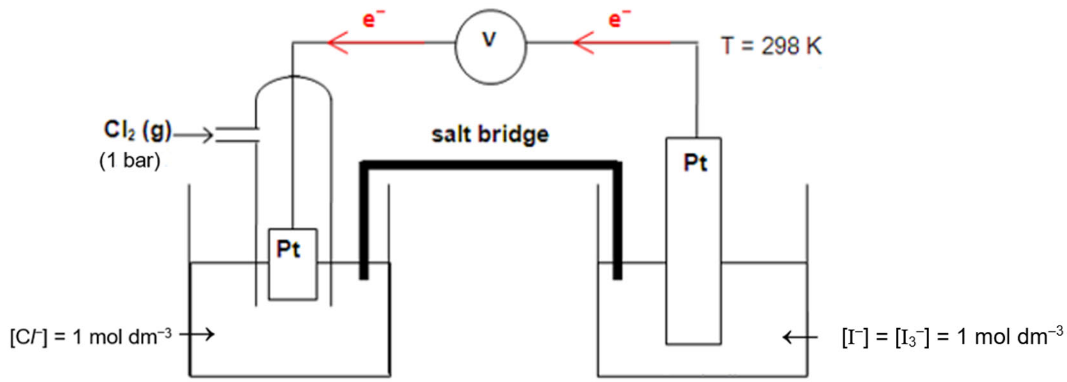
ANSWERS TO SELF-CHECK QUESTIONS

1 **(a)** The ***standard electrode potential***, $E^{\ominus}$, of a half–cell is the electromotive force, measured at **298 K**, between the **half–cell** and the **standard hydrogen electrode**, in which the concentration of any reacting species in solution is **1 mol dm^{-3}** and any gaseous species is at a pressure of **1 bar**.

(b) The **standard cell potential** ($E^{\ominus}_{cell}$) is the **potential difference** between two half–cells under **standard conditions**.

2 **(a)**



Checklist for the general labels to include when drawing the setup:

in each half-cell	connecting 2 half-cells
☐ identity of electrode	☐ external wire with voltmeter
☐ identity and concentration of electrolyte	☐ direction of electron flow
☐ gas jar, identity and partial pressure of gas (if any)	☐ salt bridge
	☐ temperature

Note: Question may also ask for the labelling of anode (–) or cathode (+).

(b) Set up the electrochemical cell as shown above. Read off the E^{θ}_{cell} value on the voltmeter. The greater the E^{θ}_{cell} value, the greater the difference in electrode potentials of the two half-cells. Monitor the direction of the flow of electrons to identify the electrode that is positively charged (the one to which electrons are flowing towards). <u>The electrode that is positively charged</u> contains the <u>stronger oxidising agent</u> which is <u>preferentially reduced</u>.

(c) $Cl_2\ (g) + 2I^-(aq) \rightarrow I_2\ (aq) + 2Cl^-(aq)$

3 **(a)** reduction half cell: $Cl_2\ /\ Cl^-$

(b) $Cl_2 + Co \rightarrow 2Cl^- + Co^{2+}$

(c) $E^{\ominus}_{cell} = +1.36 - (-0.28) = +1.64\ V$