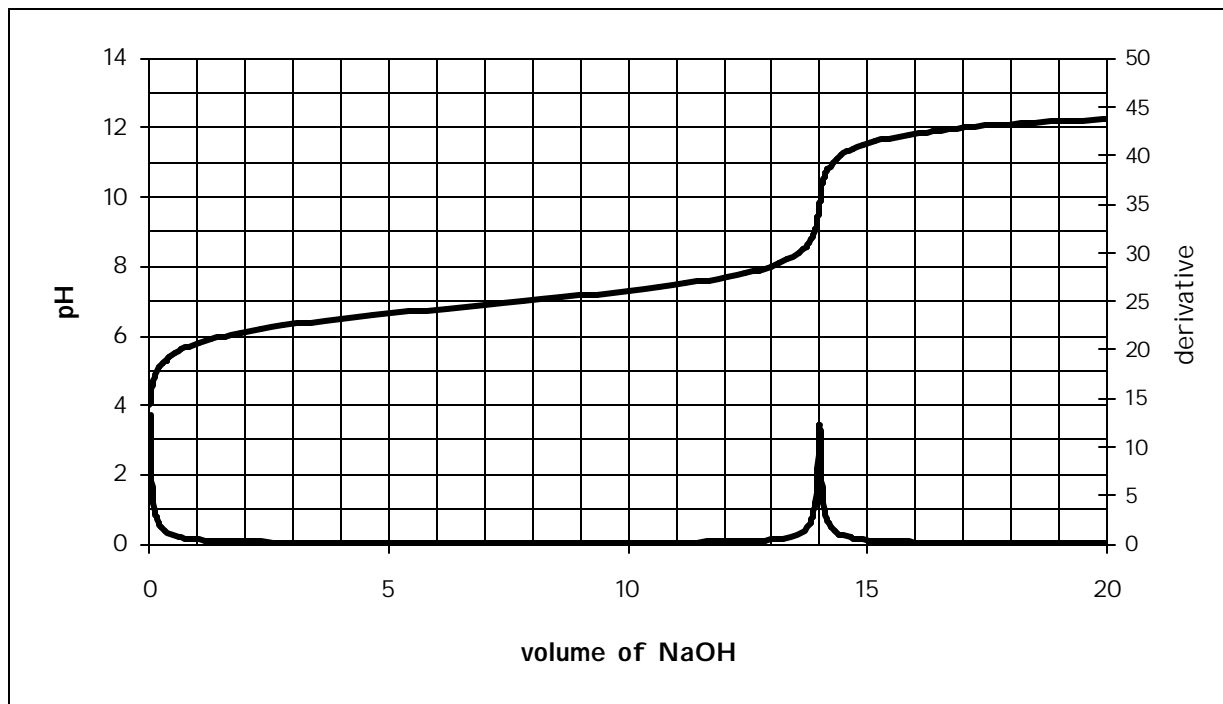
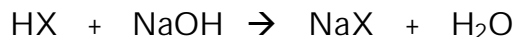


Problem Set on Tracking Reaction Progress

1. Consider the titration of a weak acid, HX, as given below. A weak acid sample of 216.1 mg was dissolved to make 15.00 mL solution. This was titrated with 0.1050 M NaOH.



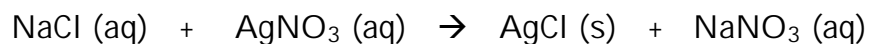
Locate the coordinates of the equivalence point. Describe how you determined this.

Calculate the molar mass of the acid.

What is the acid dissociation constant, K_a , of the weak acid? Explain or illustrate how you determined this.

What indicator would work for this titration? Explain your choice. Site the reference used.

2. The table given below contains the conductometric titration data for the following precipitation reaction. The AgNO_3 is the titrant.



Would this reaction show any visible signs? If so, what?

Plot the data and locate the equivalence point of the titration. Explain and illustrate on the graph how you determined the equivalence point.

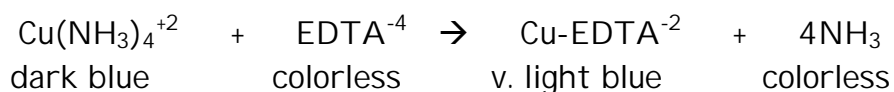
Volume, mL	Conductivity
0	2898
2.34	2875
5.15	2864
7.40	2852
10.20	2845
12.45	2829
14.70	2806
18.06	2795
19.75	2783
23.12	2990
27.06	3519
30.43	3979
34.92	4577
39.97	5313
45.03	5957

Explain what ions are causing the conductivity at the following points:

- initially
- before the equivalence point
- at the equivalence point
- after the equivalence point

The ionic conductance at 25°C and infinite dilution for the Ag^+ is $\lambda_+ = 62$.

3. Consider the spectrophotometric titration of $\text{Cu}(\text{NH}_3)_4^{+2}$ with EDTA. Here is the reaction occurring with EDTA as the titrant.



The titration is done at the λ_{max} of $\text{Cu}(\text{NH}_3)_4^{+2}$. The Cu-EDTA^{-2} complex has a very large formation constant ($K_f = 10^{18}$).

Plot the data in the table given below and locate the equivalence point.

Volume, mL	Absorbance
0	0.505
3.0	0.411
6.0	0.297
9.0	0.209
12	0.107
15	0.051
18	0.049
21	0.050
24	0.048

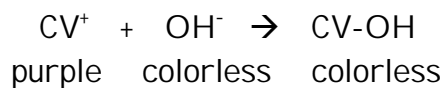
Explain the cause of the absorbance for the titration at the following points:

- initially
- before the equivalence point
- at the equivalence point
- after the equivalence point

If you titrated 100 mL $\text{Cu}(\text{NH}_3)_4^{+2}$ solution with 0.12 M EDTA, what is the initial molar concentration of the copper complex?

