Eg 2)

Oxi: $Na \rightarrow Na^{+} + e^{-}$ (x2)

Red: $Cu^{+2} + 2e^{-} \rightarrow Cu$

Redox: $2Na + Cu^{+2} \rightarrow 2Na^{+} + Cu$

Cell notation: $Na(s)/Na^{+1}(aq)(1\ mol.dm^{-3})// Cu^{+2}(aq)(1\ mol.dm^{-3})/Cu(s)$

**notice the balancing number are not included in the cell notation*

Exercise:

Write the cell notation for the following:

1)

Mg and Al

Oxi: $Mg \rightarrow Mg^{+2} + 2e^{-}$ (x3)

Red: $Al^{+3} + 3e^{-} \rightarrow Al$ (x2)

Redox: $3Mg + 2Al^{+3} \rightarrow 3Mg^{+2} + 2Al$

Cell notation:

2)

Oxi: $Zn(s) \rightarrow Zn^{+2} + 2e^{-}$

Red: $Pb^{+2}(aq) + 2e^{-} \rightarrow Pb(s)$

Redox: $Zn(s) + Pb^{+2}(aq) \rightarrow Zn^{+2}(aq) + Pb(s)$

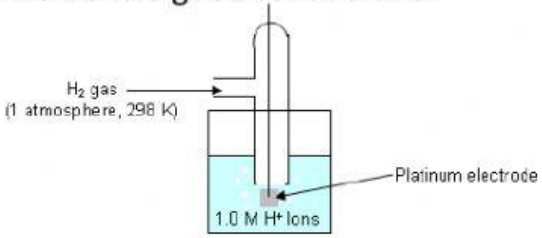
Cell notation:

Hydrogen half cell

Now you can imagine when any gas is involved it – you cannot make an electrode from a gas. Thus what happens is that you have a glass tube, in which you pump the gas into and the solution then has the same element in aqueous solution.

In the case of hydrogen gas- an acid is perfect as an electrolyte, since they all contain H^{+} ions.

The electrode is a small Platinum electrode – since it is inert and a good conductor of electricity.



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You'll also notice the value: 0V next to the hydrogen half reaction on the table of standard reduction potentials.

This is because Hydrogen is the **reference electrode**- meaning that no element by itself has a voltage, but if you connect 2 elements together that creates a potential difference. Hydrogen was thus used as the comparison to create the table.

Thus the standard H electrode is 0,00 V

The $E^{\ominus}$ is positive if electrons flow from the H to the other metal.

And negative if electrons flow from the other metal to Hydrogen.

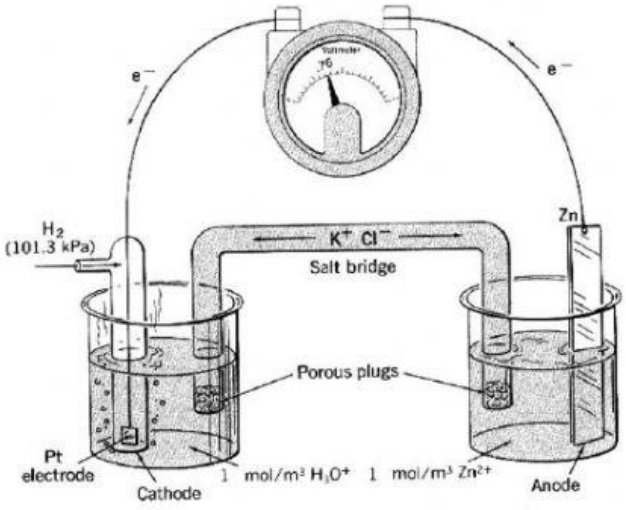
The standard-hydrogen half cell is used to compare other half cells to. It consists of a platinum electrode that is immersed in an acidic solution which contains $[H^+]$ of 1 $mol.dm^{-3}$. Hydrogen gas at a pressure of 1 atmosphere (101.3 kPa) is bubbled over the electrode. The temp is kept constant at 25°C.

Example

Oxidation half reaction: $Zn \rightarrow Zn^{+2} + 2e^-$

Reduction half reaction: $2H^+ + 2e^- \rightarrow H_2$

Redox/net reaction: $Zn + 2H^+ \rightarrow Zn^{+2} + H_2$



The only real difference is the cell notation – you need to include the Pt(s) in the hydrogen half of the notation.

Cell notation/Daniel cell notation:

$Zn(s)/\ Zn^{+2}(aq)(1\ mol.dm^{-3})//\ H^+(aq)(1\ mol.dm^{-3})/\ H_2(g)(1\ atm)$**/Pt(s)**

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Homework:

Pg 184 Q 4 (not 4.8)

Pg 185 Q5 (not 5.8)

Pg 186/187 Q6 and 7(not 6.8 or 7.8)

Pg 199 Q 9 (not 9.5)

Pg 201 Q10

Pg 203 Q8

Pg 204 Q8 (not 8.6)

Pg 208 Q10

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