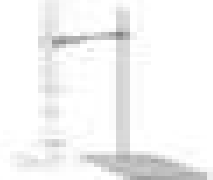




AP Chem Investigation 12 – Redox Titration Lab



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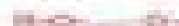
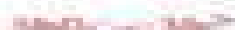
- Check to be sure you have at least 20 mL of KMnO_4 in your buret. If not, add some KMnO_4 to your buret.
- Be sure there are no air bubbles in the buret. If there are, draw some KMnO_4 out of the buret into a waste beaker until the air bubble comes out. Now run down the solution down the side with running water.
- Take an initial reading of KMnO_4 from the buret. Record this value in the data table.
- Measure out exactly 1.00 mL of H_2O_2 using a graduated pipet and add it to a clean flask.
- Add 50 mL of 0.1M H_2SO_4 to the flask using a graduated cylinder. This is so that the titration can occur in acidic conditions.
- Add water to the flask to bring the solution up to about 50-60 mL 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).
- Titrate the H_2O_2 to the endpoint (until a first-pink color persists). Try not to add even just a drop of KMnO_4 past the endpoint. Record the volume of titrant.
- Dump your solution down the sink. Leave the leftover KMnO_4 in your buret for the next class.
- 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 (mL)	1.00 mL
Initial reading of KMnO_4 volume on buret	14.11 mL
Final reading of KMnO_4 volume on buret	17.41 mL
Concentration of KMnO_4 solution on bottle	0.0250 M

Questions

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



- What volume of KMnO_4 was added to reach the endpoint?

$$17.41 - 14.11 = 3.30 \text{ mL}$$

- Knowing the volume and concentration of KMnO_4 , how many moles of KMnO_4 were added to reach the endpoint?

$$3.30 \text{ mL} \times (0.0250 \text{ mol/L}) \times (0.001 \text{ L/mL}) = 8.25 \times 10^{-5} \text{ mol}$$

- Using the balanced equation, how many moles of H_2O_2 were in your sample? **Remember 2KMnO_4 to $2\text{H}_2\text{O}_2$ in the balanced equation.*

$$8.25 \times 10^{-5} \text{ mol } 2\text{KMnO}_4 \times \text{Chem } (\text{H}_2\text{O}_2 : 2 \text{ mol } \text{H}_2\text{O}_2) = 8.25 \times 10^{-5} \text{ mol } \text{H}_2\text{O}_2$$

- Convert the moles of H_2O_2 to grams.

$$8.25 \times 10^{-5} \text{ mol } \text{H}_2\text{O}_2 \times (34.02 \text{ g/mol}) = 0.00280 \text{ g}$$

- 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 10.00 mL, calculate the percentage of your sample that was H_2O_2 .

$$(0.00280 / 0.0100) \times 100\% = 2.80\%$$

- 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 this labeled correct? Why or why not? *The 3% is because H_2O_2 breaks down over time.*

- Is it necessary to know the exact volume of...

a) H_2O_2 sample added to the flask?

b) water added to the flask?

c) KMnO_4 added to the flask?

yes

no

yes

Ap Chem Titration Lab Answers

D Keegan

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