Page 150–151 Activity: Measuring the enthalpy change for the reaction of zinc with copper sulfate solution

1 The graph should have:

- axes with scales and labels
- points plotted accurately
- a clean, smooth curve of best fit including the extrapolation.

2 **a)** 36.4 ± 0.2°C

b) $\Delta T = (36.4 \pm 0.2) - 24.1 = 12.3 \pm 0.2$°C

3 Energy given out = 50 g × 4.2 J g^{-1} K^{-1} × 12.3 K = 2583 J

4 **a)** $\frac{50}{1000}$ dm^3 × 0.25 mol dm^{-3} = 0.0125 mol

b) 0.0125 mol

5 0.0125 mol of Zn and $CuSO_4$ give out 2583 J.

1 mol of Zn and $CuSO_4$ give out $\frac{2583}{0.0125}$ J

= 206 640 J

Therefore $\Delta H_{reaction}$ = –207 kJ mol^{-1}

6 Heat loss to surroundings and to container, accuracy of thermometer

7 Insulation, lid on container, more accurate thermometer.

Pages 161–164 Practice questions

1 B [1]

2 D [1]

3 B [1]

4 B [1]

5 D [1]

6 A [1]

7 C [1]

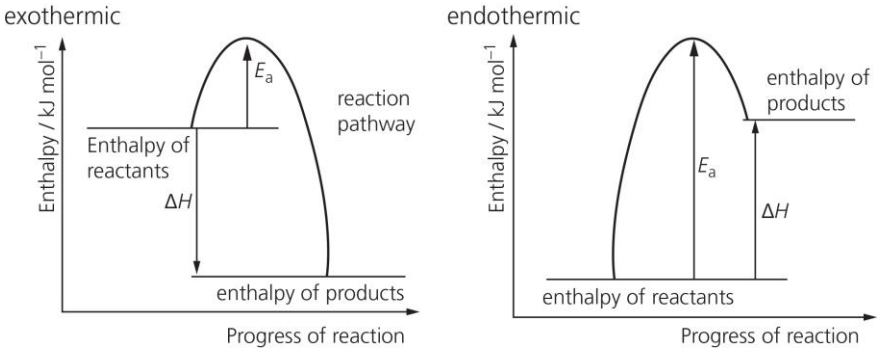
8 B [1]

9 A [1]

10 A [1]

11 In an exothermic reaction, energy is transferred from the reaction mixture to the surroundings. Chemical energy is released from the reactants and the temperature of the surroundings *increases*. An exothermic reaction is indicated by a negative Δ*H* value.

In an endothermic reaction, heat energy is transferred *from* the surroundings (or from an external source) to the reaction mixture. Endothermic reactions cause the temperature of the surroundings to decrease. Δ*H* is given a positive value.



[2 marks for each diagram; 1 for position of reactants and products, 1 for labels]

12 a) Bond enthalpy is the enthalpy change required to break 1 mol of bonds in the molecules of a gas so that the resulting gaseous (neutral) particles/atoms/radicals separate and exert no forces upon each other. [2]

For example: Cl–Cl(g) → Cl·(g) + Cl·(g) or generally: X–Y(g) → X·(g) + Y·(g) [1]

b) $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$

Bond enthalpy of reactants = 436 + 193 = 629 kJ mol^{-1} [1]

Bond enthalpies of products: 2(366) = 732 kJ mol^{-1} [1]

ΔH = sum of bond enthalpies of reactants – sum of bond enthalpies of products

629 – 732 = –103 kJ mol^{-1} [1]

13 Energy change, $q = mc\Delta T$

The mass of the reaction mixture is 48 + 2 = 50 g assuming density is 1.0 g cm^{-3}.

