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I. Lab Report and Notebook Guidelines

You laboratory notebook has to be kept using the following guidelines:

- Use only one 9.75 x 7.5 in Quad Rule Composition Book.
- Pens only! No use of pencils in the laboratory!
- Continuously number the pages in the notebook in the top outside corner.
- First right-hand page (p. 2): Name
- Second right-hand page (p. 4): Table of Content (update during the quarter).
- Subsequently, you will use the left-hand pages to write down notes (measurements, calculations, changes to the procedure, etc.) during the lab period. Right-hand pages are used for the lab report (see below).
- Do not rip any pages out of your notebook! Do not use white-out! If you make a mistake, cross line through your wrong notes/calculations. Use of white-out or removal of pages will lead to deductions on the grade for that lab (-5% per violation).
- The lab report has to follow the sequence outlined below. Always use complete, grammatically correct sentences in all sections. Do not use bullets or numbering other than those indicated below. This will lead to deductions!

As a general guideline, your lab reports should be written in a way that you or another student could use it to perform the experiments.

1. **Introduction.**
   Describe the purpose of the experiment in your own words.

2. **Experimental/Data/Observations.**
   Write a step-by-step description of what you did. Report weights, chemical reactions, and any other observation you made that are relevant to the lab. You also report things that went wrong. Make certain that you really write what you did, not what you were supposed to do according to this lab manual. In this section, a well kept diary on the left-hand pages comes in handy, since it allows you to remember what you actually did when it comes to writing the lab report. Remember, your result will, among other things, be judged by what you did or did not do. This means, if you make mistakes and you are honest about them, I may allow larger deviations from the actual value. However, if you claim you followed the instructions to the letter, then I expect your result to closely correspond to the actual value. This section ends with the report of your result, which has to include a statistical analysis (standard deviation, t-test at 95% confidence level, etc., see individual labs for required statistical analyses). All calculations have to be done in the lab book (except the Excel spreadsheets, but in this case you have to provide spreadsheet documentation). Never use loose paper to write down anything or calculate anything related to this lab.

3. **Discussion.**
   Evaluate the precision of your result(s) based on your observations. Mention possible sources of errors (even if you think you didn’t make any mistakes). This section is not supposed to be a listing of “standard mistakes made.” That means I do not want to continuously see “read the buret wrong, weighing errors, etc.” Rather, I expect you
to place your result in context. That means, is your result reasonable? Why or why not? If not, what could have gone wrong?

4. **Conclusion.**
   Report the final result, with error and units as described for each experiment. **Include your unknown number.**

II. General laboratory grading breakdown

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<td>8</td>
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</tbody>
</table>
Experiment 1
Determination of Water Hardness and Ca\(^{2+}\) and Mg\(^{2+}\) in Water by Compleximetric Titration

Background
Read p. 364 – 365, and Section 15D-9 (p. 380 – 382) in your textbook (Skoog, West, Holler, & Crouch Analytical Chemistry, An Introduction). Chapter 15 in your textbook describes the principles of complexation and will be covered in the last third of this quarter.

Procedure

A. Prepare a 0.003 M Na\(_2\)EDTA solution [Individual]
Dry Na\(_2\)EDTA at 80 °C. Weigh out dried and cooled Na\(_2\)EDTA (~0.6 g) Dissolve it in ~400-mL of water and mix well. Add a couple of drops of 2 M NaOH solution to the Na\(_2\)EDTA solution if you have difficulties dissolving it. Once the Na\(_2\)EDTA is dissolved transfer it to a 500 mL volumetric flask and fill to volume. Calculate the molarity of this solution. **Looking ahead:** the calculated molarity cannot be used because commercially available Na\(_2\)EDTA is not 100% pure. The actual molarity which is used to determine the water hardness is established through titration. **Compare the calculated and the actual molarity determined through titration in your lab report to illustrate this fact.**

B. Preparation of a 0.01 M Zn\(^{2+}\) solution, **Primary Standard** [Individual]
Weigh out (to the nearest 0.1 mg) about 0.33 g pure zinc granules and transfer to a 400-mL beaker. Dissolve in a minimum amount (~10 mL) of 6 M HCl. Cover the beaker with a watch glass and heat gently to dissolve (hot plate on lowest setting). After the zinc is in solution, wash the droplets of water on the watch glass with distilled water into the beaker, wash down the sides of the beaker, and quantitatively transfer the solution to a 500-mL volumetric flask. Let the solution cool to room temperature and then dilute to volume with distilled water. Calculate the molarity of this solution.

