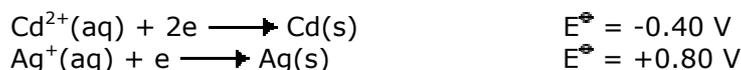


### Exercise 8.64 – Electrode potentials

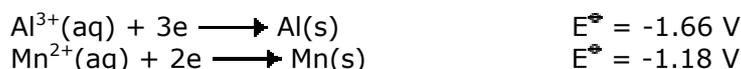
**Q864-01** Given the standard electrode (reduction) potentials:



What would be the  $E^{\ominus}$  for a cadmium-silver cell?

- A. 0.4 V
- B. 0.5 V
- C. 1.2 V
- D. 2.0 V

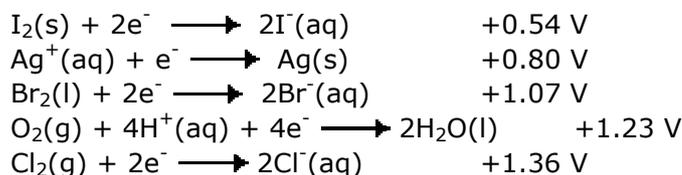
**Q864-02** The standard electrode potentials for Al and Mn are given below:



What is the potential of a cell prepared with these metals in contact with  $1.0 \text{ mol dm}^{-3}$  solutions of their ions?

- A. 0.22 V
- B. 0.48 V
- C. 2.84 V
- D. 3.43 V

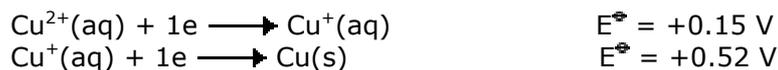
**Q864-03** Use the data below to determine  $E^{\ominus}$  cell for the reaction:



The value is:

- A. -2.45 V
- B. -0.27 V
- C. +0.27 V
- D. +2.45 V

**Q864-04** Consider these standard electrode potentials:

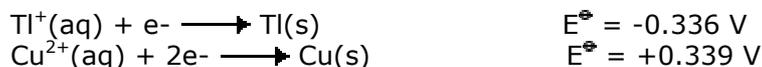


What is the standard cell potential when the two half cells are connected?

- A. -0.67V
- B. -0.37V
- C. +0.37V
- D. +0.67V

### Exercise 8.64 – Electrode potentials

**Q864-05** The standard electrode potentials of two metals are given below. Using this information what are the equation and cell potential for the spontaneous reaction that occurs?

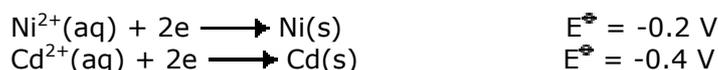


- A.  $\text{Tl}^+(\text{aq}) + \text{Cu}^{2+}(\text{aq}) \longrightarrow \text{Tl}(\text{s}) + \text{Cu}(\text{s})$   $E^\ominus = 0.003 \text{ V}$   
B.  $2\text{Tl}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \longrightarrow 2\text{Tl}^+(\text{aq}) + \text{Cu}(\text{s})$   $E^\ominus = 0.675 \text{ V}$   
C.  $2\text{Tl}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \longrightarrow 2\text{Tl}^+(\text{aq}) + \text{Cu}(\text{s})$   $E^\ominus = 1.011 \text{ V}$   
D.  $\text{Tl}^+(\text{aq}) + \text{Cu}(\text{s}) \longrightarrow 2\text{Tl}(\text{s}) + \text{Cu}^{2+}(\text{aq})$   $E^\ominus = 0.333 \text{ V}$

**Q864-06** Draw a cell diagram for the cell formed by connecting the following standard half cells.



Given:

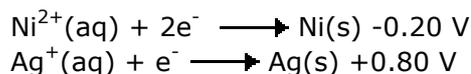


- a) Write an equation for the reaction in each half cell.  
b) Identify the species which is oxidised and the oxidising agent.

**Q864-07** An electrochemical cell is constructed from two half cells connected by a high resistance voltmeter. One half cell contains nickel in a solution of nickel nitrate and the other half cell contains silver in a solution of silver nitrate.

- a) State the conditions that must be applied to the solutions for the measurements taken in the cell to be considered as standard. [2]  
b) Outline how the two half-cells must be connected before any voltage readings can be made. [2]

Assuming that standard conditions apply use the values from the data below to calculate the potential of the cell.



Write the shorthand notation for the cell including state symbols and give the equation for the reaction occurring in the cell. [2]

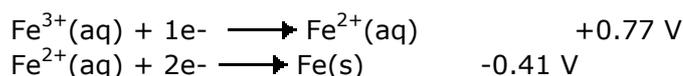
## Exercise 8.64 – Electrode potentials

**Q864-08** A half cell (A) is set up by placing a platinum electrode in a solution containing both  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$  ions at a concentration of  $1 \text{ mol dm}^{-3}$ . This half cell is then connected by means of a salt bridge to another half cell (B) containing an iron electrode in a  $1 \text{ mol dm}^{-3}$  solution of  $\text{Fe}^{2+}$  ions.

a) State the function of the salt bridge [1]

The two electrodes are connected externally.

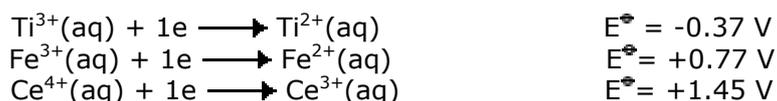
b) Use the data below to determine the cell potential. [2]



Give the redox reactions that occur in each half cell. [2]

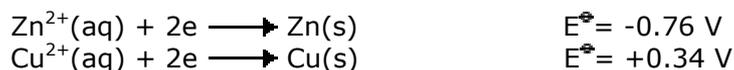
State the direction of electron flow in the external circuit. [1]

**Q864-09** The standard electrode potentials for three electrode systems are given below:



Write an equation including state symbols for the overall reaction with the greatest cell potential. [2]

**Q864-10** Calculate the cell potential of a cell made by connecting standard copper and zinc electrodes, given the following values for standard redox potentials.



a) State the direction of electron flow in the external circuit when the cell produces current.

b) Outline the changes occurring at the electrodes and in the solutions during the process [5]