So q = 50 × 4.18 × 6.7 = 1400.3J = –1.4 kJ [1 for calculation, 1 for '–' sign]

14 a) Standard temperature is 298 K (25 °C). [1]

b) Standard pressure is 100 kPa. [1]

15 a) Standard enthalpy change of formation, $\Delta H_f^{\ominus}$, is the enthalpy change when 1 mol of a substance is formed from its elements, in their natural states, under standard conditions of 298 K [1] and 101 kPa [1].

b) $3C(s) + 3H_2(g) + \frac{1}{2}O_2(g) \rightarrow CH_3CH_2CHO(l)$ [1]

(It is essential that the equation shows that only 1 mol of propanal is formed, and it is acceptable to use fractions. Your answer will be marked wrong if you multiply by 2 to remove the $\frac{1}{2}$)

16 a) Standard enthalpy change of combustion, $\Delta H_c^{\ominus}$, is the enthalpy change when 1 mol of a substance is burned completely, in an excess of oxygen, under standard conditions of and 101 kPa. [2]

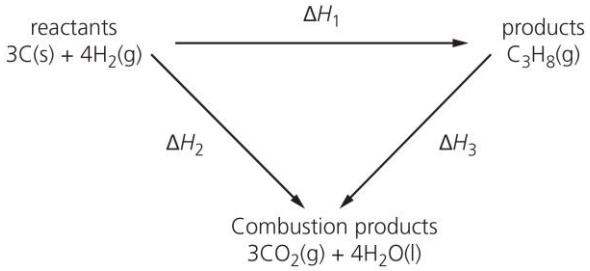
b) $CH_3COCH_3(l) + 4O_2 \rightarrow 3CO_2(g) + 3H_2O(l)$ [1]

(When balancing the equation, do not forget that there is an O in CH_3COCH_3.)

17 Step 1: Write an equation for the enthalpy of formation of propane:

$3C(s) + 4H_2(g) \rightarrow C_3H_8(g)$ [1]

Step 2: Construct an enthalpy cycle by writing the combustion products at the bottom. Both arrows point downwards.



$\Delta H_1 + \Delta H_3 = \Delta H_2$ hence $\Delta H_1 = \Delta H_2 - \Delta H_3$

ΔH_2 = 3 × [combustion of C(s)] + 4 × [combustion of $H_2(g)$]

= (3 × −394) + (4 × −286) [1]

= (−1182) + (−1144) = −2326 kJ mol^{-1} [1]

Therefore, ΔH_1 = −2326 − (−2219) = −107 kJ mol^{-1}

$3C(s) + 4H_2(g) \rightarrow C_3H_8(g)$ $\Delta H_f^{\ominus}$ = −107 kJ mol^{-1} [1]

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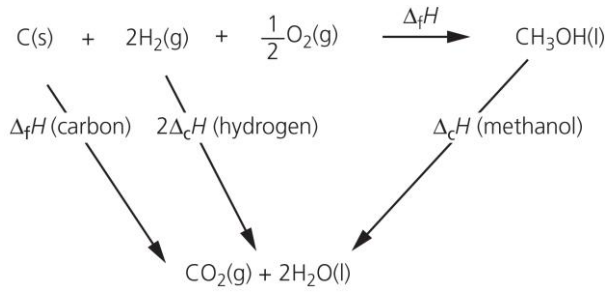
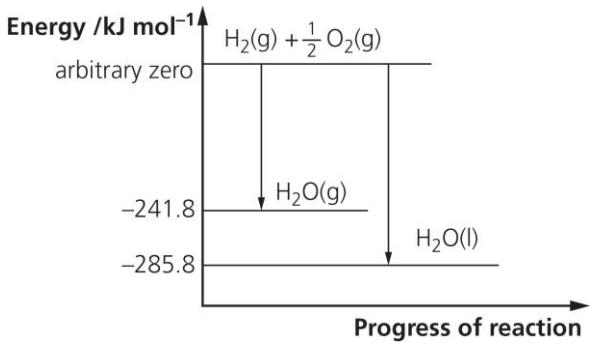


[1]

$\Delta_f H$ – 726 = (–394) + 2(–286) [1]

$\Delta_f H$ = –240 kJ mol^{-1} [1]

19 a)



[1 mark each for the positions of $H_2O(g)$ and $H_2O(l)$; 1 mark for giving the ΔH values]

b) The enthalpy change for $H_2O(l) \rightarrow H_2O(g)$ from the enthalpy profile diagram is +44 kJ. [1]

20 $C_4H_{10}(g) + 6\frac{1}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$

Bond enthalpy of reactants = 3(347) + 10(413) + 6½(498) = 8408 kJ [1]

Bond enthalpies of products = 8(805) + 10(464) = 11 080 kJ [1]

ΔH = sum of bond enthalpies of reactant – sum of bond enthalpies of products

Enthalpy of reaction = 8408 – 11 080 = –2672 kJ mol^{-1} [1]