C. Titrations of Zn\(^{2+}\) with EDTA [Individual]
Pipet 10.00 mL of the standard zinc solution into an Erlenmeyer flask. Add 10 mL of pH 10 buffer and 2 to 3 drops of Eriochrome Black T. Titrate with standard EDTA solution to a color change from red to blue. **Note:** If the color changes from red to colorless, add 2 – 3 more drops of Eriochrome Black T until color is visible. Do in triplicate and calculate the titer (mean and standard deviation) of the EDTA solution based on the molarity of the zinc solution.

D. Titrations of EDTA with Zn\(^{2+}\) [Individual]
Pipet 50.00 mL of the standard EDTA solution into an Erlenmeyer flask. Add 10 mL of pH 10 buffer and 1 to 2 drops of Eriochrome Black T. Titrate with the standard zinc solution to a color change from blue to red. Do it in triplicate and calculate the titer (mean and standard deviation) of the EDTA solution based on the molarity of the zinc solution.
Run a student’s t-test (at 95% CL) to determine if the calculated EDTA concentrations from part C and D can be pooled. If $t_{\text{calc}} < t_{\text{table}}$, pool the data from part C and D and use the average EDTA concentration for calculations in part E. If not, talk to the instructor.

E. Determination of the hardness of water [Individual]

1. Transfer 50.00 mL of the unknown water sample to a 250-mL Erlenmeyer flask, add 1 to 2 mL of pH 10 buffer and 3 to 4 drops of Eriochrome Black T indicator. Ca$^{2+}$ ions are removed by adding approximately two spatulas of (NH$_4$)$_2$CO$_3$ and swirling the flask for ~2 minutes. Titrate with the 0.003 M Na$_2$EDTA solution until the color changes from red to pure blue. Repeat the titration two more times. Note: The color change tends to be slow in the vicinity of the end point. Care must be taken to avoid overtitration.

2. Transfer 50 mL of the unknown water sample to a 250-mL Erlenmeyer flask and add 15 drops of 50 wt% NaOH solution. Test the pH to make sure it is basic. Swirl the flask for ~2 minutes to allow Mg(OH)$_2$ to precipitate fully. Weigh out about 1 g of solid hydroxynaphthol blue indicator and add it to the basic solution. Titrate this solution with the 0.003 M Na$_2$EDTA solution to the equivalence point. Repeat the titration two more times.

Report

Calculate the parts per million (ppm) of Mg$^{2+}$ and Ca$^{2+}$ in your sample. Report these data with their standard deviations (stdv.). Also report the total hardness (ppm Mg$^{2+}$ + ppm Ca$^{2+}$) with its stdv (which is obtained through error propagation, Table 6.4, p. 140 of your textbook).
Experiment 2
Ascorbic Acid Content in Vitamin C Tablets via Iodometric Titration

Background
This is an example for a redox titration. Chapter 18 in the textbook covers this topic (and will be covered in class towards the end of the quarter). Read experiment 271-3 (p. 756) for further details.

Relevant equations
\[
\begin{align*}
10^{-} + 5I^{-} + 6H^{+} & \leftrightarrow 3I_{2} + 3H_{2}O \\
I_{2} + 2S_{2}O_{3}^{2-} & \leftrightarrow 2I^{-} + S_{4}O_{6}^{2-} \\
C_{6}H_{8}O_{6} \text{ (ascorbic acid)} + I_{2} & \leftrightarrow C_{6}H_{6}O_{6} \text{ (dehydroascorbic acid)} + 2I^{-} + 2H^{+}
\end{align*}
\]

Procedure
Dry ~1 g of KIO\textsubscript{3} at 80 °C. While this is drying, boil 500 mL of distilled water.

A. Solution 1 (Indicator) [Group]
Weigh out 5 g of soluble starch in a 600-mL beaker and rub it into a paste with about 50 mL of water. Dilute to about 500 mL with the boiling water, then heat until no whitish material is visible. Cool to room temperature, add 5 mL of toluene and store in a tightly stoppered bottle.

