Edexcel International Chemistry A-level

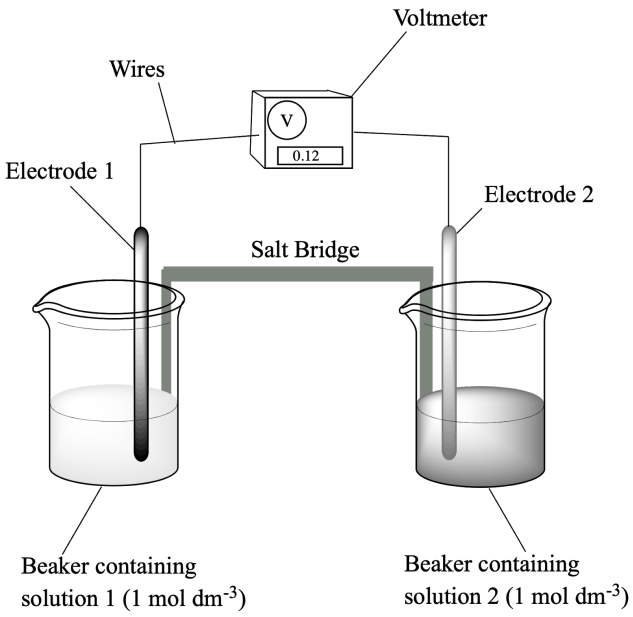
Practical 12

Investigating Electrochemical Cells

Method

1. For each half cell, clean the 2 strips of metal (if you have an ion/ion half cell) with sandpaper so there is surface area exposed for the reactions to occur.
2. Put each piece of metal into 100 cm^3 beaker containing 50 cm^3 of a solution which contains the metal's ions.
3. To create a salt bridge soak a strip of filter paper in KNO_3. Place the salt bridge so its ends are dipped into the 2 beakers with the different metals (see diagram below).
    - Used to complete the circuit by allowing movement of ions and balancing charges. Use of K^+ and NO_3^- minimises the chance of precipitates forming (soluble salts).
4. Use wires and crocodile clips to connect each piece of metal to a voltmeter and read and record the electrode potential.
    - Use a high resistance voltmeter to find the value of the E_{cell}.
5. Repeat steps 1-5 with other pairs of metals.

Diagram



Key Points

- Standard conditions: **1 mol dm^{-3}** solutions, gases at **100 kPa**, **298 K**.
- If the value is negative, the electrodes are the wrong way round.
- Remember, for a thermodynamically feasible process, EMF has to be **positive**.
- **EMF = $E_R - E_L$**, where R - right hand cell, where the reduction occurs, and E_L - left hand cell, where oxidation occurs.

Safety

- Some solutions are **too dangerous** to use at 1 mol dm^{-3} (e.g. silver nitrate which is highly oxidising).
- Zinc sulphate and iron (II) sulphate are **harmful to the environment** therefore have to be disposed of safely.

- Electrodes must be made from an **inert** substance and cannot be made from a metal that reacts with water (e.g. Mg).
- Differences between experimental values and theoretical values usually stem from the **conditions** not being standard.