

Exercise walk-through

2.1.1

Calorimetry: sign, amount and uncertainty

Pearson Edexcel · International A-Level · Chemistry

Opposite solubility trends explain the evidence

- 1 **Set up** $\text{Ba}(\text{OH})_2$ is more soluble: its equal saturated volume contains more OH^- and needs more acid.
- 2 **Reason** BaSO_4 is less soluble: equal initial cation amounts give more solid with excess sulfate.
- 3 **Conclude** Solubility controls dissolved amount; it is different from dissociation strength.

Controls distinguish chemical evidence from contamination

- 1 Set up** Match measured amounts, particle size, apparatus and heating rate; compare sample temperature at a consistent decomposition criterion.
- 2 Reason** CO_2 : milky limewater; O_2 : glowing splint relights; NO_2 : brown and acidic.
- 3 Conclude** Clean flame-test wire and keep the flame region/loading consistent. Na yellow, K lilac, Ca brick-red, Sr crimson-red; sodium carry-over can mask another colour.

Independent identification

Your turn

$\text{C}_3\text{H}_6\text{O}$ has a carbonyl IR absorption and gives a positive Fehling's test. Which structure fits all evidence?

- A. CH_3COCH_3
- B. $\text{CH}_3\text{CH}_2\text{CHO}$
- C. $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$
- D. CH_3COOH

Exercises 2.1·2.3

2.1 State standard pressure and the usual stated temperature for standard enthalpy data. [2]

2.2 A reaction warms 100.0 g of solution by 4.0 °C. Use $c = 4.18 \text{ J g}^{-1} \text{ °C}^{-1}$. Calculate

2.3 The reaction in the previous item forms 0.0300 mol of water. Calculate the molar enthalpy of neutralisation. [2]

3 Exam-style questions

Exercises 2.1·2.3

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2.1 100 kPa, usually 298 K. Temperature must be specified; standard does not mean room conditions chosen without a value.

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2.2 $q = mc\Delta T = 100.0 \times 4.18 \times 4.0 = 1672 \text{ J}$, or 1.67 kJ.

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3 Exam-style questions

2.3 $\Delta H = -q/n = -1.672 \text{ kJ}/0.0300 \text{ mol} = -55.7 \text{ kJ mol}^{-1}$.

Exercises 3.1·3.2

3.1 0.460 g of ethanol ($M = 46.0 \text{ g mol}^{-1}$) burns and warms 200 g water by $12.0 \text{ }^{\circ}\text{C}$. Use $c = 4.18 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1}$. Calculate the experimental combustion enthalpy, then explain why it can be less negative than the standard value. [5]

3.2 A thermometer has $\pm 0.1 \text{ }^{\circ}\text{C}$ uncertainty per reading. A measured rise is $2.0 \text{ }^{\circ}\text{C}$. Estimate the maximum percentage uncertainty in the rise and explain one way to lower it without claiming repeats correct the thermometer. [3]

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3.1 $q = mc\Delta T = 200 \times 4.18 \times 12.0 = 10032 \text{ J} = 10.032 \text{ kJ}$. $n = m/M = 0.460/46.0 = 0.0100 \text{ mol}$. $\Delta H = -q/n = -1003 \text{ kJ mol}^{-1}$. Heat loss, incomplete burning or fuel evaporation makes energy received per measured mole too small.

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3.2 Two readings give $\pm 0.2 \text{ }^{\circ}\text{C}$; $100(0.2/2.0) = 10\%$. A safely larger rise reduces relative reading uncertainty, or use a more precise thermometer.

Exercise 3.3

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3.3 Formation: one mole of compound; combustion: one mole of substance completely burned; neutralisation: one mole of water formed; atomisation: one mole of gaseous atoms formed. $\frac{1}{2}\text{Cl}_2(\text{g}) \rightarrow \text{Cl}(\text{g})$. Each process is under specified standard conditions.