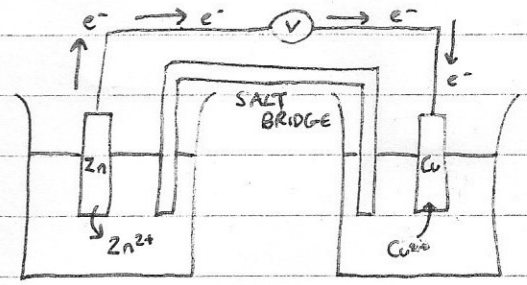
- Electrochemical cells can be made from two different metals dipped in salt solutions of their own ions connected by a wire.

- One reaction is an oxidation reaction, one a reduction. (Redox).

①



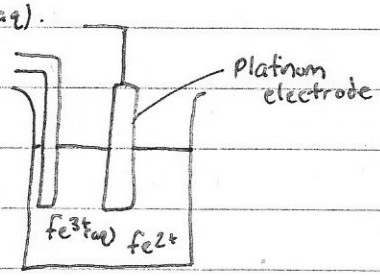
- Salt bridge made from filter paper soaked in KNO_3 (aq). Allows ions to flow through and balances out charges.

- In the above cell, zinc loses electrons more easily than copper, so the zinc electrode is oxidised to form Zn^{2+} and releases the electrons into the circuit. The same number of electrons are taken from the circuit, reducing Cu^{2+} to copper atoms.

- The voltage between the two half cells is called the cell potential or E.M.F, Ecell.

- Some times the half cells can be between two aqueous ion solutions of the same element such as $Fe^{2+}_{(aq)} / Fe^{3+}_{(aq)}$.

②



In the zinc/copper cell the reactions at the electrodes are: $Zn^{2+} + 2e^- \rightleftharpoons Zn$

$Cu^{2+} + 2e^- \rightleftharpoons Cu$

These are reversible and depending on how negative the electrode potential is, tells us how easily that metal is oxidised.

$Zn^{2+}_{(aq)} / Zn_{(s)}$	-0.76
$Cu^{2+}_{(aq)} / Cu_{(s)}$	$+0.34$

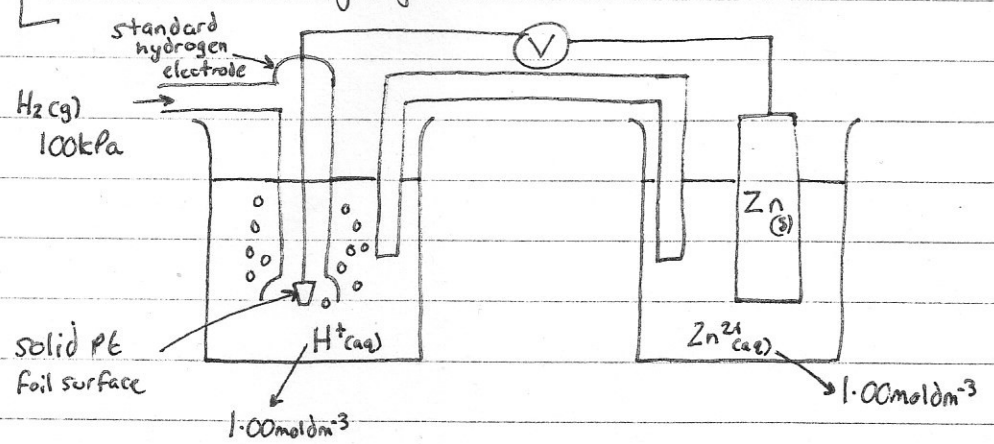
More negative means more likely to release electrons!

$\therefore$ $Zn^{2+} + 2e^- \rightleftharpoons Zn$ Zn is oxidised and Copper is reduced!

$Cu^{2+} + 2e^- \rightleftharpoons Cu$

Electrode potentials are impossible to obtain with only one half cell, and are therefore measured against each other. Science convention has decided to measure all electrode potentials against the standard hydrogen electrode.

> The standard electrode potential of a half cell is the voltage measured under standard conditions when the half cell is connected to a standard hydrogen electrode



As this cell has a solid Pt electrode it should be shown in the cell diagram even if it isn't strictly taking part:

$$Pt \,|\, H_{2(g)} \,|\, H^+_{(aq)} \,||\, Zn^{2+}_{(aq)} \,|\, Zn_{(s)}$$

The standard hydrogen electrode is always shown on the <u>left</u>. By definition the standard hydrogen electrode has an electrode potential of 0.00V.

$\therefore \quad E^\ominus_{cell} = E^\ominus_{right} - E^\ominus_{left}$

$E^\ominus_{cell} = -0.76 - (0.00V)$

$E^\ominus_{cell} = -0.76V$

The negative sign tell us which way the electrons flow.

Drawing electrochemical cells follows a particular method.

① The half cell with the more negative potential goes on the left.
② The oxidised species are in the middle of the diagram.
③ Show your state symbols.

$$\underbrace{Zn_{(s)} | Zn^{2+}_{(aq)}}\ ||\ \underbrace{Cu^{2+}_{(aq)} | Cu_{(s)}}$$

Zn is more negative so is on the left. Zn^{2+} is the oxidised form so goes in the middle.

Cu is on the right, with Cu^{2+} towards the middle.

- To calculate the cell potential: $E^{\ominus}_{cell} = E^{\ominus}_{right\ hand} - E^{\ominus}_{left\ hand}$.

$\therefore\ E^{\ominus}_{cell} = +0.34 - (-0.76)$

$E^{\ominus}_{cell} = +1.10V$

Notice the $\ominus$ symbol? This means standard conditions. As this is an equilibrium, temperature, Pressure and concentration will <u>all</u> affect electrode potentials. Standard conditions are:

Concentration	$1.00 mol dm^{-3}$
temperature	298K
pressure	100kPa