Answer on Question # 69285, Chemistry / General Chemistry

2. A 20.00 mL sample of MnO_4^{-1} is required to titrate 0.2378 g $Na_2C_2O_4$ in an acidic solution. How many ml of this same MnO_4^{-1} are required to titrate a 25.00mL sample of 0.1010 M Fe^{+2} in acidic solution.

$$2MnO_4^{-1}(aq) + 16H^+(aq) + 5{C_2O_4}^{-2}(aq) \rightarrow 2Mn^{+2}(aq) + 8H_2O(l) + 10CO_2(g)$$

(SHOW WORK)

Solution:

1. **Calculate molarity of solution** MnO_4^{-1}:

$$c(1/5KMnO_4) = \frac{1000 * m(Na_2C_2O_4)}{V(1/5KMnO_4) * E(Na_2C_2O_4)}$$

$$c(1/5KMnO_4) = \frac{1000 * 0.2378}{20.00 * 2.5 * 134.00} = 0.035493\ (M)$$

2. **Balance equation for the second reaction:**

$$MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$$

3. **Calculate volume of solution** MnO_4^{-1}:

$$V(MnO_4^-) = \frac{c(Fe^{2+}) \times V(Fe^{2+}) \times f}{c(MnO_4^-)}$$

$$V(MnO_4^-) = \frac{0.1010 \times 25.00 \times 1}{0.035493 \times 5} = 14.23\ (ml)$$

Answer: 0.035493 M; 14.23 ml.