B. Solution 2 [Individual]
Dissolve 0.05 g of Na\textsubscript{2}CO\textsubscript{3} in 500 mL of distilled water and bring to a boil. Use this water to prepare a Na\textsubscript{2}S\textsubscript{2}O\textsubscript{3} solution by dissolving about 8.7 g Na\textsubscript{2}S\textsubscript{2}O\textsubscript{3}•5H\textsubscript{2}O in these 500 mL. Let the solution cool to room temperature.

Note: The exact molarity of this solution, which is needed to calculate the amount of Vitamin C from the titrations in part E. is established in part D.

C. Solution 3 [Individual]
Prepare a ~0.01 M KIO\textsubscript{3} solution by dissolving 1 g dried KIO\textsubscript{3} in 500 mL of distilled water (in a 500-mL volumetric flask). This is your primary standard.

D. Standardization of the Na\textsubscript{2}S\textsubscript{2}O\textsubscript{3} solution [Individual]
Standardize the Na\textsubscript{2}S\textsubscript{2}O\textsubscript{3} solution by pipetting 50.00 mL of Solution 3 into a 250-mL flask, add ~2 g KI(s) and 10 mL of 0.5 M H\textsubscript{2}SO\textsubscript{4}. Titrate this solution with Solution 2 until the solution is yellow. Add 2 mL of indicator (Solution 1) and complete the titration. Repeat this titration two more times.

E. Measurement of the ascorbic acid concentration [Individual]
Weigh two tablets of Vitamin C to ±0.1 mg using the analytical balance. Using mortar and pestle, grind the tablets into a powder. Weigh about 1/4 of the ground tablets to ±0.1 mg using the analytical balance. Dissolve this amount in 60 mL of 0.3 M H\textsubscript{2}SO\textsubscript{4} in a 250 mL Erlenmeyer flask.

Pipet 50 mL of Solution 3 into the 250 mL Erlenmeyer flask containing the dissolved Vitamin C tablet, add 2 g of KI(s) and dissolve. If the color does not change to dark red, add more Solution 3 in 10 mL increments until the solution is deep red.
Note: If around 100 mL of Solution 3 (total, i.e. 50 mL original to 5 x 10 mL increments) are required to obtain a color change to red, use less (~1/8) of the tablet powder to prepare the second and third titration solution.

Then titrate with Solution 2 as described before, including the addition of indicator (Solution 1) just before the end point. However, in this case no yellow color will be observed. Add the indicator when the solution turns from dark red to lighter red. Repeat this titration at least two more times. Run the next trials with either about ¼ or 1/8 (see Note above) of the powdered tablets. Remember to record the weight of the full tablet as well as of the amount of powder you used in each trial for the final lab report. Additionally, do not forget to record how much total Solution 3 you used in each trial to obtain a dark red solution.

Report

Report your final result as mg ascorbic acid/tablet Vitamin C ± stdv.
Experiment 3
Ion Specific Electrodes: Fluoride Content in Water and in Toothpaste

Background
Read experiment 27J-4 (p. 761) and sections 19D (p. 475 ff.) and 19F (p. 493 ff.) in your textbook.

Procedure
A. Preparation of a calibration curve [Individual]
   1. A stock solution of 0.10 M $F^-$ should be used to prepare, by serial dilution, 100 mL of each of the following concentrations of the desired ion: $5 \times 10^{-2}$, $10^{-2}$, $5 \times 10^{-3}$, $10^{-3}$, and $5 \times 10^{-4}$ M along with a blank. The ionic strength of each standard solution should be identical; consequently a high concentration of an unreactive ionic compound or buffer is added to each. In this case, the compound of the commercial “total ionic strength adjusting buffer” (TISAB) complexes iron and aluminum which interfere with the $F^-$ determination and adjusts the pH to 5 as well as maintaining constant ionic strength.

      Add 10 ml of TISAB II to each 100-mL flask before filling to volume.

      2. Measure the potential of each of the above solutions (use a 50-mL plastic beaker) at room temperature in the order of lowest concentration to highest using the appropriate ion-specific electrode. Be sure to rinse the electrode with distilled water and blot dry between measurements. Measurements should be recorded while the solution is being stirred and after the potential has stabilized. Repeat the measurements by measuring the same set of solutions two more times going from least concentrated to most concentrated each time.

      Plot the electrode potential vs. pF (-log[$F^-$]). Examine the data and ascertain the minimum concentration that can be measured with this electrode. Using only the appropriate data, calculate the slope of the plot.

