Acid–base back titration

This resource accompanies **The essential guide to teaching quantitative chemistry** in *Education in Chemistry* which can be viewed at: rsc.li/44KTlPQ. Read the article to discover misconceptions and further ideas for teaching reacting masses and limiting reagents. Use the technician notes to prepare the practical and find the experimental procedure in the student worksheet.

Learning objectives

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

Learners will address all three learning objectives throughout the experiment. They should successfully carry out the practical, write relevant balanced equations and utilise these to carry out mole calculations. Learners will draw their findings together in a conclusion and evaluation, comparing their experimental and theoretical data.

Scaffolding

Some learners will be keen to work through the whole task independently. Others will benefit from you checking the calculation steps as they go along, ensuring they have correctly balanced the equations and found the masses and moles before they write their conclusion and evaluation. Use the PowerPoint to model each step. Remove the example tables in the student sheet to encourage learners to draw their own or leave them in for support.

As an extension, ask learners to titrate the original HCl against the $NaOH$ to check the concentration is correct. They can also investigate the effect of CO_2 dissolving on the pH of water following the procedure at rsc.li/3L9K6Qr.

Safety and hazards

- Read our standard health and safety guidance, available from rsc.li/3IAmFA0 and carry out a risk assessment before running any live practical.
- Eye protection must be worn.
- The flask containing hydrochloric acid and calcium carbonate may get warm.
- Instruct learners to take care not to spill solutions, particularly phenolphthalein, on their skin. If they do get any on their skin, rinse well.
- Fill the burette at eye level.

Results

Example data tables. Masses and volumes will be learner dependent.

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

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

Answers

Presuming mass of $CaCO_3 = 1.10$ g

Note this will vary for each group, but the working will remain the same. Use the accompanying moles calculator spreadsheet to check learners' answers.

1. $2HCl + CaCO_3 \rightarrow CO_2 + H_2O + CaCl_2$
2. Moles of $HCl =$ concentration $\times$ volume

$$= 1.00 \times 0.0500$$

$$= 0.0500 \text{ moles}$$

Formula mass of $CaCO_3 = 40 + 12 + (16 \times 3)$

$$= 100$$

Moles of $CaCO_3 = \frac{\text{mass}}{\text{formula mass}}$

$$= \frac{1.103}{100}$$

$$= 0.0110 \text{ moles}$$

Ratio needed $HCl : CaCO_3$

$2 : 1$

Therefore, you need twice as many moles of HCl compared to $CaCO_3$.

We have more than twice as many moles of HCl, therefore $CaCO_3$ is limiting.

3. From the balanced equation there should be a $1:1$ ratio of $CO_2 : CaCO_3$.

   Therefore, moles of CO_2 = moles of $CaCO_3 = 0.011$ moles

   Mass of gas = moles x formula mass

   $= 0.01103 \times (12 + (16 \times 2)$

   $= 0.485$ g

   Moles of acid left over = $0.05 - 2(0.01103)$

   $= 0.0280$ moles

4. $HCl + NaOH \rightarrow NaCl + H_2O$
5. Presuming an average titre of 12.40 ml (as an exemplar, learners' titres will vary).

   Moles NaOH $=$ concentration $\times$ volume

   $= 0.40 \times \frac{12.40}{1000}$

   $= 0.00496$ moles

   From balanced equation moles HCl = moles NaOH

   Therefore, moles of HCl (in 10 ml) = 0.00496

   Therefore, total leftover moles of HCl = 0.00496×5

   $= 0.0248$ moles

Conclusion and evaluation

Learners will compare their experimental data for the mass of carbon dioxide produced and the moles of hydrochloric acid in excess with theoretical data to write their conclusion and evaluation. The latter should include why they think any discrepancies might have occurred, the strengths and limitations of their experiment, any improvements they would make to the procedure or how they carried it out and why.

Potential limitations might include:

- Carbon dioxide will dissolve in water, affecting the results (see rsc.li/3L9K6Qr). At 20°C 0.168 g/100 g water will dissolve (*Lange's Handbook of Chemistry*, 10th edn, table 5.1) although most of this exists as dissolved gas rather than carbonic acid.
- The final volume may not be 50 ml due to $CaCl_2$ dissolving, this would mean that the 10 ml taken out of the solution would not be one fifth of the total.
- Some of the mass of $CaCO_3$ may be lost on transfer.

Possible improvements might include:

- The mass of $CaCO_3$ can be calculated as mass added (ie, mass of container and solid before addition – mass of container and any remaining solid).
- Reaction carried out at a higher temperature to reduce the amount of CO_2 dissolved, which decreases with temperature. Alternatively, dissolved CO_2 could be calculated (bubble a known volume of CO_2 through water, capturing excess. Difference in initial and final volume will give the dissolved volume).
- Rather than using aliquots for the titration, the entire volume could be used to ensure the volume is as expected. Repetition of the entire reaction would then be needed for repeats.