Unit 16: Electrochemistry

Electrolytic Cells

This worksheet will cover electrolytic cells and their differences from galvanic cells. It will discuss the components, half-reactions, and mathematics of electrolytic cells. As you progress through the worksheet, you will develop the skills necessary to distinguish characteristics of different electrochemical cells (galvanic vs. electrolytic), write half-reactions for a common electrolytic cell, and utilize the relationship of current, time, Faraday's constant, and charge.

Practice Problems:

1. What distinguishes electrolytic cells from galvanic/voltaic cells? Draw an example of an electrolytic cell and an example of a galvanic/voltaic cell.

2. Electrons flow from anode/cathode (circle one) to anode/cathode (circle one) in electrolytic cells.

3. Write the half-reactions for the molten NaCl redox reaction (2 NaCl (l) → 2 Na (s) + Cl_2 (g)). Identify the elements being oxidized and reduced.

4. Determine the minimum voltage necessary for the molten NaCl redox reaction to proceed using Standard Reduction Potentials.

$Na^+ + e^- \rightarrow Na$ (s)	– 2.71 V
$Cl_2 + 2e^- \rightarrow 2\ Cl^-$	+ 1.36 V

5. What mass of iron can be produced when a current of 45.2 amps is passed through molten $FeCl_3$ for 1.50 hours?

ANSWER KEY

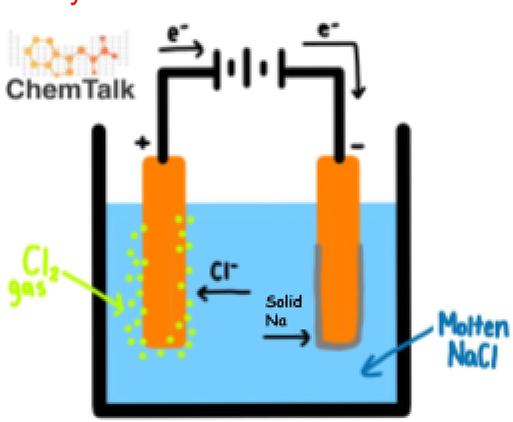
Electrolytic Cells

1. What distinguishes electrolytic cells from galvanic/voltaic cells? Draw an example of an electrolytic cell and an example of a galvanic/voltaic cell.

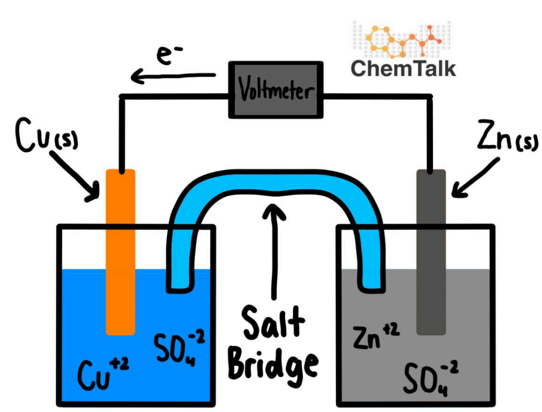
Electrolytic and galvanic cells are both electrochemical cells that involve redox reactions but operate in different ways. These are the main distinctions between them:

- Spontaneity
  - Spontaneous chemical reactions occur in galvanic cells, whereas non-spontaneous reactions are driven by an external source of electrical energy in electrochemical cells.
- Cell components
  - Galvanic cells consist of two separate half-cells connected by a salt bridge. Each half-cell contains an electrode in an electrolyte solution.
  - Electrolytic cells typically consist of one container with an anode and a cathode immersed in an electrolyte.
- Energy flow
  - In galvanic cells, chemical energy is converted to electrical energy, and the cell produces a spontaneous electric current.
  - In electrolytic cells, electrical energy is supplied to drive a nonspontaneous reaction.
- Anode and cathode
  - In galvanic cells, the anode undergoes oxidation, and the cathode undergoes reduction. The anode is negative and the cathode is positive.
  - In electrolytic cells, the anode is positive and the cathode is negative. The anode attracts anions, and the cathode attracts cations. During electrolysis, oxidation occurs at the anode, and reduction occurs at the cathode.
- Cell potential
  - Galvanic cells have a positive cell potential because the redox reaction is spontaneous.
  - Electrolytic cells have a negative cell potential because electrical energy must be supplied to drive the non-spontaneous reaction.

