Electrochemistry

There are two kinds of electrochemical cell:

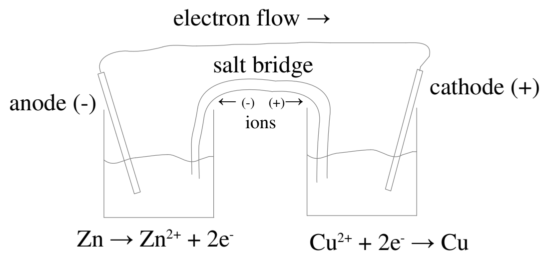
- Galvanic cells (also known as voltaic cells)
  - produce electrical energy from spontaneous chemical reactions
- Electrolytic cells
  - use electrical energy from an external source to cause non-spontaneous chemical reactions

In all cases

- oxidation occurs at the anode
- reduction occurs at the cathode

Remember OIL RIG: Oxidation Is Loss of electrons, Reduction Is Gain of electrons.

An example of a galvanic cell, a zinc-copper cell (a Zn / Zn^{2+} half-cell and a Cu / Cu^{2+} half-cell):



The *salt bridge* (shown labelled) and/or the *electrolyte* (ionic solution):
- complete the circuit
- allows ions to flow

Rechargeable galvanic cells can reverse the electrode reactions by applying an external electrical supply.

- During discharging, the original oxidiser and reducer are used up, producing electricity.
- During recharging, the oxidiser and reducer are regenerated by application of electricity in the opposite direction.

Fuel cells are galvanic cells in which the electrode reactants are available in continuous supply (the electrodes are not consumed) and the electrolyte is not consumed.

Fuel cells compared to other galvanic cells:

Advantages	**Disadvantages**
• higher operating efficiency (and mass-to-power ratio) • consistent operation • electrodes and electrolyte are not consumed minimal maintenance is required	• possibility of contamination ruining the catalyst or electrolyte • the high purity fuels required are costly • many cells require high temperatures • catalysts can be costly • some electrolytes are corrosive

Electrolytic cells can be used in the production of metals.

For example sodium metal is produced by electrolysis of molten NaCl: