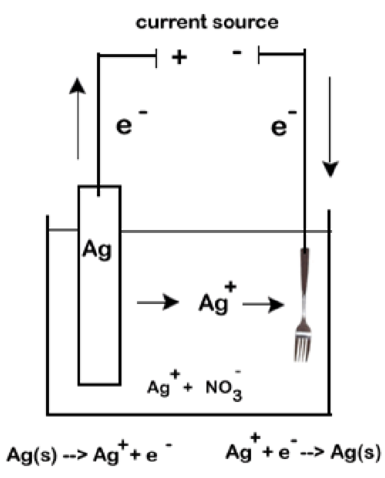
Note - there is only one metal involved.

The Ag^+ that will be deposited on the fork as pure silver once it undergoes reduction will come from the electrolytic solution.

This causes the solution to become negatively charged because there is now more NO_3^- present than Ag^+. The silver bar then undergoes oxidation and replaces the Ag^+ that has been removed.

The flow of electrons goes from the anode of the battery, through the external circuit, and into the cathode of the electrolytic cell.

Electrons are produced by the oxidation of silver in the anode. These flow up through the external circuit into the cathode of the battery.



7.5 - Electrolysis

Ex 1) First, determine the cell reaction for the electrolysis of NaI(aq) using inert electrodes. Then, calculate the minimum voltage that must be applied in order for this reaction to occur.

Ex 2) An iron nail is suspended in a solution of $CuSO_{4\,(aq)}$ by an inert platinum wire. An external power source if connected so the nail becomes the cathode and a copper electrode becomes the anode. If sufficient voltage is applied, what anode and cathode half-reactions occur in this cell?

7.5 Electrolysis Assignment

1. What is the difference between an anode and a cathode in a electrolytic vs. an electrochemical cell?

2. Predict the anode and cathode half reactions during the electrolysis of a 1.0M solution of $KI_{(aq)}$. Assume that inert electrodes are used.

3. Draw a labelled diagram of an apparatus that would plate a nickel-plated knife with silver.

7.5 - Electrolysis

4. Will silver metal in a solution of chloride ions produce silver ions and chlorine gas? Explain.

5. Can a 1M solution of iron (III) sulfate be stored in a container of nickel metal? Explain.

6. How many grams of hydrogen gas would be produced from the oxidation of 5.00g of magnesium metal?