Electrolytic cell:



Galvanic/voltaic cell:



2. Electrons flow from anode/cathode (circle one) to anode/cathode (circle one) in electrolytic cells.

In an electrolytic cell, electrons flow from the anode to the cathode due to the application of an external voltage. The anode, connected to the positive terminal of the battery, undergoes oxidation, releasing electrons into the external circuit (the wire connecting electrodes). These electrons then travel through the external circuit to the cathode, connected to the negative terminal of the battery. At the cathode, a reduction reaction takes place, and electrons are accepted from the external circuit.

3. Write the half-reactions for the molten NaCl redox reaction (2 NaCl (l) → 2 Na (s) + Cl_2 (g)). Identify the elements being oxidized and reduced.

First, let's write the half-reaction for Na. We know that the ionic Na^+ in molten NaCl must gain electrons to convert into its elemental form, Na:

$$2\ Na^+ + 2e^- \rightarrow 2\ Na\ (s)$$

Next, we can write the half-reaction for Cl^-. The ionic Cl^- must lose its electrons to become neutral Cl_2 gas.

$$2\ Cl^- \rightarrow 2e^- + Cl_2$$

Since Na^+ atoms are gaining electrons to become neutral, **Na is being reduced** (LEO GER acronym). Cl^- atoms are losing electrons to become neutral, so **Cl is being oxidized.**

4. Determine the minimum voltage necessary for the molten NaCl redox reaction to proceed using Standard Reduction Potentials.

$Na^+ + e^- \rightarrow Na\ (s)$	– 2.71 V
$Cl_2 + 2e^- \rightarrow 2\ Cl^-$	+ 1.36 V

First, we can determine the $E°_{cell}$ of this electrolytic cell using the SRP's above.

$$E°_{cell} = E°_{reduction,\ cathode} - E°_{reduction,\ anode}$$

We know that the reduction of Na^+ is occurring at the cathode and the oxidation of Cl^- is occurring at the anode, thus:

$$E^\circ_{cell} = E^\circ_{reduction,\ Na+} - E^\circ_{reduction,\ Cl-}$$
$$E^\circ_{cell} = -2.71\ V - 1.36\ V$$
$$E^\circ_{cell} = -4.07\ V$$

Since the E°_{cell} is negative, we know the redox reaction is nonspontaneous, and **we must supply an external voltage of 4.07 V for this reaction to occur.**

5. What mass of iron can be produced when a current of 45.2 amps is passed through molten $FeCl_3$ for 1.50 hours?

First, we can write the half-reaction of Fe in this reaction. We know that the Fe^{+3} ions in molten $FeCl_3$ must be converted to solid Fe.

$$Fe^{+3}\ (l) + 3e^- \rightarrow Fe\ (s)$$

From this, we know that for every mole of Fe, there must be 3 moles of electrons in the reduction reaction.

Next, we can find the charge in Coulombs using *Q = I * t (we must also convert 1.50 hours to seconds by multiplying it by 3600 seconds/hour):*

$$Q = I * t$$
$$Q = 45.2\ \text{amps} * (1.50\ \text{hours} \times 3600\ \text{seconds/hour})$$
$$Q = 244080\ \text{Coulombs}$$

Finally, we can use dimensional analysis and Faraday's constant (96500 Coulombs = 1 mole e^-) to find the deposited mass of Fe.

244080 Coulombs	1 mol e^-	1 mol Fe	55.845 g Fe	= 47.1 g Fe
	96500 Coulombs	3 mol e^-	1 mol Fe	

The mass of iron produced would be 47.1 g.