



# 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 $H_2O_2$ sampled	1.00 mL
Initial reading of $K_2Cr_2O_7$ volume on buret	19.11 mL
Final reading of $K_2Cr_2O_7$ volume on buret	17.41 mL
Concentration of $K_2Cr_2O_7$ solution on bottle	0.0088 M

### 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?

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

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

$$1.70 \text{ mL} \times (0.0088 \text{ mol/L}) \times (1000 \text{ mL/L}) = 0.015 \text{ mol}$$

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

$$1.70 \times (1/3) \text{ mol } 6K_2Cr_2O_7 \times (2 \text{ mol } H_2O_2 / 6 \text{ mol } K_2Cr_2O_7) = 0.11 \text{ mol } H_2O_2$$

- Convert the moles of  $H_2O_2$  into grams.

$$0.11 \text{ mol } H_2O_2 \times (34.02 \text{ g/mol}) = 3.74 \text{ g}$$

- 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  $H_2O_2$ .

$$(3.74 / 20.00) \times 100\% = 18.7\%$$

- 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? (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? 38.30 mL

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

# Ap Chem Titration Lab Answers

**Michael Seilmaier**



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