From the world of RedOx:

You are given the following schematic representation of a galvanic cell:

\[ \text{Fe} / \text{FeSO}_4 (0.50 \text{ M}) \parallel \text{MnO}_2^{4+} (0.10 \text{ M}), \text{MnO}_2(g), \text{pH} = 1.0 / \text{Pt} \]

NOTE: The FeSO_4 is a completely soluble salt, while MnO_2 is an insoluble nonconductive solid in its pure state.

a) Identify the two electrodes. Which electrode is the cathode, and which is the anode? What are their polarities (+ or -)? For the Fe^3+ / Fe redox couple, Fe is the anode for the MnO_2 / Mn^4+ redox couple, Pt is the electrode.

b) Write the two half reactions for this galvanic cell. Which is the oxidation and which is the reduction?

\[ \text{Fe} \rightarrow \text{Fe}^{3+} + 2e^- \quad \text{oxidation} \]

\[ \text{MnO}_2 + 4H^+ + 2e^- \rightarrow \text{Mn}^{2+} + 2H_2O \quad \text{reduction} \]

c) Write the overall cell reaction. Adding the two above removes the 2e^-.

\[ \text{You get:} \quad \text{Fe} + \text{MnO}_2 + 4H^+ \rightarrow \text{Fe}^{3+} + \text{Mn}^{2+} + 2H_2O \]

d) Calculate the reduction potential of each half cell.

For \( \text{MnO}_2 \) system:
\[ E = E^0_{\text{MnO}_2} + \frac{0.0591}{2} \log \left( \frac{[\text{Mn}^{2+}][\text{H}_2\text{O}]}{[\text{MnO}_2]} \right) \]
\[ = 1.23 + \frac{0.0591}{2} \log \left( \frac{1}{4} \right) \]
\[ = 1.23 - 0.089 = 1.14 \]

For \( \text{Fe} \) system:
\[ E = E^0_{\text{Fe}^{3+}/\text{Fe}^{2+}} - \frac{0.0591}{2} \log \left( \frac{[\text{Fe}^{3+}]}{[\text{Fe}^{2+}]} \right) \]
\[ = -0.440 - \frac{0.0591}{2} (-0.303) \]
\[ = -0.440 + 0.089 = -0.351 \]

e) Calculate the overall cell potential, \( E_{\text{cell}} \)

\[ E_{\text{cell}} = E_R - E_L = 1.14 - (-0.45) = +1.59 \text{V} \]

From the world of Kinetics:
By analyzing kinetic data for the reaction $2A + B \rightarrow 2C$, we discover that the reaction has a first order dependence on A and a second order dependence on B.

a) What is the mathematical form of a rate law that expresses these facts

_Generically_, the rate law $\nu = k [A]^2 [B]^2$

_in terms of A_, $-\frac{d[A]}{dt} = \frac{1}{2} k [A]^2 [B]^2$

_in terms of B_, $-\frac{d[B]}{dt} = k [A]^2 [B]^2$

_in terms of C_, $\frac{d[C]}{dt} = \frac{1}{2} k [A]^2 [B]^2$

b) If we double the initial concentration of B how will the initial rate change?

The rate will change by a factor of $2^2 = 4$

c) If the rate of consumption of A is 2.5 M/s, what is the rate of consumption of B? What is the rate of production of C?

A is being used up twice as fast as B

So if A is consumed 2.5 M/sec, then B will be consumed 1.25 M/sec

The rate of production of C will be equal to the rate of consumption of A, 2.5 M/sec

(it will increase as A decreases at the same rate)

Please write and sign the pledge: