2023 Preliminary Exam/End of Year Practical Exam answers

1 (a) (i) [5]

Results table:

Records initial burette readings, final burette readings and volume added with correct headings and units in a titration table [1]

	Trial 1	Trial 2
Final burette reading / cm^3	20.90	20.90
Initial burette reading / cm^3	0.00	0.00
Volume of S / cm^3	20.90	20.90
Best titration results (✓)	✓	✓

All burette readings for all accurate titres in titration table are recorded to nearest 0.05 cm^3 [1]

Titration results:
Accuracy for average titre of consistent readings

- within 0.20 cm^3 of supervisor's average value [max 2]
- within 0.30 cm^3 of supervisor's average value [max 1]

	Student's average
+0.30 cm^3 [1]	20.90
+0.20 cm^3 [2]	20.80
Supervisor's average/cm^3 [2]	**20.60**
-0.20 cm^3 [2]	20.40
-0.30 cm^3 [2]	20.30

Concordance:
At least two titre values are within 0.20 cm^3 [1]

(ii) [1]

Calculation of appropriate average volume of **S** in 2 d.p from closest titre values (titres should be identified either in the table or by a tick, or in the calculation)

e.g. Average volume of **S** = $\frac{20.90 + 20.90}{2}$ = 20.90 cm^3

(b) (i) [1]

No. of moles of sodium hydroxide = $\frac{20.90}{1000} \times 0.120$ = 0.00251 mol

No. of moles of sodium hydroxide present
$= \frac{(a)(ii)}{1000} \times 0.120$ [1]

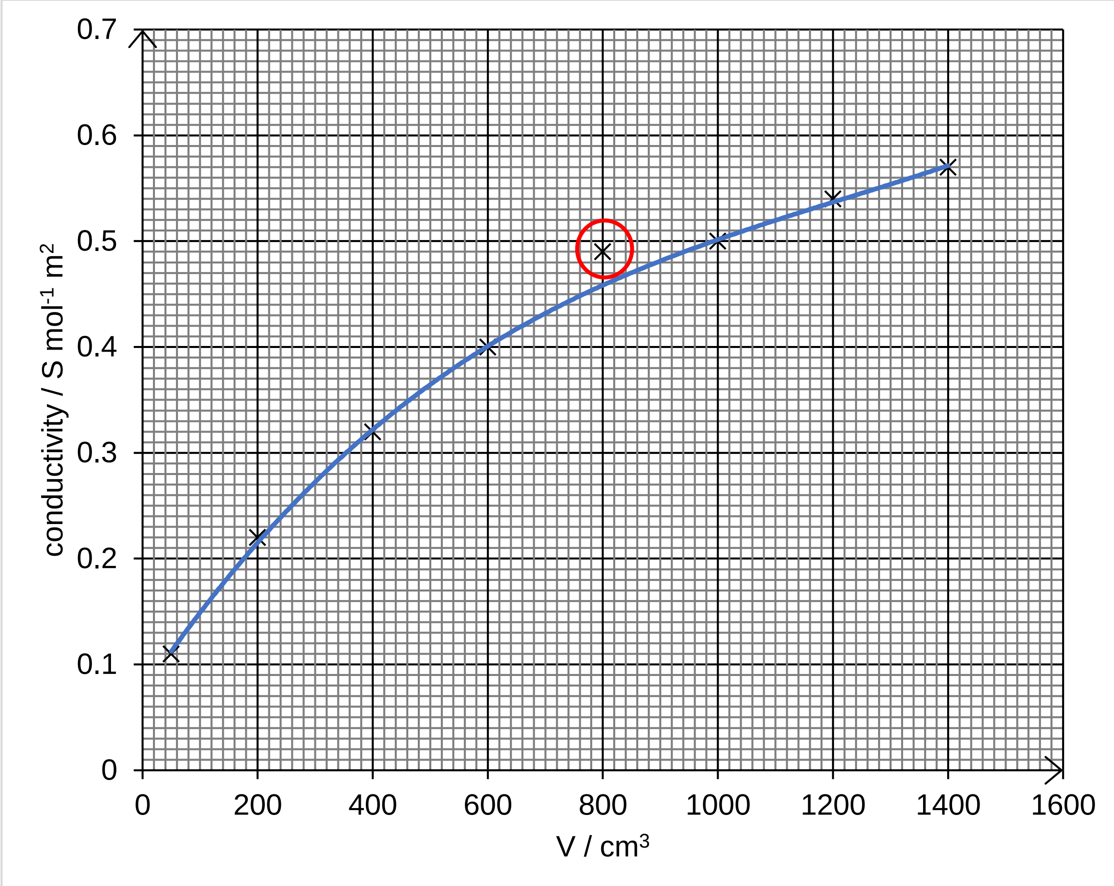
(ii) [1]

