GALVANIC CELLS

- Consists of two half-cells connected by a porous boundary (salt bridge) that can produce electricity spontaneously
- Each electrode is in its own electrolyte, forming a half-cell
- The two half-cells are connected by a salt bridge usually filled with sodium nitrate and by an external conductor to make a complete circuit

- The cathode is the electrode where reduction takes place
- The anode is the electrode where the oxidation takes place
    - The substance higher up on E° table will undergo reduction

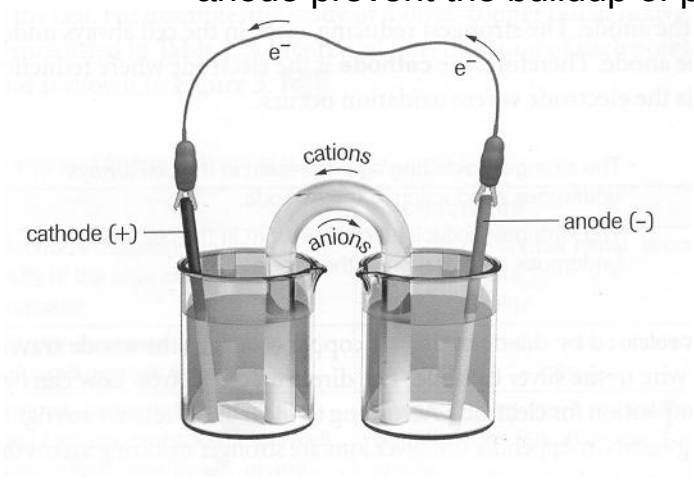
- It can be represented using an abbreviated notation where a single line (|) represents a phase boundary and a double line (||) represents a physical boundary
    - Cathode (+) | electrolyte || electrolyte | anode (-)

- The electrodes are usually metal and the half-reactions take place on the surface of the metal submerged in an electrolyte containing the same metallic ion
- Sometimes an electrode consisting of the same “material” as the electrolyte is impossible to use so an inert electrode is used instead (carbon or platinum)
    - Pt | H^{+1} || Al^{+3} | Al

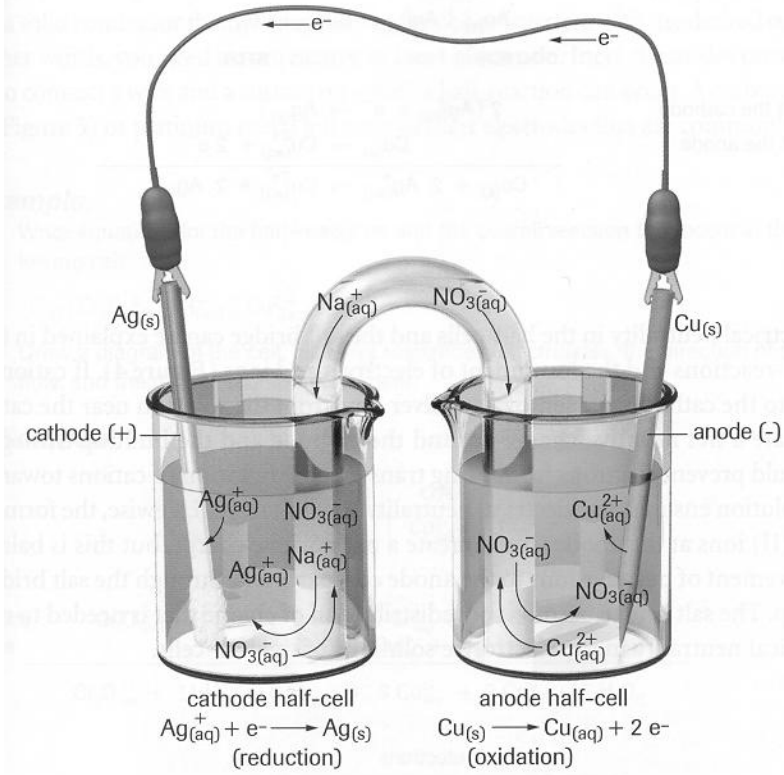
- Electrons released by the oxidation (at the anode) travel through the connecting wire to the cathode where reduction takes place

- The salt bridge ensures electrical neutrality
    - If the positive ions did not move towards the cathode, the loss of positive ions (reduction) that is occurring near the cathode would create a net negative charge
    - The buildup of negative charge would prevent electrons from being transferred
    - Likewise, the production of positive ions (oxidation) at the anode would create a net positive charge

- The migration of negative ions from the salt bridge to the anode prevent the buildup of positive charge



eg) Construct a galvanic cell consisting of 1 M silver nitrate and 1 M copper (II) nitrate solutions, copper and silver metal. Label: external circuit, internal circuit, anode, cathode, charges on the electrodes, site of oxidation, site of reduction, both half-reactions, ion flow and electron flow. Calculate E°cell.



E°cell = E°red – E° oxid
= .8 - .34
= 0.46 V