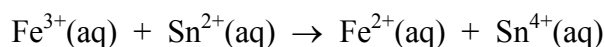
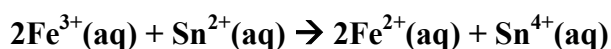


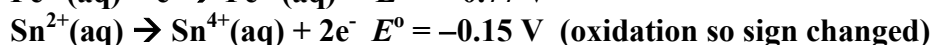
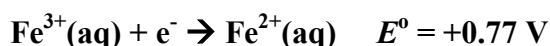
- Calculate the initial cell potential for the following *unbalanced* reaction at 25 °C from the standard electrode potentials. Assume the concentration of all species is initially 1 M.



The balanced equation is:



The relevant half cell reactions and potentials are:



Hence,

$$E^{\circ}_{\text{cell}} = (+0.77 - 0.15) \text{ V} = +0.62 \text{ V}$$

Answer: +0.62 V

Calculate the equilibrium constant, K , for the reaction at 25 °C.

Using $E^{\circ} = \frac{RT}{nF} \ln K$ and $T = (25 + 273) \text{ K} = 298 \text{ K}$,

$$0.62 \text{ V} = \frac{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298 \text{ K})}{2 \times (96485 \text{ C mol}^{-1})} \ln K$$

$$K = 9.4 \times 10^{20}$$

Note that the reaction involves 2 electrons ($n = 2$) as each Sn^{2+} is oxidised to Sn^{4+} and two Fe^{3+} are reduced to two Fe^{2+} .

Answer: 9.4×10^{20}

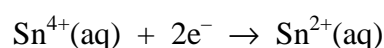
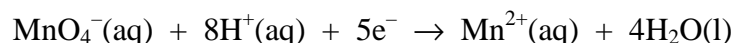
Marks
2

- Give the oxidation number of carbon in each of the following.

CF ₂ Cl ₂ (g)	+4
Na ₂ C ₂ O ₄ (s)	+3
HCO ₃ ⁻ (aq)	+4
C(s)	0

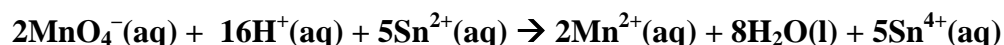
3

- Consider a voltaic cell that uses the following half-reactions:



Write a balanced equation for the overall reaction.

The two cell potentials are +1.51 V (MnO₄⁻/Mn²⁺) and +0.15 V (Sn⁴⁺/Sn²⁺). The least positive (Sn⁴⁺/Sn²⁺) is reversed so that it is the oxidation reaction. Combining the half cells in this way gives, after balancing:



Which species is the oxidising agent?

MnO₄⁻(aq)

Which species is the reducing agent?

Sn²⁺(aq)

Calculate the standard cell potential. (Refer to the table of standard reduction potentials.)

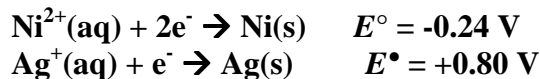
As noted above, the Sn⁴⁺/Sn²⁺ half cell is reversed so that its oxidation potential is -0.15 V. Hence, the standard cell potential is:

$$E_{\text{cell}}^{\circ} = E_{\text{red}}^{\circ} + E_{\text{ox}}^{\circ} = (+1.51) + (-0.15) = +1.38 \text{ V}$$

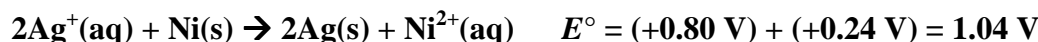
Answer: +1.38 V

- A galvanic cell is made of a Ni^{2+}/Ni half cell with $[\text{Ni}^{2+}] = 1.00 \times 10^{-3} \text{ M}$ and a Ag^+/Ag half cell with $[\text{Ag}^+] = 5.00 \times 10^{-2} \text{ M}$. Calculate the electromotive force of the cell at 25°C .

The reduction potentials for the two half cells are (from the data sheet):



The $\text{Ni}^{2+}(\text{aq}) / \text{Ni}(\text{s})$ cell has the lower electrode potential and it is the one that is reversed. Hence, the cell reaction and the standard cell potential are



At non-standard concentrations, the electrode potential is given by the Nernst equation:

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

For this two electron reduction, $n = 2$ and $Q = \frac{[\text{Ni}^{2+}(\text{aq})]}{[\text{Ag}^+(\text{aq})]^2}$

$[\text{Ni}^{2+}(\text{aq})] = 1.00 \times 10^{-3} \text{ M}$ and $[\text{Ag}^+(\text{aq})] = 5.00 \times 10^{-2} \text{ M}$. Hence, at $T = 25^\circ\text{C} = (25 + 273) \text{ K} = 298 \text{ K}$:

$$E = (1.04 \text{ V}) - \frac{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298 \text{ K})}{(2 \times 96485 \text{ C mol}^{-1})} \ln\left(\frac{1.00 \times 10^{-3}}{(5.00 \times 10^{-2})^2}\right) = 1.06 \text{ V}$$

Answer: **1.06 V**

Calculate the equilibrium constant of the reaction at 25°C .