No. of moles of hydrochloric acid in 250 cm^3 of solution R
$= 0.00251 \times \frac{250}{25} = 0.0251$ mol

No. of moles of hydrochloric acid in 250 cm^3 of solution R
$= answer\ in\ b(i) \times \frac{250}{25}$ [1]

		(iii)	No. of moles of hydrochloric acid that reacted with Q $= \frac{100}{1000} \times 0.75 - 0.0251 = 0.499$ mol No. of moles of hydrochloric acid that reacted with Q $= \frac{100}{1000} \times 0.75 - b(ii) =$ Calculates initial amount of acid [1] Subtraction of the amount of acid used from the initial amount of acid [1]	2
		(iv)	$Q + 2HCl \rightarrow QCl_2 + H_2$ Allowed: $Mg + 2HCl \rightarrow QCl_2 + H_2$	1
		(v)	No. of moles of Q $= \frac{0.0499}{2} = 0.0250$ mol Relative atomic mass of Q $= \frac{0.60}{0.0250} = 24.0$ No. of moles of Q $= \frac{(iii)}{2}$ [1] Relative atomic mass of Q $= \frac{0.60}{\frac{(iii)}{2}} = \ldots$ to 1 d.p [1]	2
	(c)	As some of the mixture splashed out of the conical flask, some unreacted acid is also lost from the reaction mixture in the conical flask. [1] This would result in a lower amount/volume of sodium hydroxide required to neutralise the remaining acid ***/*** This would result in a greater calculated amount of hydrochloric acid that reacted with the metal. [1] This leads to a greater calculated amount of Q and subsequently a smaller calculated value of the relative atomic mass of Q. [1] • describe **impact** of acid spray on the composition of mixture (in particular, the acid that is unreacted) [1] • how lesser acid present in the conical flask affects the **results of titration –** whether the volume of NaOH used/amount of NaOH calculated or the subsequent **calculations** involving hydrochloric acid reacted or leftover [1] • effect on the **calculated value of the relative atomic mass** of Q [1]		3
			[Total: 16]	

2	**(a)**	**(i)**	Upon adding B, A light brown/cream/off-white precipitate was formed. Upon adding C, a white precipitate was formed and effervescence was observed. The colourless, odourless gas evolved formed a white precipitate in limewater. The gas is carbon dioxide.	
		(ii)	Upon adding A, a white precipitate was formed. Upon adding B, effervescence was observed. The colourless, odourless gas evolved formed a white precipitate in limewater. The gas is carbon dioxide.	
			Marking points for 2(a): • formation of light brown/cream/off-white precipitate [1] • formation of white precipitate for **both** (i) and (ii) [1] • observation for formation of gas and test for gas [1] • identity of gas [1]	4
	(b)	**(i)**	White precipitate formed upon heating	1
		(ii)	Light blue precipitate formed. Light blue precipitate/solid turned black/dark brown upon heating	1
		(iii)	Blue solution turned green/greenish-blue upon heating.	1
			Note from 2022 Examiner's Report as a reference for marking points: • need to differentiate between the colour of solutions and the colour of precipitates • need to identify the gas formed	
	(c)		A – aqueous silver nitrate; C – dilute hydrochloric acid In (a)(i), when C was added, carbon dioxide was produced. Hence C could be hydrochloric acid or sodium carbonate. In (a)(ii), when B was added, carbon dioxide was produced. Hence B could be hydrochloric acid or sodium carbonate too. However, when C was added to A, a white precipitate of silver chloride was formed, suggesting that C is aqueous hydrochloric acid. Hence, A is aqueous silver nitrate. • identity of A **and** C [1] • explanation of how students arrived at their answer [1]	1 1
	(d)		$Na_2CO_3 + CuSO_4 \rightarrow CuCO_3 + Na_2SO_4$	1
	(e)	**(i)**	$Al(OH)_3$ accept formula only	1
		(ii)	It is acidic	1
		(iii)	Cations with +3 charge are acidic, cations with +2 charge are not acidic. • Relate charge to acidity	1
			[Total: 13]	

