

# 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  $\text{MnO}_4^-$  ion and  $\text{Fe}^{2+}$  ion in acidic solution.



2. How many moles of  $\text{Fe}^{2+}$  ion can be oxidized by  $1.4 \times 10^{-2}$  moles  $\text{MnO}_4^-$  ion in the reaction in Question 1?

1 mole of  $\text{MnO}_4^-$  reacts with 5 moles of  $\text{Fe}^{+2}$   
 then  $1.4 \times 10^{-2}$  moles of  $\text{MnO}_4^-$  reacts with x moles  $\text{Fe}^{+2}$

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

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

3. A solid sample containing some  $\text{Fe}^{2+}$  ion had a total mass of 0.9791 g. It required 18.2 mL, 0.02034 M  $\text{KMnO}_4$ , to titrate the  $\text{Fe}^{2+}$  in the dissolved sample to a pink end point.

$\text{MnO}_4^-$  molarity (M) = 0.02034 M  
 volume of  $\text{KMnO}_4^-$  (V) = 18.2 ml

- a. How many moles of  $\text{MnO}_4^-$  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  $\text{Fe}^{2+}$  were present in the sample?

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

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

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

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

$$0.01754 \text{ moles}$$

# Advance Study Assignment Experiment 30 Answers

**JL Elias**



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