Simple Titration Lab

Introduction:
Titrations are used to determine the exact concentration of a solution of unknown concentration.

Objective:
You are going to determine the concentration of 100.0 mL of a HCl solution using 0.1 M NaOH. 
YOU WILL BE DOING THREE TRIALS OF THE TITRATION.

Procedure Part One: setting up the titration
1. Place a waste beaker under the buret and close the stopcock on the buret.
2. Using a funnel, fill the buret about half way with 0.1 M NaOH.
3. Open the stopcock and let a little of the the NaOH run into the waste beaker. This is to clean the tip of the buret and to get any air bubbles out. Discard the contents of the waste beaker down the sink.
4. Using a funnel, fill the rest of the buret with 0.1 M NaOH to some value between 0 and 1 mL and accurately record the initial volume in the data table. Notice that the buret is numbered backwards. BE CAREFUL when reading the buret (see the example below). You might need to stand on a chair to get to eye level to properly read the meniscus.

Example:

5. Using a graduated cylinder, pour exactly 100.0 ml the HCl into a clean Erlenmeyer flask. Add 3-4 drops of phenolphthalein to the HCl in the flask.

6. Place the flask containing the HCl underneath the buret tip (make sure the buret is CLOSED). It helps to put a piece of white paper underneath the flask so you can see the color change easier once you start titrating.
Procedure Part Two: practice titration

The first trial will be a practice run. You will likely mess it up. That is okay.

1. Open the stopcock to start adding the base to the acid quickly until you start to see a color change then close the stop cock. **Swirl the flask.**

2. Now start to add the NaOH a few drops at a time. **Make sure you are swirling the Erlenmeyer flask as you go so it stays evenly mixed.**

3. As the color starts to persist try to add a drop or two at a time until you get a very pale pink that doesn’t disappear. This point is your end point or equivalence point.

4. Record the final volume of base used in the data table provided. Subtract the initial volume from the final volume to get the amount of base used. Record this in the data table provided. Now you have an idea of how much NaOH you will need to add to your acid to neutralize it.

Example:

![Buret diagram]

Starting volume

Volume Change

35.4 - 14.6 = 20.8 mL

Ending volume

7. Empty the flask of Acid/base into the sink, rinse well and dry.

Procedure Part Three: Second and third Titration: the REAL ones.

1. Prepare another sample of exactly 100.0 mL of HCl with 3-4 drops of phenolphthalein in a clean Erlenmeyer flask.

2. If you need to, add more NaOH to the buret and accurately record the initial volume of NaOH in the buret in the data table provided. Make sure a waste beaker is under the tip just in case some spillage occurs.
3. Place the flask containing the HCl underneath the buret tip (make sure the buret is CLOSED). It helps to put a piece of white paper underneath the flask so you can see the color change easier once you start titrating.

4. Add NaOH solution to the beaker of acid, fast at first (say in 1 ml increments), then slowly (1 drop at a time) when you near the equivalence point volume that you determined in the practice titration. **Make sure you are swirling the Erlenmeyer flask as you go so it stays evenly mixed.**

5. Stop titrating when the solution in the flask turns **VERY LIGHT pink** and stays that way for at least 10 seconds. Record the volume of NaOH needed to reach this point in the data table provided.

6. Empty the flask in the sink and rinse well.

7. Repeat steps 1-6 for a third titration.

**Data Table:**

<table>
<thead>
<tr>
<th>Trial #</th>
<th>Amount of HCl in flask</th>
<th>Initial buret reading</th>
<th>Final buret reading</th>
<th>Volume NaOH added</th>
</tr>
</thead>
<tbody>
<tr>
<td>#1: practice</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>#2: real one</td>
<td></td>
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<td></td>
</tr>
<tr>
<td>#3: real one</td>
<td></td>
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</tbody>
</table>

**Calculations. Show all your work.**

1. Write the balanced equation for the titration:
2. Determine the average volume of 0.1 M NaOH needed to titrate 100.0 mL of the HCl solution, using only the data from the 2\textsuperscript{nd} and 3\textsuperscript{rd} titrations.
3. Determine the concentration of the HCl using the average volume of the 0.1M NaOH calculated in #2.