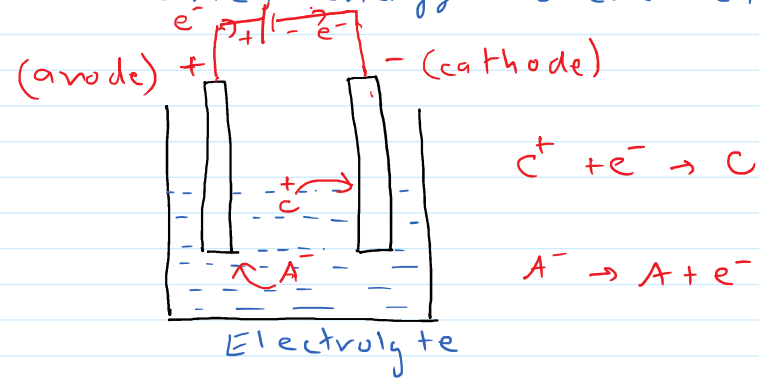
Electrolytic cell

In this cell a non-spontaneous redox reaction is made to happen using external source of electricity. This cell converts electrical energy into chemical energy.



Qualitative analysis

To find out nature of products obtained at anode and cathode

Quantitative analysis

Quantatity / amount of products obtained at anode and cathode. (Faraday's Laws)

Qualitative analysis

Factors affecting cell reactions at electrodes:

i) Nature of electrolyte

For example

NaCl, $MgBr_2$, $ZnCl_2$, $CuSO_4$ all give different products.

ii) State of electrolyte

$NaCl(l) \rightarrow Na^+ + Cl^-$

$NaCl(aq) \rightarrow Na^+ + Cl^- + H^+ + OH^- + H_2O$

Liquid electrolyte and aqueous electrolyte contain

different kind of ions in electrolyte.

iii) Concentration of electrolyte:

If concentration is not mention, then always take high concentration. Products of electrolysis may vary upon concentration in case of certain electrolytes, for example electrolysis of dil. NaCl and conc. NaCl aqueous solution give different products.

IV) Charge density

In some case products also depend upon charge density. For example in case of H_2SO_4.

V) Nature of electrode

a) Inert electrode:

It does not participate in cell reaction.

Eg: Platinum.

b) Active electrode

It participates in cell reaction and hence affect products of electrolysis.

Eg: Ag, Au, Zn.

Rules to analyse electrolysis products:

i) If not mentioned, assume inert electrode.

ii) Always analyse reaction at anode first.

iii) If anode is of metal, then always that metal will oxidise at anode irrespective of presence of other ion.

Case I (Molten salts)

1- NaCl. (using platinum electrodes)

$NaCl(l) \rightarrow Na^+ + Cl^-$

Ions at anode (+)

Cl^-

Ions at cathode (−)

Na^+

Reaction at anode

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

Reaction at cathode

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

Net reaction

$2Na^+ + 2Cl^- \rightarrow Cl_2(g) + Na(s)$

anode cathode.

2. $MgBr_2$ (using platinum electrodes)

$MgBr_2(l) \rightarrow Mg^{2+} + 2Br^-$

Ions at anode (+)

Br^-

Ions at cathode (−)

Mg^{2+}

Reaction at anode

$2Br^- \rightarrow Br_2 + 2e^-$

Reaction at cathode

$Mg^{2+} + 2e^- \rightarrow Mg(s)$

Net reaction

$2Br^- + Mg^{2+} \rightarrow Br_2(g) + Mg(s)$

(anode) (cathode)