21 a) $C_2H_6(g) + Br_2(l) \rightarrow C_2H_5Br(l) + HBr(g)$

Bond enthalpy of reactants = 347 + 6 × 413 + 193 = 3018 kJ [1]

Bond enthalpies of products = 347 + 5 × 413 + 290 + 366 = 3068 kJ [1]

ΔH = sum of bond enthalpies of reactant – sum of bond enthalpies of products

3018 – 3068 = –50 kJ [1]

$C_2H_6(g) + Br_2(l) \rightarrow 2CH_3Br(g)$

Bond enthalpy of reactants = 347 + 6 × 413 + 193 = 3018 kJ [1]

Bond enthalpies of products = 2(3 × 413 + 290) = 3058 kJ [1]

ΔH = sum of bond enthalpies of reactant – sum of bond enthalpies of products

3018 – 3058 = –40 kJ [1]

b) Reaction 1 is more exothermic than reaction 2, so overall could be more likely to occur [1], although since both reactions are exothermic the actual pathway of the reaction might affect what actually happens [1].

c) The bond enthalpies used are average figures and will not be quite the same as the experimentally determined values. [1]

Mean bond enthalpies are calculated assuming substances are gases whereas bromine in both reactions and C_2H_5Br in reaction 1 are liquids. [1]

22 Energy change, q = mcΔT

Mass = 25 + 50 = 75g (this is the mass of the two solutions when mixed assuming density is 1.0 g cm^{-3} [1]

So q = 75 × 4.18 × 4.6 = 1442.1J = 1.4421 kJ [1]

There are equimolar amounts of both reactants, moles of NaOH = $(\frac{50}{1000})$ × 0.5 = 0.025 mol [1]

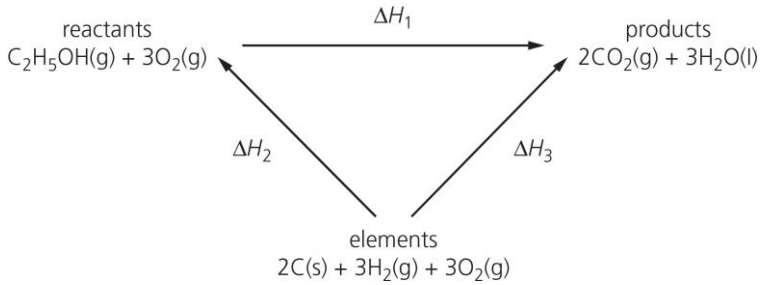
Moles of nitric acid = $(\frac{25}{1000})$ × 1.0 = 0.025 mol [1]

Therefore $\Delta H = \frac{1.4421}{0.025}$ = –57.7 kJ mol^{-1} [1 mark for answer, 1 mark for '–' sign]

23 Step 1: Write an equation for the enthalpy of combustion required:

$C_2H_5OH(g) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$ [1]

Step 2: Construct an enthalpy cycle by writing the elements at the bottom. Both arrows point upwards.



$\Delta H_1 + \Delta H_2 = \Delta H_3$ hence $\Delta H_1 = \Delta H_3 - \Delta H_2$

ΔH_2 = formation of $C_2H_5OH(g)$ = −277 kJ mol^{-1}

ΔH_3 = 2 × [formation of $CO_2(g)$] + 3 × [formation of $H_2O(l)$]

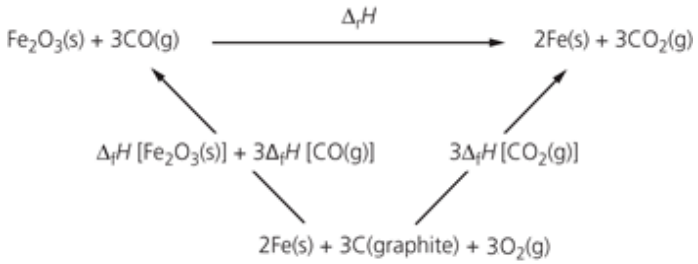
= (2 × −394) + (3 × −286) [1]

= (−788) + (−858) = −1646 kJ mol^{-1} [1]

Therefore, ΔH_1 = (−1646) − (−277) = −1369 kJ mol^{-1}

$C_2H_5OH(g) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$ ΔH = −1369 kJ mol^{-1} [1]