3	(a)	 correct plotting of points [1] line of best fit, ignore anomalous point (shown in red, but students don't need to circle) [1] labelling of axes [1] suitable scale [1]	4
	(b)	The greater the volume of water used/the greater the dilution, the greater the conductivity. Hence, the ionisation of ethanoic acid increases with increasing volume of water/increasing dilution. • Relationship between volume of water used and conductivity [1] • Relationship between conductivity (and amount of ions) and degree of ionisation [1]	2
	(c)	Tap water may contain traces of ions and this will cause the measured conductivity to increase. OR Pure water does not contain any traces of ions and this will not interfere with the results. OR Ions in tap water may react with ethanoic acid, reducing the amount of ions present. This causes the measured conductivity to decrease. • Connection between presence/absence of ions and how it affects/will not affect the results	1
		[Total: 7]	

4 **Method 1**

1. Measure 100 g of water using a electronic mass balance/measure 100 cm^3 of water using a measuring cylinder. Pour it into a beaker.

2. Add a known mass of solid ammonium chloride (eg. 70 g, or any mass greater than 40 g) to the water.
3. Stir continuously using a glass rod until no more solid can dissolve. (Leave the solution to stand for 30 min)
4. Filter to obtain the undissolved* ammonium chloride / the **residue**
5. Dry between sheets of filter paper.

6. Weigh the undissolved* ammonium chloride using an electronic mass balance.
7. Subtract 70 – mass undissolved = mass dissolved

8. Compare with the value (40) in the table. If it is the same or close, it is correct.

- measure of 100 g water (or a known mass) or 100 cm^3 of water [1]
- dissolve solid (with known mass specified) [1]
- obtain undissolved solid (record mass) [1] *mention 'undissolved'/'insoluble'/'solid' once
- how the results are used [1]

Method 2

1. Measure 100 g of water using a electronic mass balance/measure 100 cm^3 of water using a measuring cylinder. Pour it into a beaker.

2. Add excess solid ammonium chloride and stir continuously / until no more solid can dissolve.
3. Filter to obtain the **filtrate** / ammonium chloride solution*.

4. Measure the mass of ammonium chloride solution* using an electronic mass balance. Record the mass.
5. Subtract the mass of solution by 100 to get the mass of ammonium chloride dissolved.

6. Compare with the value (40) in the table. If it is the same or close, it is correct.

- preparation of water [1]
- dissolve excess solid, conduct filtration [1]
- obtain the mass of aqueous* ammonium chloride [1] *mention 'aqueous'/'solution' once
- how the results are used [1]

Method 3

(but not advisable because the solution would still contain traces of chloride ions even if the maximum mass of silver chloride has been precipitated.)

1. Measure 100 g of water using a electronic mass balance/measure 100 cm^3 of water using a measuring cylinder. Pour it into a beaker.

2. Add excess solid ammonium chloride and stir continuously / until no more solid can dissolve
3. Filter to obtain the **filtrate** / ammonium chloride solution.
4. Add an excess of 0.1 mol/dm^3 aqueous silver nitrate.
5. Filter to collect the precipitate/residue/silver chloride. Wash the residue with distilled water.

	6. Dry between sheets of filter paper. 7. Measure the mass of silver chloride obtained using an electronic mass balance. 8. Calculate the number of moles of silver chloride and equate to the number of moles of ammonium chloride 9. Calculate the mass of ammonium chloride obtained. 10. Compare with the value (40) in the table. If it is the same or close, it is correct. • preparation of water [1] • dissolve solid (with known mass), conduct precipitation method [1] • obtain the mass of silver chloride [1] • how the results are used [1] Allowed: • Weighing scale, weighing balance, mass balance, electronic balance, electrical mass balance, electric weighing mass(?) - (Best to use ‘**electronic mass balance**’) Not accepted: • Collect the filtrate and evaporate to dryness (because ammonium chloride sublimes) • Collect the filtrate and conduct crystallisation (no crystals will appear because the investigation is at room temperature and the solution is saturated)			
			[Total: 4]	