Titration Lab

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Abstract:

In this lab, we used the technique of titration to determine the molarity of an acid. This was a concept learned in class, and this lab demonstrated how this concept could be applied to the real world. My particular HCl concentration I found to be 0.112M HCl in part 2 of the lab. I determined this value by dripping a base of a known concentration into the HCl with phenolphthalein in the solution until the solution was a very light pink. This meant all the acid had been neutralized. Using the recorded volume of acid, the molarity of the base and the volume of the base, I was able to find the molarity of the acid. However, there were certainly some errors. Often, drops of NaOH would stay on the sides of the buret, and so I was unable to obtain a very accurate volume of base used. This might have caused my calculated molarity to be higher as a greater volume of base was "used." Conversely, we lost some KHP in the beginning so the molarity of our base might be a little low, leading to a lower molarity of acid.

Purpose:

We will determine the molarity of an unknown acid using titration. We will apply concepts learned in class to a real-world situation. These concepts include molarity, titration, volume and the equation relating the three: $\frac{MV}{F} = \frac{MV}{F}$

Materials:

- 50 ml buret
- NaOH (≈0.1 M)
- buret clamp
- Ring stand
- KHP
- 250 ml Erlenmeyer flask
- 100 ml graduated cylinder
- HCl – (0.080M)

- Beakers
- HCl - unknown Conc.
- Vinegar
- Phenolphthalein
- Distilled water

Procedure:

Rinse the buret with distilled water. Place in buret clamp on ring stand

- Pour about 10 ml of NaOH into buret and rinse inside surfaces thoroughly. Allow base to run out of buret tip.
- Conditioning the buret
- Fill the buret slightly above the 0.0 mark with 0.1M NaOH. Allow the base to drain until the bottom of the meniscus is past 0.0 ml. Do not stop on zero. Go past it. The buret tip must also be full.
- Standardizing the base solution
- Determine the amount of KHP that would react with 35ml of a 0.1M solution of NaOH. ( the molar ratio of KHP to NaOH is 1:1 in this reaction) (Show this calculation in your Calculation Section)
- Confirm this mass with your teacher
- Add that amount of KHP to an erlenmeyer flask (Record to 0.001g)
- Add ≈50 mL of distilled water
- Add one dropperful of phenolphthalein to the flask of acid
- Place the flask on a sheet of white paper under the buret containing the base solution.
- Swirling gently, begin the titration by adding the base solution to the flask drop by drop. Continue until the faint pink color remains for 30 seconds.
- Note the final volume reading on the scale of the buret. Discard the solution down the sink.

Titrating an Acid of Known Concentration

- Refill the Buret
- Use a graduated cylinder and quantitatively measure 50mL of the 0.080M HCl and add to a 250 ml Erlenmeyer flask.
- Repeat steps 8-11 above

Titrating an unknown solution

- Refill the Buret

- Repeat the titration with 35mL HCl of an unknown concentration. (Each person of your group must do separate unknown acid)
- Repeat steps 8-11 above

Titrating a weak acid

- Refill the Buret
- Repeat the titration with 5mL of vinegar
- Repeat steps 8-11 above

Data:

| | Volume (or mass) of acid | Volume of base initial | Volume of base final |
|---|---|---|---|
| KHP and NaOH | .7785g KHP | .80mL NaOH | 42.41mL NaOH |
| NaOH and HCl (.08M) | 49.5mL HCL | .29mL NaOH | 39.69mL NaOH |
| NaOH and HCl (unknown) | 30.0mL HCl (unknown) | .17mL NaOH | 36.81mL NaOH |
| NaOH and vinegar | 5.15mL vinegar | .33mL NaOH (1$^{st}$ time) .51mL NaOH (2$^{nd}$ time) | 49.65mL NaOH (1$^{st}$ time) 2.60mL NaOH (2$^{nd}$ time) |

Calculations:

1. Find the mass of KHP needed for standardization

35mL of NaOH put into liters find the number of moles in 35mL given the molarity and use the molar ratio and molar mass of KHP to determine the number of grams of KHP needed

$$35mL \times \frac{1L}{1000mL} \times \frac{.1molNaOH}{1L} \times \frac{1molKHP}{1molNaOH} \times \frac{204.221gKHP}{1molKHP} = .715g_{KHP}$$

2. Determine the actual molarity of NaOH solution

.7785g of KHP converted into moles using the molar mass and find the number of moles of NaOH and divide by the volume

$$.7785g \times \frac{1molKHP}{204.221gKHP} \times \frac{1molNaOH}{1molKHP} = .003812mol_{NaOH}$$

$$M = \frac{moles}{volume} = \frac{.003812mol}{(42.41-.80)/1000} = \frac{.003812mol}{.04161L} = .09161M$$