24 $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$ [1]

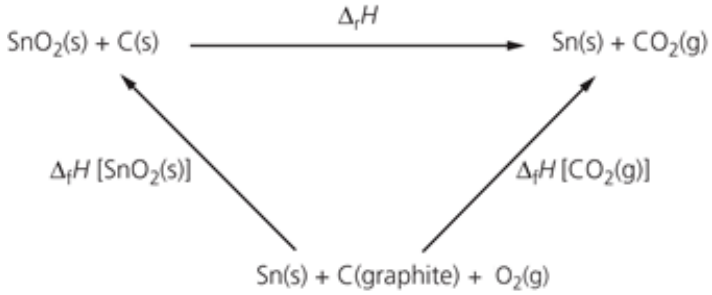


$\Delta_r H$ + (–824) + 3(–110) = 3(–394) [1]

$\Delta_r H$ – 1154 = –1182 [1]

$\Delta_r H$ = –28 kJ mol^{-1} [1]

25 a) Reaction 1:

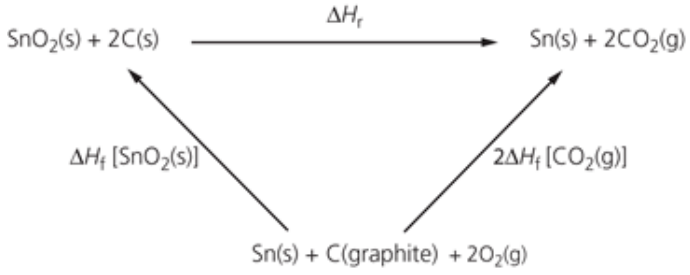


$\Delta_r H + \Delta_f H$ [$SnO_2(s)$] = ΔH_f [$CO_2(g)$] [1]

$\Delta_r H$ – 581 = – 394 [1]

$\Delta_r H$ = +187 mol^{-1} [1]

Reaction 2:



$\Delta_r H + \Delta_f H$ [$SnO_2(s)$] = $2\Delta H_f$ [CO(g)] [1]

$\Delta_r H$ – 581 = 2 × (–110) [1]

$\Delta_r H$ = + 361 kJ [1]

b) Both reactions are endothermic. However the first reduction producing carbon dioxide is less endothermic (+187 kJ) than the process forming carbon monoxide (+361 kJ). [1]

This would suggest that it would require less energy to carry out the first process which also uses less carbon. [1]

However the calculations use standard enthalpies of formation and the situation might be different at the high temperatures required in the furnace to manufacture the tin. [1]

26 $q = mc\Delta T = 100 \times 4.18 \times 6.97 = 2913.46$ J = 2.913 kJ [1]

$\Delta H = \frac{q}{n}$; q = −2.913 kJ and ΔH = −726 kJ mol^{-1} [1]

Hence, $n = \frac{q}{\Delta H} = \frac{-2.913}{-726} = 4.013 \times 10^{-3}$ mol [1]

mass, m, of alcohol burned = 117.416 − 117.288 = 0.128 g [1]

Hence, molar mass of the alcohol, $M = \frac{m}{n} = \frac{0.128}{4.013 \times 10^{-3}} = 31.90 = 32$ g mol^{-1} [1]

The general formula for an alcohol is $C_nH_{2n+1}OH$, and the mass of OH is 17. It follows that the mass of C_nH_{2n+1} = 32 − 17 = 15, hence $C_nH_{2n+1} = CH_3$. The alcohol is methanol, CH_3OH. [1]

27 a) Total mass of the mixed solutions is 50 + 100 = 150 g [1]

Energy change, $q = mc\Delta T = 150 \times 4.18 \times 9 = 5643$ J = 5.643 kJ [1]

The sodium hydroxide is in excess so that the temperature rise is controlled by the moles of the hydrochloric acid. [1]

The amount in moles of HCl, $(n) = cV = 2.0 \times \frac{50}{1000} = 0.10$ mol [1]

$\Delta H = \frac{q}{n} = \frac{5.643}{0.10} = -56.4$ kJ mol^{-1} [1 mark for answer, 1 for sign]

(The minus sign indicates the reaction is exothermic)

b) The temperature rise will be the same (9 °C). [1]