B. Aqueous unknown [Individual]
   1. Preparation of samples: Pipet 10, 25, and 50 mL of the unknown solution into three separate 100-mL volumetric flasks. Add 10 mL of the TISAB solution and fill to volume.

   2. Standard Addition: Pipet 50.0 mL of the diluted unknown (50/100 dilution) into a plastic beaker and measure the potential. Measure the potential of this solution and determine its concentration from the above calibration curve. Pipet 5.00 mL of the standard solution which is 10 times more concentrated than the concentration of the diluted aqueous unknown as determined from the calibration curve into a 100-mL beaker. Mix and record the potential of this solution in the same fashion, without making any adjustment of the pH/mV meter. Calculate the concentration of the unknown.
3. Measure the potential of the 1/100 and 10/100 dilutions of the unknown. Determine the concentration of $F^{-}$ in these solutions from the calibration curve and/or the standard addition method.

4. Summarize your determinations of the fluoride concentration in your unknown (correcting for dilutions) in a table in your lab notebook as outlined below.

<table>
<thead>
<tr>
<th>Sample</th>
<th>mV reading</th>
<th>[$F^{-}$]uncorrected</th>
<th>[$F^{-}$]corrected</th>
</tr>
</thead>
<tbody>
<tr>
<td>50/100</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>50/100 + 5</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>10/100</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>10/100 + 5</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1/100</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1/100 + 5</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

C. Fluorinated toothpaste [Individual]

Weigh out accurately around 750 mg of toothpaste (make sure $F^{-}$ is the fluorine species) into a 250-mL beaker. Add 10 mL of TISAB solution, heat to almost boiling on a hot plate (careful, the suspension easily bumps), and let simmer for 2 minutes; the toothpaste will not completely dissolve. Let cool. Transfer quantitatively to a 100-mL volumetric flask and dilute to volume.

Use the same procedure as described in part B. for the aqueous unknown to determine the fluoride content of the toothpaste. Report your data in tabular form as described below and remember that you prepared different dilutions of the toothpaste solution.

Repeat the measurements by measuring the same set of solutions two more times going from least concentrated to most concentrated each time.

<table>
<thead>
<tr>
<th>Sample</th>
<th>mV reading</th>
<th>[$F^{-}$]uncorrected</th>
<th>[$F^{-}$]corrected</th>
</tr>
</thead>
<tbody>
<tr>
<td>50/100</td>
<td></td>
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<td></td>
</tr>
<tr>
<td>50/100 + 5</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>25/100</td>
<td></td>
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<td></td>
</tr>
<tr>
<td>25/100 + 5</td>
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<td>10/100</td>
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<td></td>
<td></td>
</tr>
<tr>
<td>10/100 + 5</td>
<td></td>
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</tr>
</tbody>
</table>

D. Density of toothpaste [Group]

Devise a way to determine the density of the toothpaste and use that method to establish the density of toothpaste. Describe your method and show your calculations.

Report

Calculate and report the fluoride content of the water sample as $\text{mol/L} \pm \text{stdv}$. Calculate the amount of fluoride in the toothpaste and report it as $\text{wt/v\%}$, using, among other things, the density of the toothpaste you determined.
Experiment 4
Absorption Spectrophotometry. Iron Content of Vitamin Supplements.

Background
Read p. 594-604 as well as 27L-1 and 27L-2 (p. 764 – 765) in your textbook for further information about this experiment.

Procedure:
A. Preparation of Fe standard solution (0.01 mg/ml).[Group]
   Weigh (to the nearest 0.1 mg) 0.0702 g of reagent grade Fe(NH$_4$)$_2$(SO$_4$)$_2$$•$6H$_2$O into a 1-L volumetric flask. Dissolve in 50 mL of water that contains 1-2 mL of 98% wt% H$_2$SO$_4$; dilute to the mark and mix well.

B. Dissolution of the Vitamin tablet [Individual]
   Weigh two tablets of the iron supplement to ±0.1 mg using the analytical balance. Using mortar and pestle, grind the tablets into a powder. Weigh about 1/2 of the ground tablets to ±0.1 mg using the analytical balance and place it in a 125-mL flask or beaker. Add 25 mL of 6 M HCl and boil gently (in the hood) for 15 minutes. This procedure dissolves the Fe. Transfer this solution quantitatively to a 100-mL volumetric flask and dilute to volume. Dilute 5.00 mL of this solution to volume in a 100.00-mL volumetric flask. This final diluted solution is the one that will be analyzed for iron content in part D. Looking ahead: Remember the dilutions when you calculate the Fe content of the Vitamin tablet.