The equilibrium constant is related to the standard electrode potential:

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

Hence, with $E^\circ = 1.04 \text{ V}$ and $n = 2$:

$$\ln K = E^\circ \times \frac{nF}{RT} = (1.04 \text{ V}) \times \frac{(2 \times 96485 \text{ C mol}^{-1})}{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298 \text{ K})} = 81.0$$

so,

$$K = 1.51 \times 10^{35}$$

Answer: **1.51×10^{35}**

ANSWER CONTINUES ON THE NEXT PAGE

Calculate the standard free energy change of the reaction at 25 °C.

The standard free energy change, ΔG° , is related to the standard electrode potential, E° :

$$\begin{aligned}\Delta G^\circ &= -nFE^\circ \\ &= -(2) \times (96485 \text{ C mol}^{-1}) \times (1.04 \text{ V}) = -201000 \text{ J mol}^{-1} = 201 \text{ kJ mol}^{-1}\end{aligned}$$

Alternatively, the free energy change is related to the equilibrium constant, K :

$$\begin{aligned}\Delta G^\circ &= -RT \ln K \\ &= -(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (298 \text{ K}) \times (1.51 \times 10^{35}) = -201 \text{ kJ mol}^{-1}\end{aligned}$$

Answer: **-201 kJ mol⁻¹**

Indicate whether the reaction is spontaneous or not. Give reasons for your answer.

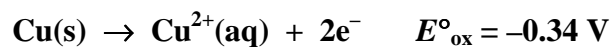
- $E^\circ > 0$ so the reaction is spontaneous.
- $\Delta G^\circ < 0$ so the reaction is spontaneous.
- K is much greater than 1 so the reaction is spontaneous.

Express the overall reaction in the shorthand voltaic cell notation.



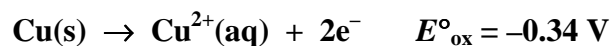
- Explain why copper dissolves in dilute HNO₃, but not in dilute HCl?

Copper does not dissolve in dilute HCl(aq) it is not oxidized to Cu²⁺(aq) under these conditions:



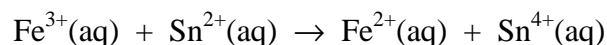
The overall cell potential $E^{\circ} = E^{\circ}_{\text{ox}} + E^{\circ}_{\text{red}} = (-0.34 + 0) = -0.34 \text{ V}$. The cell potential is negative and the reaction is not spontaneous.

Copper does dissolve in dilute HNO₃ due to oxidation by the NO₃⁻ ion:



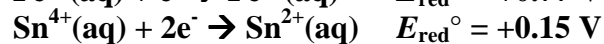
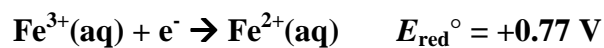
The overall cell potential $E^{\circ} = E^{\circ}_{\text{ox}} + E^{\circ}_{\text{red}} = (-0.34 + 0.96) = +0.62 \text{ V}$. The cell potential is positive and the reaction is spontaneous.

- Consider the following *unbalanced* reaction at 25 °C:

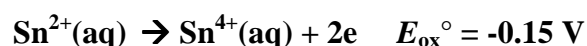


Calculate the standard cell potential.

The two half cells and the standard reduction potentials are:



As the $\text{Sn}^{4+} / \text{Sn}^{2+}$ cell has the least positive reduction potential, it is reversed and becomes the oxidation half cell:



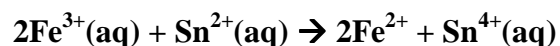
Hence, the standard cell potential is:

$$E^{\circ} = (+0.77 \text{ V}) + (-0.15 \text{ V}) = +0.62 \text{ V}$$

Answer: **+0.62 V**

Calculate the equilibrium constant, K , for the reaction at 25 °C.

Using the half cell reactions above, the balanced equation is:



It involves the transfer of 2 electrons. The equilibrium constant for this 2 electron reaction is therefore given by:

$$E^{\circ} = (RT/nF)\ln K$$

$$\ln K = E^{\circ} \times (nF / RT) = (0.62) \times (2 \times 96485) / (8.314 \times 298) = 48.3$$

$$K = 9.4 \times 10^{20}$$

Answer: **9.4×10^{20}**