

Experiment 30 Advance Study Assignment: Determination of Iron by Reaction with Permanganate—A Redox Titration

Name _____

Section _____

1. Write the balanced net ionic equation for the reaction between MnO_4^- ion and Fe^{2+} ion in acidic solution.



2. How many moles of Fe^{2+} ion can be oxidized by 1.4×10^{-2} moles MnO_4^- ion in the reaction in Question 1?

1 mole of MnO_4^- reacts with 5 moles of Fe^{2+}
 then 1.4×10^{-2} moles of MnO_4^- reacts with x moles Fe^{2+}

$$x = \frac{1.4 \times 10^{-2} \text{ moles} \times 5 \text{ moles}}{1 \text{ mole}} = 7 \times 10^{-2} \text{ moles } \text{Fe}^{2+}$$

$$7 \times 10^{-2} \text{ moles } \text{Fe}^{2+}$$

3. A solid sample containing some Fe^{2+} ion had a total mass of 0.9791 g. It required 18.2 mL, 0.02034 M KMnO_4 to titrate the Fe^{2+} in the dissolved sample to a pink end point.

$$\text{MnO}_4^- \text{ molarity (M)} = 0.02034 \text{ M}$$

$$\text{volume of } \text{KMnO}_4 = V = 18.2 \text{ mL}$$

- a. How many moles of MnO_4^- ion were required? $= 18.2 \times 10^{-3} \text{ Liters.}$

$$\text{moles of } \text{MnO}_4^- \text{ ion required} = M \times V$$

$$= 0.02034 \frac{\text{moles}}{\text{Liter}} \times 18.2 \times 10^{-3} \text{ Lit} = 0.3701 \times 10^{-3} \text{ moles}$$

$$0.3701 \times 10^{-3} \text{ moles}$$

- b. How many moles of Fe^{2+} were present in the sample?

$$\text{mass of } \text{Fe}^{2+} \text{ ions} = 0.9791 \text{ g.}$$

$$\text{molecular weight of } \text{Fe}^{2+} = 55.8 \text{ g/mol.}$$

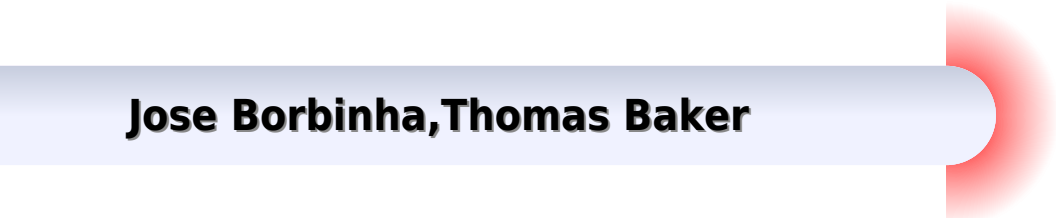
$$\text{moles of } \text{Fe}^{2+} = \frac{\text{mass}}{\text{molecular weight}} = \frac{0.9791 \text{ g}}{55.8 \text{ g/mol}}$$

$$\text{moles of } \text{Fe}^{2+} = 0.01754 \text{ moles.}$$

$$0.01754 \text{ moles}$$

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Jose Borbinha, Thomas Baker



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