1.11 ELECTRODE POTENTIALS AND ELECTROCHEMICAL CELLS

Electrochemical Cell

Circuits with two different metals dipped in salt solutions of their own ions and connected with a wire

- Salt bridge connects the solutions
  - Allows ions to flow through to balance the charges
  - Filter paper soaked in a chemical which will not interfere with the reaction, e.g. KNO_3
- Electrons flow from the more reactive metal to the less reactive metal
  - Redox process
- Reversible reactions take place at each electrode
  - Half equations
    - Combined and balance to form overall equation
    - Always written with the reduction reaction as the forward reaction, e.g. $Zn^{2+}_{(aq)} + 2e^- \rightleftharpoons Zn_{(s)}$
  - When two half cells are joined to make a cell, the metal with the more negative potential is oxidised

Cell potential (EMF or E_{cell})

- Voltage between 2 cells
  - Measured against the standard hydrogen electrode (SHE)
    - SHE uses inert platinum electrode
    - Standard conditions
      - 298 K, 100 kPa and all solutions of concentration 1.00 mol dm^{-3}

Conventional representation of cells

- Single line separates the different phases, double line indicates salt bridge
- reduced form | oxidised form || oxidised form | reduced form
  - $Zn_{(s)}|Zn^{2+}_{(aq)}||Cu^{2+}_{(aq)}|Cu_{(s)}$

Electrochemical Series

- List of standard electrode potentials
  - The more positive the electrode potential, the more likely the metal will be reduced
    - The more negative potential is always subtracted from the more positive potential
- Use values to calculate standard cell potential
  - $E^{\ominus}cell = E^{\ominus}reduced - E^{\ominus}oxidised$

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1.11 ELECTRODE POTENTIALS AND ELECTROCHEMICAL CELLS: COMMERCIAL APPLICATIONS

Non-Rechargeable Batteries

The reactions are difficult or impossible to reverse by supplying a current

Fuel Cells

A cell which continuously produces a voltage when supplied with a fuel and oxygen

Alkaline hydrogen-oxygen fuel cell

Water is the only waste product

Electrode reactions:

$2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$

$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

Do not need to be recharged

Rechargeable Batteries

Recharged by supplying a current which forces electrons to flow in the opposite direction, reversing the reactions

Lithium cell

Electrode reactions:

$Li^+ + CoO_2 + e^- \rightarrow Li + [CoO_2]^-$

$Li \rightarrow Li^+ + e^-$

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