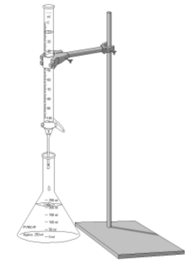
AP Chem Investigation 12 – Redox Titration Lab



Name ___KEY___________________________________ Period _______

1. Check to be sure you have at least 20 mL of $KMnO_4$ in your buret. If not, add some $KMnO_4$ to your buret.
2. Be sure there are no air bubbles in the buret. If there are, stream some $KMnO_4$ out of the buret into a waste beaker until the air bubble comes out. You can dump this solution down the sink with running water.
3. Take an initial reading of $KMnO_4$ from the buret. Record this value in the data table.
4. Measure out exactly 1.00 mL of H_2O_2 using a graduated pipet and add it to a clean flask.
5. Add 10 mL of 3M H_2SO_4 to the flask using a graduated cylinder. This is so that the titration can occur in acidic conditions.
6. Add water to the flask to bring the solution up to about 30-50mL of volume. (This is just to improve your ability to see the color change at the endpoint. Adding water will NOT affect the number of moles of H_2O_2 present in the flask, so it won't affect the volume of titrant required to reach the equivalence point).
7. Titrate the H_2O_2 to the endpoint (until a faint pink color persists). Try not to add even just 1 drop of $KMnO_4$ past the endpoint! Record the volume of titrant
8. Dump your solution down the sink. Leave the leftover $KMnO_4$ in your buret for the next class.
9. Check the label on the $KMnO_4$ bottle and record the molarity of the $KMnO_4$ solution in your data section.

Data Table

Volume of H_2O_2 titrated	1.00 mL
Initial reading of MnO_4^- volume on buret	19.15 mL
Final reading of MnO_4^- volume on buret	37.61 mL
Concentration of MnO_4^- solution on bottle	0.0188M

Questions:

1. Balance the equation below using the half-reaction method. Show all work.

$$MnO_4^-{}_{(aq)} + H_2O_{2(aq)} \rightarrow O_{2(g)} + Mn^{2+}{}_{(aq)}$$

$MnO_4^- \text{-->} Mn^{2+}$

$(5e^- + 8H^+ + MnO_4^- \text{-->} Mn^{2+} + 4H_2O) \times 2$

$H_2O_2 \text{-->} O_2$

$(H_2O_2 \text{-->} O_2 + 2H^+ + 2e^-) \times 5$

$\cancel{10e^-} + \cancel{1}6H^+ + 2MnO_4^- \text{-->} 2Mn^{2+} + 8H_2O$

$5H_2O_2 \text{-->} 5O_2 + \cancel{10H^+} + \cancel{10e^-}$

$\mathbf{5H_2O_2 + 6H^+ + 2MnO_4^- \text{-->} 2Mn^{2+} + 8H_2O + 5O_2}$

2. What volume of MnO_4^- was added to reach the endpoint?

37.61 - 19.15 = **18.46 mL**

3. Knowing the volume and concentration of $KMnO_4$, how many moles of MnO_4^- were added to reach the endpoint?

18.46 mL x (1L/1000mL) x (0.0188mol/L) = **3.46 x 10^{-2} mol**

4. Using the balanced equation, how many moles of H_2O_2 were in your sample? *Recognize $KMnO_4$ IS MnO_4^- in the balanced equation.

3.46 x 10^{-2} mol $KMnO_4$ x (5mol H_2O_2/ 2 mol MnO_4^-) = **8.68 x 10^{-4} mol H_2O_2**

5. Convert the moles of H_2O_2 into grams.

8.68 x 10^{-4} mol H_2O_2 x (34.02g/1mol) = **0.0297g**

6. Imagine that your sample of H_2O_2 was mostly water and contained very little H_2O_2 molecules. If it was mostly water, it would be expected to have a density similar to water. The density of water is 1.00 g/mL. Let's assume the sample had a density of 1.00 g/mL. Using the mass of H_2O_2 calculated in the previous question along with the fact that your sample was exactly 25.00 mL, calculate the percentage of your sample that was H_2O_2. (0.0297/1.00) x 100% = **2.97%**

7. The label on the bottle states that the bottle of H_2O_2 is a 3% solution. Compare this value to your calculation in the previous question. Is the label correct? Why or why not? No. This is because H_2O_2 breaks down over time.

8. Is it necessary to know the exact volume of…

a) H_2O_2 sample added to the flask? yes

b) water added to the flask? no

c) $KMnO_4$ added to the flask? yes

Review Questions:

1. A 25.00mL sample of oxalic acid, $H_2C_2O_4$, was titrated with a standardized solution of $KMnO_4$ under acidic conditions. To reach the endpoint, 17.30mL of 0.0200M $KMnO_4$ was needed. At that point, a pink color persisted.

$MnO_4^-{}_{(aq)} + H_2C_2O_{4(aq)} \rightarrow CO_{2(g)} + Mn^{2+}{}_{(aq)}$

a. What does the pink color suggest?

You've reached the endpoint and reacted all moles of oxalic acid.

b. Balance the equation using the half-reaction method. Show all work.

$(5e^- + 8H^+ + MnO_4^- \text{-->} Mn^{+2} + 4H_2O) \times 2$ $(H_2C_2O_4 \text{-->} 2H^+ + 2CO_2 + 2e^-) \times 5$

$\cancel{10e^-} + \cancel{1}6H^+ + 2MnO_4^- \text{-->} 2Mn^{+2} + 8H_2O$

$5H_2C_2O_4 \text{-->} \cancel{10H^+} + 10CO_2 + \cancel{10e^-}$

$\mathbf{5H_2C_2O_4 + 6H^+ + 2MnO_4^- \text{-->} 2Mn^{+2} + 8H_2O + 10CO_2}$

c. How many moles of MnO_4^- reacted with the oxalic acid?

17.30 mL x (1L / 1000 mL) x (0.0200 mol/L) = **0.000346 mol MnO_4^-**

d. Based on your previous answer, how many moles of oxalic acid were present in the original sample?

0.000346 mol MnO_4^- x (5 mol $H_2C_2O_4$ / 2 mol MnO_4^-) = **0.000865 mol $H_2C_2O_4$**

e. Calculate the molarity of the $H_2C_2O_4$ solution, given your previous answer and the fact that 25.00 mL were used?

0.000865 mol / 0.025 L = **0.0346 M**

*25 mL = 0.025L

2. Our titration was performed in an acidic solution. Research the products of the redox reaction between MnO_4^- ions and H_2O_2 in a basic solution. How might the products under basic conditions impact your ability to know when you've reached the endpoint of the titration?

In basic solution, one of the products is manganese dioxide, which is a dark black-brown precipitate. This precipitate makes the solution cloudy, so that it is difficult to tell when you've reached the endpoint.