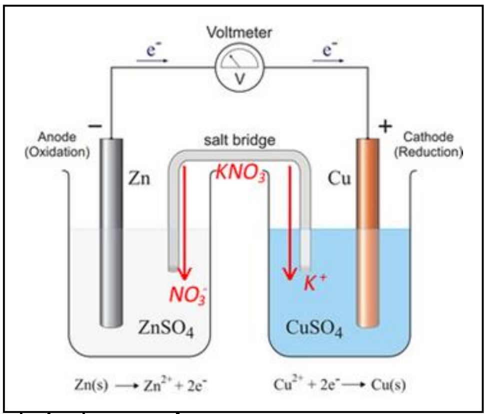
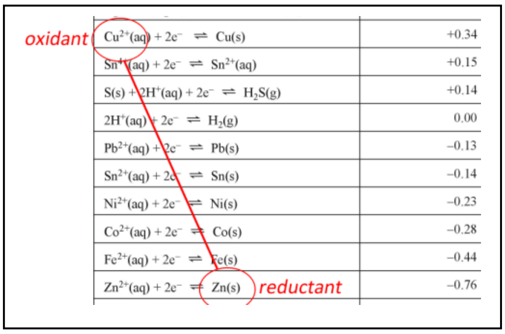
Redox reactions – Galvanic cells
Lesson 4a

Galvanic cells convert chemical potential energy into electrical energy. A redox reaction takes place where the oxidation and reduction half reactions are separated and electrons forced to travel from the oxidation half reaction to the reduction half reaction via an external circuit. Now before we progress any further it should be noted that there are two types of galvanic cells, primary and secondary. Primary cells are non-rechargeable, single use, whereas secondary cells are rechargeable and will be dealt with later under electrolysis.



A few things to note when referring to galvanic cells.

- Energy is converted from chemical to electrical
- The oxidant and reductant are separated into two half-cells
- Oxidation occurs at the electrode known as the anode
- The anode has a negative polarity
- Reduction occurs at the electrode known as the cathode
- The positive electrode is the ***cathode*** while the negative electrode is the ***anode***
- Electrons travel spontaneously from anode to cathode ie. from negative to positive via an external circuit
- The two half-cells are connected by a salt bridge containing an ionic substance, usually KNO_3(aq). Positive ions (K^+) flow from the salt bridge into the half cells containing the cathode while negative ions (NO_3^-) flow from the salt bridge into the half-cell containing the anode.
- In order for a spontaneous reaction to occur the oxidant must be higher than the reductant on the E° table, as shown on the right.
  It is important to note, that having the oxidant higher than the reductant indicates that a reaction will occur, however, the rate of the reaction may be so slow that no reaction is observed.

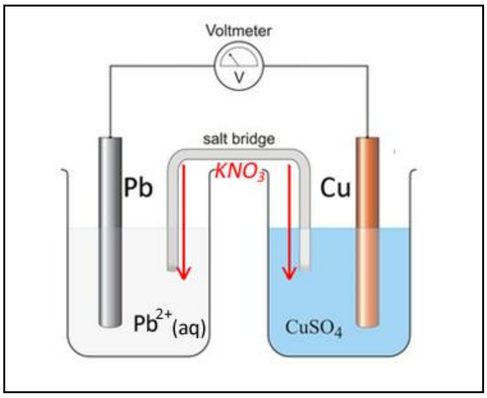


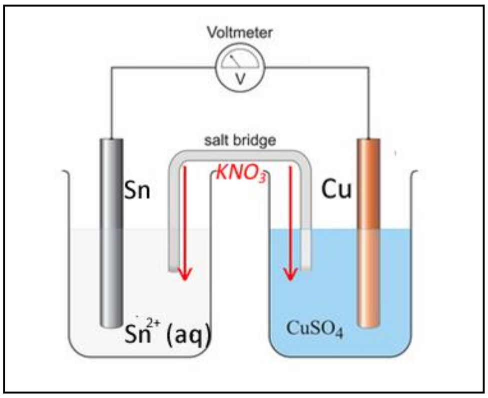
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$		+0.34
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$		+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(g)$		+0.14
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$		0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$		−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$		−0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$		−0.23
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$		−0.28
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$		−0.44
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$		−0.76

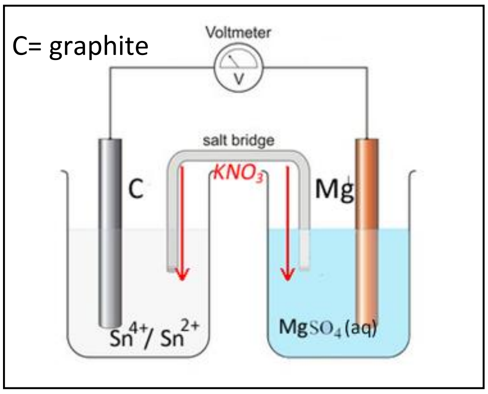
- The half-cell containing the anode has the reductant and its conjugate oxidant, in this case Zn/Zn^{2+} , and the half-cell containing the cathode has the oxidant and its conjugate reductant, in this case Cu^{2+}/Cu
- If one of the species in the half-cell is a metal it is often used as the electrode. In this case copper and zinc are both used as electrodes. However a conducting material, such as graphite, can also be used.
- The voltage produced by the galvanic cell, at standard conditions (25°C, 1 atm pressure and 1.0 M concentration) is given by the formula below
  Cell potential difference = higher half-cell E° – lower half-cell E°
  In this case it is cell potential = +0.34V - -0.76V = 1.10V
  NOTE – ***if the reductant and the oxidant come into direct contact, heat energy is produced rather than electrical energy. The galvanic cell physically isolates the oxidant and reductant and facilitates the production of electrical energy by forcing electrons to travel via an external circuit from the reductant to the oxidant.***

A battery is a galvanic cell that is used to produce electrical energy to power a device. Temperature plays a critical role in the life of a battery. High temperatures can severely reduce the life of a battery by accelerating side reactions which do not produce current. Side reactions may involve the removal of reactant or electrolyte or speeding up the corrosion of electrodes. Low temperatures, however, slow reaction rates and hence decrease the current supplied by the battery

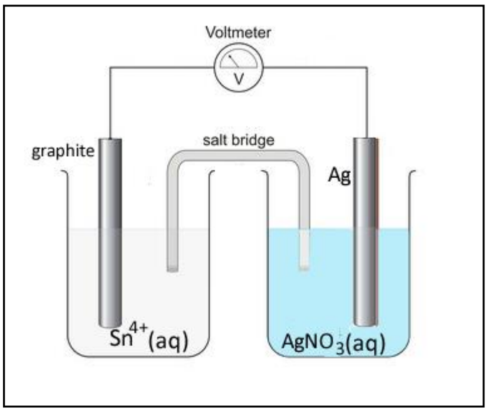
1) Consider the galvanic cells shown on the right.
   For each cell give the cell potential and label the:
- Direction of electron flow
- Anode and cathode
- Direction of K^+ ions flowing from the salt bridge
- Direction of NO_3^- ions flowing from the salt bridge







2) Will a reaction occur in the galvanic cell on the right?
   Explain.



3) The overall redox reaction in a galvanic cell is given below.
$Cu(s) + Cl_2(g) \rightarrow Cu^{2+}(aq) + 2Cl^{-}(aq)$

i. Give the oxidation and reduction half equations
ii. Draw a fully labelled diagram, as in 1) above, of the galvanic cell.
(Consider the shape of a gas electrode)

4) Consider the Galvanic cell shown on the right.
What type of energy is produced by this device if at all.