Sketch a dashed line (---) on the graph to show the reaction if the Cu-EDTA^{-2} complex was colorless.

4. Consider the crystal violet kinetics run data given in the table below. This is following a reaction over time by using spectrophotometry. The reaction is given below.



Time	Absorbance
0 sec	0.50
3	0.47
6	0.44
9	0.40
12	0.38
15	0.33
18	0.29
21	0.27
24	0.23
27	0.20
30	0.18
33	0.16
36	0.12
39	0.09

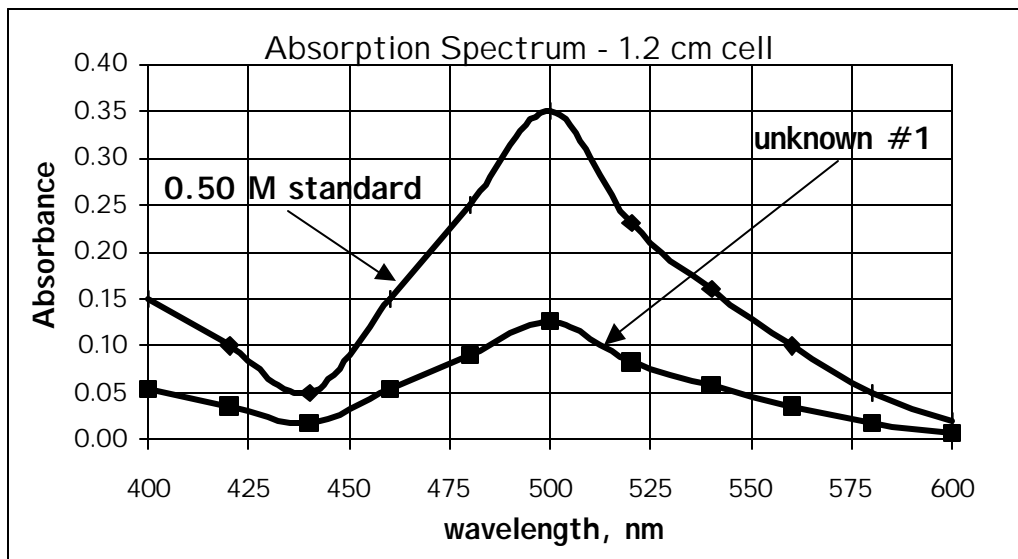
Plot a graph of the data. How does the absorbance change with time? Why?

Calculate the average initial rate of the reaction for this run. Explain your method.

How good a fit is your line to the data?

Sketch a dashed line (----) on the graph above that shows a run with a greater rate of reaction.

5. A student determined the absorption spectrum of a standard solution and then repeated the process for an unknown. What is the concentration of the unknown #1 solution? Explain your method and state the wavelength used.

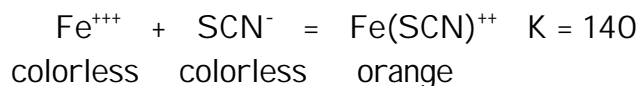


The student went to run the absorption spectrum of unknown #2 and found that it had too high an absorbance ($A > 1$). The student set the wavelength to 460 nm and ran the unknown #2, which gave an absorbance of 0.59. Calculate the concentration of unknown #2. Explain your method. Why did this work?

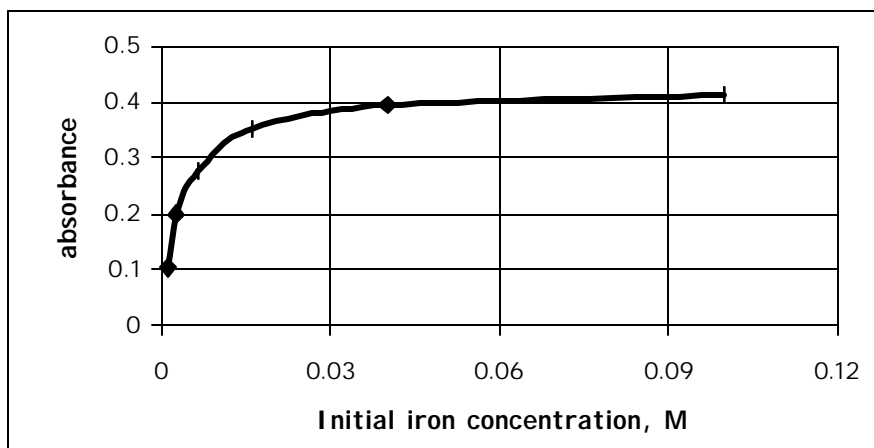
The determinations above are based on a single-point calibration, which was used in the equilibrium constant determination. When performing a single-point calibration, a very important assumption is being made. Explain this assumption.

To analyze unknown #2 at a wavelength of 500 nm, what could be done to the sample to lower the absorbance?

6. Consider the determination of the equilibrium constant for the reaction given below.



The plot shows the behavior of the absorbance as the initial iron concentration, $(\text{Fe}^{+++})_0$, is increased with the $(\text{SCN}^-)_0$ remaining constant as in the experiment. Why is the absorbance changing?



In this experiment the first test tube (0.10M Fe^{+3}) was the standard used to measure the other test tubes as the iron concentration was decreased. We assume that all the SCN^- is tied up in the complex. Or the following equation must hold:

$$(\text{SCN}^-)_0 = (\text{Fe}(\text{SCN})^{++})$$

How does the absorbance behave as the concentration of the initial iron concentration, $(\text{Fe}^{+++})_0$, is increased?

If you were to extend the graph to higher initial iron concentration, how would it behave? Explain or illustrate on the graph.

If you used an initial concentration of 0.20M Fe^{+3} , what would the absorbance do?