The amount that has reacted has not changed since the sodium hydroxide was in excess in the original experiment. [1]

The amount in moles of each reagent is now the same and the heat is distributed over the same total volume of solution. [1]

c) The amount that has reacted is still the same. [1]

The heat evolved is dispersed over only 100 cm^3 of solution instead of 150 cm^3. [1]

This means the temperature rise will now be $\frac{150}{100} \times 9 = 13.5$°C. [1]

28 Reaction 1:

Heat released, $q = mc\Delta T = 100 \times 4.18 \times 7.5 = 3135$ J = 3.135 kJ [1]

The M_r of $MgSO_4$ = 24.3 + 32.1 + 64.0 = 120.4 g mol^{-1}

Amount in mol, n, used = $\frac{m}{M} = \frac{8.11}{120.4} = 0.0674$ [1]

$\Delta H = \frac{q}{n} = \frac{3.135}{0.0674} = -46.5$ kJ mol^{-1} [1]

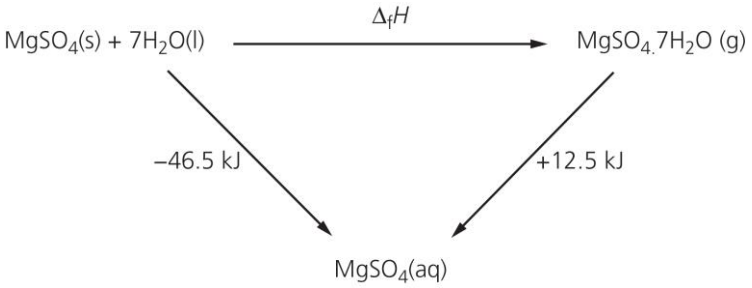
Reaction 2:

Heat absorbed, $q = mc\Delta T = 100 \times 4.18 \times 1.0 = 418$ J = 0.418 kJ [1]

The M_r of $MgSO_4.7H_2O$ = 24.3 + 32.1 + 64.0 + (7 × 18) = 246.4 g mol^{-1}

Amount in mol, n, used = $\frac{m}{M} = \frac{8.23}{246.4}$ = 0.0334 [1]

$\Delta H = \frac{q}{n} = \frac{0.42}{0.0334}$ = +12.5 kJ mol^{-1} [1]



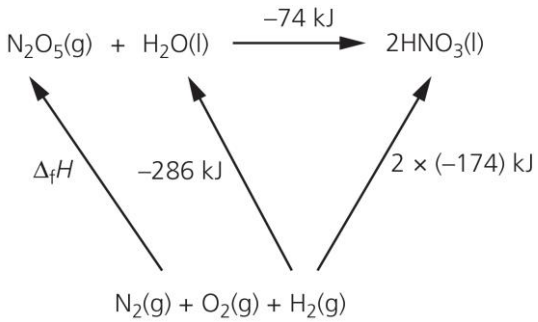
$\Delta_r H$ + 12.6 = –46.8

$\Delta_r H$ = – 59.3 kJ mol^{-1} (accept 59.4 or 60 depending on rounding) [1]

29 a) $N_2(g) + 2\frac{1}{2}O_2(g) \rightarrow N_2O_5(g)$ [1]

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ [1]

b)



[1 mark, plus 1 mark for cycle]

$\Delta_f H$ –286 –74 = 2 × (–174)

$\Delta_f H$ = –348 + 286 + 74 = +12 kJ mol^{-1} [1]

30 a) $2N_2H_4 + N_2O_4 \rightarrow 3N_2 + 4H_2O$ [1]

b) $(CH_3)_2NNH_2 + 2N_2O_4 \rightarrow 3N_2 + 4H_2O + 2CO_2$ [1]

c) The Aerozine 50 mixture contains 50% of each component, so 2200 kg contains 1100 kg of hydrazine and 1100 kg of 1,1-dimethylhydrazine. [1]

M_r of N_2H_4 = 32, so amount in moles of 1100 kg of N_2H_4 = $\frac{1\,100\,000}{32}$ = 34 375 mol [1]

The equation in (**a**) shows that this requires $\frac{34\,375}{2}$ = 17 185.5 mol of N_2O_4 for complete reaction. [1]

M_r of N_2O_4 = 92, so mass of N_2O_4 = $\frac{17\,185.5 \times 92}{1000}$ = 1581.3 kg [1]