3. Find the molarity of acids using base

Use equation relating molarity and volume. Subtract initial base reading from final and use for volume multiply by molarity of base calculated above (molar ratio, F, is one-to-one, so I will leave this piece out) and divide by the volume of the acid

Known HCl

$$\frac{MV(B)F(A)}{V(A)F(B)} = M$$

$$\frac{(.0916M)\times(39.69-.29mL)}{49.5mL} = .0729M$$

Unknown HCl

$$\frac{MV(B)F(A)}{V(A)F(B)} = M$$

$$\frac{(.09161M)\times(36.81-.29mL)}{30.0mL} = .112M$$

Vinegar

$$\frac{MV(B)F(A)}{V(A)F(B)} = M$$

$$\frac{(.09161M)\times(49.65-.33+2.60-.51mL)}{(5.15mL)} = .914M$$


Error Analysis:

1a. If the buret is dirty, there is a possibility that this substance will affect the results if the buret still has a higher or lower molarity titrant in it previously. If it is a higher molarity, then you will use less titrant than you would have normally, and get a lower molarity for your unknown substance. If it is a lower molarity, you will have to use more titrant, and the unknown solution will be calculated to be higher than in reality. If titrant clings to the sides of the buret, you will have a higher molarity than in actuality, as the volume of the base is greater in calculations than it really is.

1b. If the buret is not conditioned with NaOH beforehand, there will be water left from cleaning the buret, and the molarity of the NaOH will become lower, and will result in a calculated molarity that is too low.

1c. If the buret tip is not filled at the beginning, then the recorded volume is higher than in reality. You will have the wrong value as you are measuring the displacement of NaOH, and not filling the tip will increase the displacement.

1d. Using a different volume of acid than what is written will increase the calculated molarity. Having more acid will mean you also have more moles, but you have a too low volume, so the molarity will be higher, as volume and molarity are inversely proportional.

1e. Adding 50mL of distilled water may or may not affect the molarity. If you use the former volume of the acid, before the 50mL, the results won't be affected. However, if the water is included in the calculations as part of the volume, this will lead to a lower molarity than in reality as volume and and molarity are inversely proportional and the number of moles of acid has not changed.

1f. Losing some of the KHP on the paper will result in a lower number of moles reacting in the titration, and a higher calculated molarity of NaOH. This is because fewer moles of NaOH will be actually needed, and so a lower volume will be needed. The calculations however, will show that with the original mass of KHP, a greater number of moles would have been needed than in reality, and data will show that a lesser volume was used. This will lead to higher calculated molarity of the NaOH.

2. Often, drops of NaOH would cling to the sides, leading to a greater recorded volume than in actuality of the base. As explained in 1a, this would cause a higher calculated molarity as the volume of the base is higher and is directly proportional to the molarity of the acid in the equation MV/F etc. Another error that is like this one is that my group never tapped the buret to get rid of all the air bubbles. Like the other, the volume recorded was not quite what it really was, and a greater volume was needed as part of this volume was air.

3. In our lab, we calculated the NaOH to be ~0.09, which could be due to a slight loss of KHP in the beginning. Having the base have a lower calculated molarity than in actuality would cause the calculated molarity of the acids to be lower as well. Additionally, it is quite possible that the buret was not clean enough and some higher concentration base was left in the buret. As explained in 1a, this would cause you to use less titrant, and have a lower calculated molarity.

Conclusions:

1.     A solution of malonic acid, H2C3H2O4, was standardized by titration with 0.1000M NaOH solution.  If 30.26 mL of the base was required to neutralize completely 14.25 mL of the malonic acid, what was the molarity of the acid solution?

H2C3H2O4 + 2NaOH ☐ Na2C3H2O4 + 2H2O

$$30.26mL \times \frac{1L}{1000mL} \times \frac{.1molNaOH}{1L} \times \frac{1molH2C3H2O4}{2molNaOH} = .001513 \text{mol} H_2C_3H_2O_4$$

$$M = \frac{mol}{L} = \frac{.001513mol}{.01425L} = .1062M \ H_2C_3H_2O_4$$

2.     A solid mixture contains KHP and an inert solid. A 1.00g sample was added to 100mL of water and titrated with a 0.1100M solution of NaOH.  If 30.35mL of the base were needed to neutralize completely the sample, what was the percent KHP in the solid mixture?

$$30.35mL \times \frac{1L}{1000mL} \times \frac{.1100mol}{1L} \times \frac{1molKHP}{1molNaOH} \times \frac{204.221g}{1mol} = .6818g$$

$$\% = \frac{.6818g}{1g} \times 100 = 68.2\%$$