

Summary

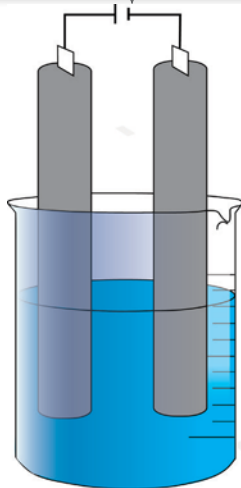
Electrochemical cells:

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- Redox reactions
- Oxidation at the anode
- Reduction at the cathode
- Anions move to the anode.
- Cations move to the cathode.
- External circuit: e^- flow from the anode to cathode.

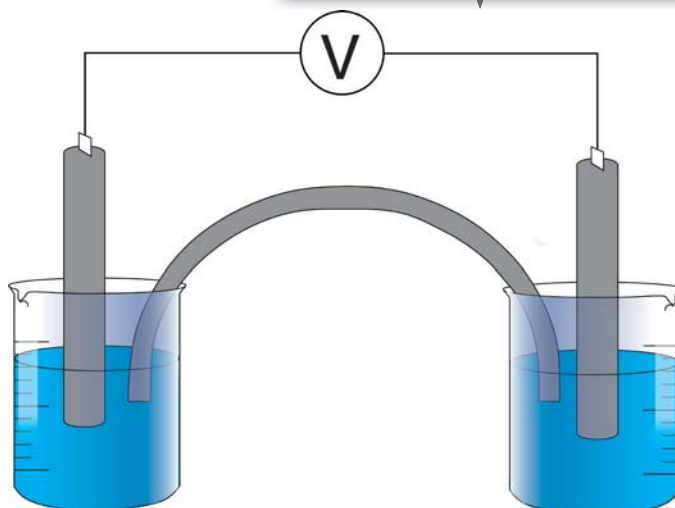
Electrolytic cells

Non-spontaneous redox reaction



Voltaic/galvanic cells

Spontaneous redox reaction



- Weak reducing agent & weak oxidising agent
- Endothermic
- $E_{\text{electrical}} \rightarrow E_{\text{chemical}}$
- $E_{\text{cell}}^0 < 0$
- Anode: oxidation: positive
- Cathode: reduction: negative
- One electrolyte
- Direct current power source
- Electrodes unreactive unless it takes part in the reaction, e.g. electroplating and purification of Cu.

- Strong reducing agent & strong oxidising agent
- Exothermic
- $E_{\text{chemical}} \rightarrow E_{\text{electrical}}$
- $E_{\text{cell}}^0 > 0$
- Anode: oxidation: negative
- Cathode: reduction: positive
- Two separate electrolytes
- Salt bridge
- Electrodes take part in the redox reaction.



Summary

Salt bridge:

- Connects two half-cell electrolytes in a galvanic cell.
- Completes the circuit.
- Ensures electrical neutrality by causing the anions to move to the anode half-cell and by causing the cations to move to the cathode half-cell.
- Contains a solution of spectator ions.
- $\text{KNO}_3(\text{aq})$ is an ideal salt bridge electrolyte.

Cell notation:

- Anode (s) / anode electrolyte (aq) // cathode electrolyte (aq) / cathode (s)

Electrolysis:

- Non-spontaneous electrochemical process during which electrical current flows through an electrolyte between the two electrodes, due to the movement of ions. It leads to the formation of new substances through oxidation and reduction of the ions in the electrolyte.
- In a melted ionic substance:
 - anions move to the anode and are oxidised.
 - cations move to the cathode and are reduced.
- In ionic solutions:
 - Anions move to the anode and compete with (OH^-) in water to be oxidised.
- Preferably Cl^- ; Br^- ; I^- are oxidised, rather than H_2O .
- Preferably H_2O is oxidised, rather than PO_4^{3-} ; SO_4^{2-} ; CO_3^{2-} ; NO^- to form $\text{O}_2(\text{g})$.
 - Cations move to the cathode and compete with the (H^+) in water to be reduced.
- Ag^+ , Cu^{2+} , Sn^{2+} , Ni^{2+} are preferably reduced rather than H_2O and form the corresponding metals.
- H_2O is preferably reduced rather than Na^+ , Al^{3+} , Mg^{2+} and forms $\text{H}_2(\text{g})$.

Electrolysis in industry:

- Extraction of metals such as aluminium, magnesium, sodium and gold
- The manufacturing of halogens such as chlorine
- The purification (refining) of metals such as copper and zinc
- Manufacturing of NaOH to make soap.
- Manufacturing of hydrogen gas to make ammonia.

Standard conditions:

- Electrolyte concentration: $1 \text{ mol}\cdot\text{dm}^{-3}$
- Temperature: 25°C
- Pressure (if a gas is present): $101,3 \text{ kPa}$ (1 atmosphere)

Standard hydrogen electrode:

- Platinum (Pt) electrode in an electrolyte with an $[\text{H}^+(\text{aq})] = 1 \text{ mol}\cdot\text{dm}^{-3}$ and $\text{H}_2(\text{g})$ at $101,3 \text{ kPa}$
- Standard temperature of 25°C maintained.
- Reference electrode
- Standard reduction potential of $0,00 \text{ V}$



Summary

Equilibrium and the current through the galvanic cell:

- Current is the rate at which charge moves past a point.
- $I = \frac{Q}{\Delta t}$
- If less charge moves past a point per time unit (per second) the current decreases.
- Forward reaction rate gradually decreases since [cations at anode] increases and [cations at cathode] decreases.
- Therefore current gradually decreases.

Equilibrium and the emf through a galvanic cell:

- Forward reaction rate gradually decreases and the reverse reaction rate gradually increases, since [cations at anode] increases and [cations at cathode] decreases.
- Condition of dynamic chemical equilibrium is reached and $\text{emf} = 0,00 \text{ V}$.

Electrical conductors:

- Metals: delocalised electrons move freely between positive atom cores.
- Ionic solution: dissociated ions move freely between molecules of the solvent (e.g. water).
- Melted ionic substance: ions move freely amongst each other.
- Ionic solid: non-conductors, since ions are in fixed positions in a crystal lattice.

Spontaneity of redox reactions:

- Redox reaction between strong reducing agents and strong oxidising agents relative to one another, are spontaneous.
- Redox reaction between weak reducing agents and weak oxidising agents relative to one another, are non-spontaneous.
- A reactive metal (strong oxidising agent) displaces a less reactive metal (weak reducing agent) from a solution of its salt.

Notes