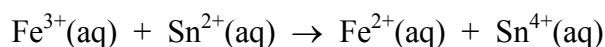
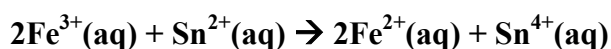


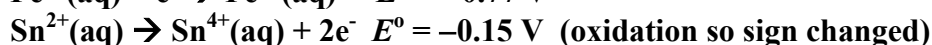
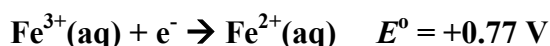
- Calculate the initial cell potential for the following *unbalanced* reaction at 25 °C from the standard electrode potentials. Assume the concentration of all species is initially 1 M.



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  $\text{Sn}^{2+}$  is oxidised to  $\text{Sn}^{4+}$  and two  $\text{Fe}^{3+}$  are reduced to two  $\text{Fe}^{2+}$ .

Answer:  $9.4 \times 10^{20}$