Acid–base back titration

Learning objectives

1. Apply your knowledge of mole calculations to practical experimental data.
2. Write balanced equations for practical experiments.
3. Review practical techniques.

Introduction

In this experiment you will carry out an investigation, calculate the expected mass of product and compare the practical and theoretical data. To complete the investigation, you will write balanced chemical equations, carry out mole calculations, make conclusions and evaluate.

Equipment

Do the experiment individually or as a pair/small group. You will require:

- Two $100\ \mathrm{ml}$ conical flasks
- $25\ \mathrm{ml}$ or $50\ \mathrm{ml}$ pipette
- $10\ \mathrm{ml}$ pipette
- Pipette fillers sized to match pipettes
- Mass balance measuring three decimal places $(0.001\ \mathrm{g})$*
- Burette
- Clamp and retort
- White tile
- Cotton wool
- Safety equipment: safety glasses

*If a mass balance measuring three decimal places is not available then use a mass balance measuring two decimal places $(0.01\ \mathrm{g})$. This will be less accurate though.

Chemicals

- Hydrochloric acid, $1.00\ \mathrm{mol\ dm^{-3}}$
- Calcium carbonate chips (approx $1.00\ \mathrm{g}$)
- Sodium hydroxide, $0.400\ \mathrm{mol\ dm^{-3}}$
- Phenolphthalein indicator solution in dropper bottles

Procedure

1. Prepare tables for your results before you start your experiments.
2. Using the pipette, accurately measure out 50 ml of 1.00 mol dm^{-3} hydrochloric acid and add this to the conical flask. Place a piece of cotton wool in the top of the flask.
3. Measure the mass of the flask, cotton wool and hydrochloric acid. Record this in the table.
4. Using the three decimal places (dp) mass balance, accurately weigh out between 1.000 and 1.500 g (1.00 to 1.50 g for the two dp mass balance) of calcium carbonate and record the mass in the table.
5. Add the calcium carbonate to the hydrochloric acid and swirl the flask until all of the calcium carbonate has reacted.
6. Measure the final mass of the flask and contents and record in the table.
7. Fill the burette with sodium hydroxide solution (0.400 mol dm^{-3}) using a funnel and pouring at your eye level.
8. Using a 10 ml pipette, measure out 10 ml of your reaction mixture into a clean 100 ml conical flask.
9. Add a few drops of phenolphthalein to the flask containing your reaction mixture. The mixture should remain colourless.
10. Titrate your mixture with the sodium hydroxide from your burette, recording how much sodium hydroxide you require to turn the mixture pink.
11. Rinse the conical titration flask with distilled water.
12. Repeat steps 6–11 until you have two results, excluding the rough titration, which are within 0.05 ml of each other (or you run out of reaction mixture).
13. Calculate the average titre to use in the calculations.

Results

	Mass (g)
Flask, cotton wool and hydrochloric acid	
Calcium carbonate	
Total mass of reactants and flask	
Total mass once reaction is complete	
Mass lost	

	Volume (ml)		
	Rough titration	Repeat 1	Repeat 2
Initial			
Final			
Total added			

Calculations

1. Write a balanced equation for the reaction of hydrochloric acid and calcium carbonate.
2. Using your measured mass and volume, calculate the number of moles of hydrochloric acid and calcium carbonate used and whether the hydrochloric acid or calcium carbonate is limiting.
3. Using your answer to question 2, calculate the mass of gas that should form and the moles of acid that should be left over.
4. Write a balanced equation for the reaction of hydrochloric acid and sodium hydroxide.
5. Using your answer to question 4 and your titration data, calculate the moles of hydrochloric acid in your sample (remember this is 10 ml from the total 50 ml)

Conclusion and evaluation

Use your theoretical and experimental data to write a conclusion and evaluation. This should include:

1. How do the theoretical (from step 3) and experimental (from step 5) data for the mass of carbon dioxide produced and the moles of hydrochloric acid in excess compare?
2. Was the actual mass loss greater or less than expected?
3. What is the percentage difference between the theoretical and actual mass lost and is this within the uncertainty of the measurements you have made?
4. Why might any discrepancies have occurred?
5. What are the strengths and limitations of your experiment and what improvements would you make to the procedure or how you carried it out? Explain why.