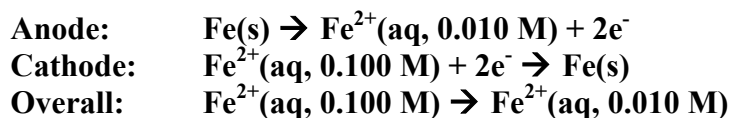


- What is the electrochemical potential of the following cell at 25 °C?



Marks
3

As this is a concentration cell, $E^\circ = 0 \text{ V}$. The cell notation corresponds to the 0.100 M solution being the cathode, where reduction occurs, and the 0.010 M solution being the anode, where oxidation occurs. The two half cells are:



The potential is given by the Nernst equation for this two electron reaction:

$$E = E^\circ - \frac{RT}{nF} \ln Q$$

$$= (0 \text{ V}) - \frac{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298 \text{ K})}{2 \times 96485 \text{ mol}^{-1}} \ln \frac{(0.010)}{(0.100)} = +0.0296 \text{ V}$$

Answer: **+0.0296 V**

- Calculate the mass of aluminium which can be produced with the same quantity of electricity that is used to produce 1.00 kg of copper metal.

2

As the molar mass of Cu is 63.55 g mol^{-1} , 1.00 kg corresponds to:

$$\text{number of moles} = \text{mass} / \text{molar mass}$$

$$= 1.00 \times 10^3 \text{ g} / 63.55 \text{ g mol}^{-1} = 15.7 \text{ mol.}$$

Reduction of Cu^{2+} requires 2 mol of electrons. Hence, the number of electrons requires to produce 15.7 mol is:

$$\text{number of moles of electrons} = 2 \times 15.7 \text{ mol} = 31.5 \text{ mol}$$

Reduction of a mole of Al^{3+} requires 3 mol of electrons. Hence, the number of moles of aluminium produced by 31.5 mol of electrons is:

$$\text{number of moles of aluminium} = 31.5 / 3 \text{ mol} = 10.5 \text{ mol}$$

As the molar mass of aluminium is 26.98 g mol^{-1} , this corresponds to:

$$\text{mass} = \text{number of moles} \times \text{molar mass} = 10.5 \text{ mol} \times 26.98 \text{ g mol}^{-1} = 283 \text{ g.}$$

Answer: **283 g**

ANSWER CONTINUES ON THE NEXT PAGE

- Explain why Na(s) cannot be obtained by the electrolysis of aqueous NaCl solutions.

2

From the table of standard reduction potentials:



Water has a much greater reduction potential than Na^+ and hence is preferentially reduced, even when the overpotential of water is considered.