C. Generating a standard curve (A vs. concentration) from standard iron solution. [Individual]
   Standards: Into a series of 5 100-mL volumetric flasks, pipet 5.00, 10.00, 15.00, 20.00, and 25.00 mL of the standard iron solution.
   Blank: Add 10.00 ml of distilled water to a 100-mL volumetric flask containing no iron.

   To each of the above flasks, add 10 mL of the phenanthroline solution, 1.0 mL of the hydroxylamine hydrochloride solution, and 10 mL of the sodium acetate solution before filling to volume. Wait 15 minutes for the red color to develop. Fill to volume and mix well.

   Obtain the absorbance of the above solutions at 520 nm using one of the Spec-20 spectrophotometers.

D. Standard Addition [Individual]
   Pipet 10.00 mL aliquots of the diluted vitamin tablets into six separate 100-mL volumetric flasks. Then add the following amounts of the standard iron solution to one of each of the flasks containing diluted iron unknown solution: 0.00, 5.00, 10.00, 15.00, 20.00, and 25.00 mL.

   To each of the above flasks, add 10 mL of the phenanthroline solution, 1.0 mL of the hydroxylamine hydrochloride solution, and 10 mL of the sodium acetate solution before filling to volume. Wait 15 minutes for the red color to develop. Fill to volume and mix well.
E. Measuring the Absorbance Spectrum [Individual]

Obtain and record the absorbances of the above solutions at 520 nm using one of the Spec-20 spectrophotometers.

After taking the first readings for the Vitamin Fe samples (Part D.), the student has to prepare additional samples to obtain accurate and precise evaluation of the iron content. This is done by preparing replicate samples from a new digest of the remaining ground tablet not previously used in Part B.

Report

Using an Excel spreadsheet, plot both the standard curve (Part C.) as well as the two standard addition curves (Part E.) following the example on p. 601 and 603 of the textbook. Remember to include a spreadsheet documentation. Print the spreadsheet and linear regression and glue it into your lab report where appropriate. You should have obtained four separate ppm values for Fe$^{2+}$. Two are obtained from the spreadsheets from Part E. while the other two are obtained by solving the linear regression equation calculated from Part C. using the absorbance values recorded for the solutions prepared in Part D. that contain no (0.00) standard Fe$^{2+}$ solution. Convert each ppm Fe$^{2+}$ value in a mg Fe$^{2+}$/tablet value. Average both the ppm Fe$^{2+}$ and the mg Fe$^{2+}$/tablet values and report your result as \( \text{ppm } \text{Fe}^{2+} \pm \text{stdv} \) and as \( \text{mg } \text{Fe}^{2+}/\text{tablet} \pm \text{stdv} \).
Experiment 5
The pKₐ of an Indicator

Background
Read Chapter 22 in your textbook, with specific emphasis on 22A-1 to 22 A-3, and 22B.

Procedure

A. Absorption spectra of Bromothymol Blue at various pH values [Group]

Obtain three complete absorption spectra of bromothymol blue at three pH values; use conditions of approximately pH 1, pH 7, and pH 13.

1. pH ≈ 1. Carefully pipet 1 mL of bromothymol blue stock solution into a clean 25-mL volumetric flask. Add a few milliliters of distilled water, then 4 drops of concentrated HCl, and finally dilute to the mark with distilled water. Invert several times to effect mixing, and add about 5 mL of this solution to a cuvette that has been rinsed with this solution. Record the absorbance spectrum of this solution between 365 and 800 nm using the UV-Vis spectrophotometer. Indicate the color of the solution.

2. pH ≈ 7.2. Pipet 1 mL of indicator into a 25-mL volumetric flask and add 5 mL each of 0.10 M Na₂HPO₄ and KH₂PO₄ from a pipet. Dilute to the mark and obtain the spectrum as above. Indicate the color of the solution.

3. pH ≈ 13. To 1 mL of indicator in a 25-mL volumetric flask, add 12 drops of 4 M NaOH. Dilute to the mark and obtain the spectrum as above. Indicate the color of the solution. If prepared properly, the three curves should intersect each other at a single point, called an isosbestic point. If they do not intersect at one point, remake the appropriate solution and remeasure it until a proper isosbestic point is obtained.

