Combining half-cells

In the cell shown, copper has a more positive $E^\circ$ value (+0.34V) than zinc (-0.76V).

- the zinc metal is more reactive
- it dissolves to give ions
  \[ \text{Zn}_\text{(s)} \rightleftharpoons \text{Zn}^{2+\text{(aq)}} + 2e^- \]
- the electrons produced go round the external circuit to the copper
- electrons are picked up by copper ions and copper is deposited
  \[ \text{Cu}^{2+\text{(aq)}} + 2e^- \rightleftharpoons \text{Cu}_\text{(s)} \]
- The voltage of the cell is 1.10V.

Cell diagrams

These give a diagrammatic representation of what is happening in a cell.

- Place the cell with the more positive $E^\circ$ value on the RHS of the diagram.
- Drawing it out as shown indicates that ...
  - the cell reaction goes from left to right
  - the electrons go round the external circuit from left to right
  - the cell voltage is $E^\circ$(RHS) - $E^\circ$(LHS). In this way it must be positive
  - oxidation takes place at the anode and reduction at the cathode

Conclusion

The reaction(s) will proceed from left to right

\[ \text{OXidation} \quad \text{Zn}_\text{(s)} \rightleftharpoons \text{Zn}^{2+\text{(aq)}} + 2e^- \quad \text{at the ANODE} \]
\[ \text{REDuction} \quad \text{Cu}^{2+\text{(aq)}} + 2e^- \rightleftharpoons \text{Cu}_\text{(s)} \quad \text{at the CATHODE} \]

Electrons

Go from the anode to the cathode via the external circuit

Cell reaction

\[ \text{Zn}_\text{(s)} + \text{Cu}^{2+\text{(aq)}} \rightleftharpoons \text{Zn}^{2+\text{(aq)}} + \text{Cu}_\text{(s)} \]