



AP Chem Investigation 12 – Redox Titration Lab



Name KIT Period

- Check to be sure you have at least 20 mL of $K_2Cr_2O_7$ in your buret. If not, add some $K_2Cr_2O_7$ to your buret.
- Be sure there are not air bubbles in the buret. If there are, stretch some $K_2Cr_2O_7$ 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.
- Take an initial reading of $K_2Cr_2O_7$ 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 20 mL of 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 green color persists). Try not to add even just a drop of $K_2Cr_2O_7$ past the endpoint! Record the volume of titrant.
- Dump your solution down the sink. Leave the leftover $K_2Cr_2O_7$ in your buret for the next class.
- Check the label on the $K_2Cr_2O_7$ bottle and record the molarity of the $K_2Cr_2O_7$ solution in your data section.

Data Table

| | |
|--|------------------|
| Volume of H_2O_2 (mL) | <u>1.00 mL</u> |
| Initial reading of $K_2Cr_2O_7$ volume on buret | <u>19.11 mL</u> |
| Final reading of $K_2Cr_2O_7$ volume on buret | <u>17.41 mL</u> |
| Concentration of $K_2Cr_2O_7$ solution on bottle | <u>0.00880 M</u> |

Questions

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



- What volume of $K_2Cr_2O_7$ was added to reach the endpoint?

$$17.41 - 19.11 = \underline{1.70 \text{ mL}}$$

- Knowing the volume and concentration of $K_2Cr_2O_7$, how many moles of $K_2Cr_2O_7$ were added to reach the endpoint?

$$1.70 \text{ mL} \times (0.00880 \text{ mol/L}) \times (1000 \text{ mL/L}) = \underline{1.496 \times 10^{-2} \text{ mol}}$$

- Using the balanced equation, how many moles of H_2O_2 were in your sample? (Remember $2K_2Cr_2O_7$ to $6H_2O_2$ in the balanced equation.)

$$1.496 \times 10^{-2} \text{ mol } K_2Cr_2O_7 \times (3 \text{ mol } H_2O_2 / 2 \text{ mol } K_2Cr_2O_7) = \underline{2.244 \times 10^{-2} \text{ mol } H_2O_2}$$

- Convert the moles of H_2O_2 into grams.

$$2.244 \times 10^{-2} \text{ mol } H_2O_2 \times (34.02 \text{ g/mol}) = \underline{0.763 \text{ g}}$$

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

$$(0.763 \text{ g} / 3.00 \text{ g}) \times 100\% = \underline{2.54\%}$$

- 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? (The 3% is based on H_2O_2 breaking down over time.)

- If necessary to know the exact volume of...

a) H_2O_2 sample added to the flask? 1.70 mL

b) water added to the flask? 48.30 mL

c) $K_2Cr_2O_7$ added to the flask? 1.70 mL

Ap Chem Titration Lab Answers

Arturo Cuomo



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