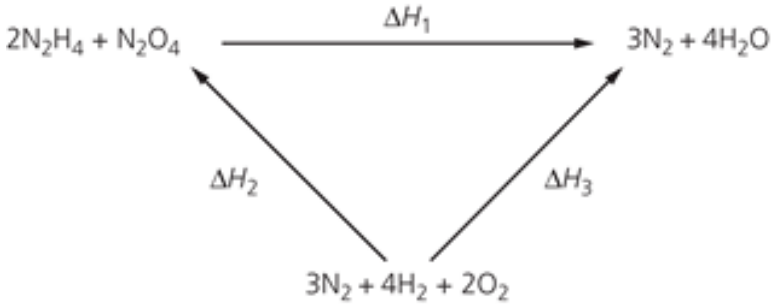
M_r of $(CH_3)_2NNH_2$ = 60, so amount in moles of 1100 kg of $(CH_3)_2NNH_2$ = $\frac{1\,100\,000}{60}$ = 18 333.3 mol [1]

The equation in (**b**) shows that this requires 18 333.3 × 2 = 36 666.7 mol of N_2O_4 for complete reaction. [1]

M_r of N_2O_4 = 92, so mass of N_2O_4 = $\frac{36\,666.7 \times 92}{1000}$ = 3373.3 kg [1]

Therefore, the lunar module would require 1581.3 + 3373.3 = 4594.6 kg (accept 4.59 x10^3 kg) of N_2O_4 for take-off. [1]

d) The enthalpy cycle for hydrazine is as follows:



[1]

ΔH_2 = (2 × 50.6) + (−19.5) = 81.7 kJ mol^{-1} [1]

ΔH_3 = 4 × (−241.8) = −967.2 kJ mol^{-1} [1]

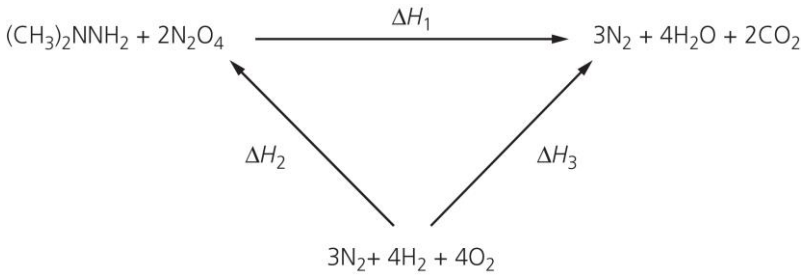
$\Delta H_1 + \Delta H_2 = \Delta H_3$

So $\Delta H_1 = \Delta H_3 - \Delta H_2$

Therefore, ΔH_1 = −967.2 − 81.7 = −1048.9 kJ mol^{-1} [1]

This is for the combustion of 2 mol of hydrazine, which is 64 g. For 1100 kg of hydrazine the enthalpy change is therefore $\frac{1100 \times 1000}{64}$ × 1048.9 = −1.80 × 10^7 kJ [1]

The enthalpy cycle for dimethylhydrazine is as follows:



[1]

ΔH_2 = 49.3 − (2 × 19.5) = 10.3 kJ mol^{-1} [1]

ΔH_3 = [4 × (−241.8)] + [2 × (−393.7)] = −1754.6 kJ mol^{-1} [1]

$\Delta H_1 + \Delta H_2 = \Delta H_3$

So $\Delta H_1 = \Delta H_3 - \Delta H_2$

Therefore, ΔH_1 = −1754.6 − 10.3 = −1764.9 kJ mol^{-1} [1]

This is for the combustion of 1 mol of 1,1−dimethylhydrazine, which is 60 g. For 1100 kg of hydrazine the enthalpy change is therefore $\frac{1100 \times 1000}{60}$ × 1764.9 = −3.24 × 10^7 kJ [1]

Total enthalpy change = −5.04 × 10^7 kJ [1]

e) Use $q = mc\Delta T$, rearrange to $\Delta T = q/mc$

So $\Delta T = \frac{1000 \times 5.04 \times 10^7}{350\,000\,000 \times 4.2}$ = 34.3 K (accept 34 K because using 4.2 J g^{-1} K^{-1} for heat capacity of water) [1]