Balanced Redox Reaction

Given the reduction potentials below, determine the balanced redox reaction and the cell potential for a voltaic cell. Calculate the free energy change for the reaction.

$$MnO_4^- + 8\,H^+ + 5\,e^- \rightarrow Mn^{2+} + 4\,H_2O \qquad E^o_{red} = +1.51\,V$$

$$SO_4^{2-} + 2\,H^+ + 2\,e^- \rightarrow SO_3^{2-} + \ H_2O \qquad E^o_{red} = +0.17\,V$$

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Given the reduction potentials below, determine the balanced redox reaction and the cell potential for a voltaic cell. Calculate the free energy change for the reaction.

Solution: Step 1.A voltaic cell should have a positive cell potential. Therefore, we choose the smaller positive value and write that reduction reaction in reverse as an oxidation.

$$MnO_4^- + 8\,H^+ + 5\,e^- \rightarrow Mn^{2+} + 4\,H_2O \qquad E^o_{red} = +1.51\,V$$

$$SO_3^{2-} + \;H_2O \;\rightarrow SO_4^{2-} + 2\,H^+ + 2\,e^- \qquad E^o_{ox} = -0.17\,V$$

Step 2. Then we find the least common factor between the two reactions to eliminate the electron term in both reactions when we sum them.

$$2(MnO_4^- + 8\,H^+ + 5\,e^- \rightarrow Mn^{2+} + 4\,H_2O) \qquad E^o_{red} = +1.51\,V$$

$$5(SO_3^{2-} + \;H_2O \;\rightarrow SO_4^{2-} + 2\,H^+ + 2\,e^-) \qquad E^o_{ox} = -0.17\,V$$

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Given the reduction potentials below, determine the balanced redox reaction and the cell potential for a voltaic cell. Calculate the free energy change for the reaction.

Step 3. Add the reactions. Remember that the cell potential is not affected by the factors used to balance the equation.

The balanced reaction is

$$2\ MnO_4^- + 5\ SO_3^{2-} + 6\ H^+ \rightarrow 2\ Mn^{2+} + 3\ H_2O + 5\ SO_4^{2-}$$

And the cell potential is

$$E^o_{red} + E^o_{ox} = E^o_{cell}$$

$$1.51\ V + (-0.17\ V) = +1.34\ V$$

Step 4. Calculate the free energy

$$\Delta_{rxn}G^o = -nFE^o_{cell} = -(10)(96472)(1.34) = -1{,}290\ kJ/mol$$