B. Absorbance spectra of solutions with different pHs [Group]

Measurements will be made on solutions with seven different pH values other than the three solutions studied thus far.

<table>
<thead>
<tr>
<th>mL Indicator</th>
<th>mL H₂PO₄⁻</th>
<th>mL HPO₄²⁻</th>
<th>pH</th>
</tr>
</thead>
<tbody>
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<td>5.0</td>
<td>0.0</td>
</tr>
<tr>
<td>2.</td>
<td>1.0</td>
<td>5.0</td>
<td>1.0</td>
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<tr>
<td>3.</td>
<td>1.0</td>
<td>10.0</td>
<td>5.0</td>
</tr>
<tr>
<td>4.</td>
<td>1.0</td>
<td>5.0</td>
<td>10.0</td>
</tr>
<tr>
<td>5.</td>
<td>1.0</td>
<td>1.0</td>
<td>5.0</td>
</tr>
<tr>
<td>6.</td>
<td>1.0</td>
<td>1.0</td>
<td>10.0</td>
</tr>
<tr>
<td>7.</td>
<td>1.0</td>
<td>0.0</td>
<td>5.0</td>
</tr>
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</table>
Pipet the above quantities for a given pH into a 25-mL volumetric flask and dilute to the mark. Record the absorbance spectrum of each solution between 365 and 800 nm using the UV/Vis spectrophotometer. Indicate the colors of the solutions. If prepared properly, all spectra should intersect each other at the isobestic point. If they do not intersect at one point, remake the appropriate solution and remeasure it until each solution intersects the isosbestic point. Calculate the pH of the above solutions given the fact that the second ionization constant for H$_3$PO$_4$ is $6.32 \times 10^{-8}$ ($pK_{a2} = 7.20$).

C. Determining the $pK'_{a}$ value [Individual]

Refer to the spectra you have recorded and select two wavelengths for which exact absorbance values will be used. You should select a wavelength to the left of the isosbestic point and one to the right of this point. Choose wavelengths where the acid and base forms of the indicator show a maximum difference in their absorbance (peak areas).

Combining the absorbance values at the two selected wavelengths (obtained in Procedure A) with the data obtained in Procedure B, plot absorbance (vertically) vs. pH (horizontally) for each of the two wavelengths studied. Connect the points with a smooth curve. The midpoint of each curve corresponds to equal concentrations of the acid and the base form of the indicator. From each graph determine the $pK'_{a}$ of the indicator. The two values thus obtained for the $pK'_{a}$ may differ slightly.

Draw two horizontal lines across each of your a vs. pH plots; one corresponding to the absorbance of the pH $\approx 1$ solution of the indicator, and the other corresponding to the absorbance of the pH $\approx 13$ solution. The first line (I) gives the absorbance of the indicator when present entirely in the acid form. The second line (II) gives the absorbance of the indicator when present entirely in the base form. Any deviation of the actual absorbance from these two lines is a measure of the extent to which one form of the indicator has been converted to the other form.

Consider the absorbance reading obtained for the pH 6.5 solution. By subtracting the absorbance at pH 6.5 from the absorbance at pH 1 (line I of the curve), a measure of the amount of Ind$^{2-}$ in the solution at pH 6.5 can be obtained. By subtracting the absorbance at pH 13 (line II of the curve) from the absorbance at pH 6.5, a comparable measure of the amount of HInd$^{-}$ in the pH 6.5 solution can be obtained. The ratio Ind$^{2-}$ to HInd$^{-}$ may be obtained by dividing the results of the two preceding calculations.

Determine the Ind$^{2-}$/HInd$^{-}$ ratio that corresponds to each of the points plotted on the absorbance-vs.-pH graph at the wavelength to the right of the isosbestic point, and plot log(Ind$^{2-}$/HInd$^{-}$) (vertically) vs. pH on a graph. The point where the line (obtained from linear regression) crosses the vertical axis at zero corresponds to equal concentrations of the basic and acidic forms of the indicator. From the pH at which the line crosses the vertical axis on your plot, determine the $pK'_{a}$ of the indicator.

Report

Report your three values for the $pK'_{a}$, the average $pK'_{a}$ and the standard deviation.