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INTERPRETING GRAPHICS

USE WITH SECTION 22.3

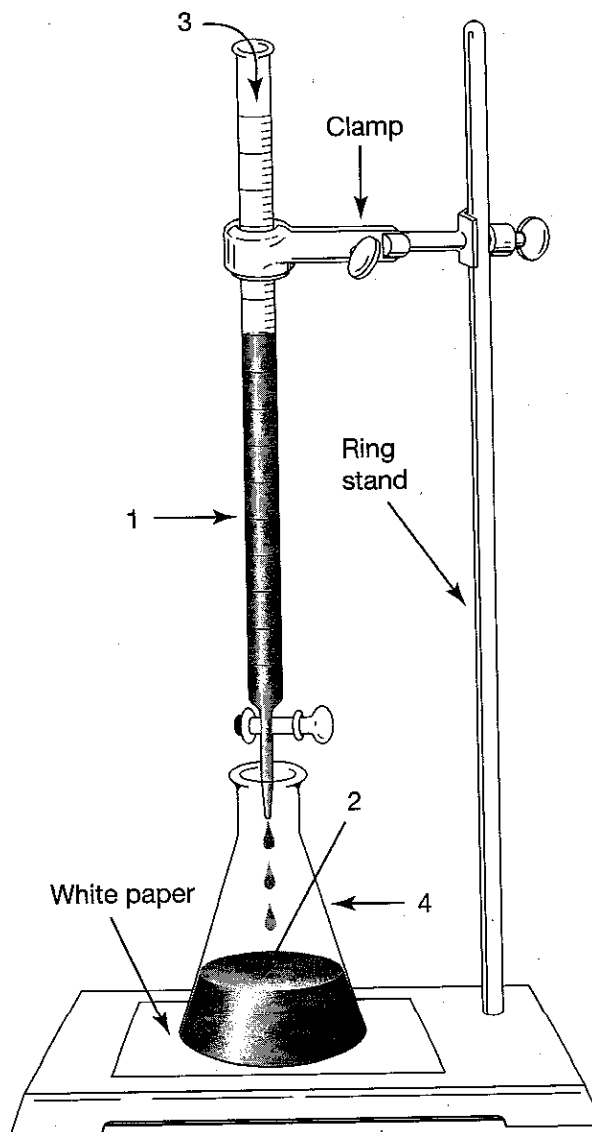
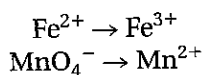


Figure 1 Titration of iron(II) ion (Fe^{2+}) with a standard solution of 0.0200M potassium permanganate (KMnO_4).

To determine the relative amount of iron in a sample of iron ore, a chemist dissolved 2.938 g of the ore in 50.0 mL of dilute sulfuric acid (H_2SO_4) in a reaction flask. The colorless solution was then titrated to the end point with potassium permanganate. The half-reactions for the oxidation and reduction processes that occur during this titration are:



Use the data in Table 1 and what you have learned about oxidation–reduction reactions to answer the following questions.

Table 1 Analysis of an Unknown Iron-Containing Ore

Initial Volume of KMnO_4	48.65 mL
Final Volume of KMnO_4	23.35 mL
Volume of MnO_4^-	
Moles MnO_4^-	
Moles Iron(II), Fe^{2+}	
Mass of Iron	
% of Iron in Ore	

1. Match each component from the following list with the correct number shown in Figure 1. The same number may be used more than once.

- _____ a. oxidizing agent
- _____ b. reducing agent
- _____ c. standard solution of 0.0200M KMnO_4
- _____ d. acidic solution of iron(II) ion, Fe^{2+}
- _____ e. reaction flask
- _____ f. buret

2. Use the half-reaction method to balance the equation for the redox reaction between permanganate ion and iron(II) ion. Write the net ionic equation only.

3. Explain what the *end point* of this particular titration means in terms of the reacting species in solution. How does the chemist recognize the end point when it occurs?

4. Use the stoichiometry of the balanced equation given in your answer to question 2 and the fact that the molar mass of Fe is 55.85 g to complete Table 1 above. Use the space below to show your work.