ENERGETICS Core (SL & HL)

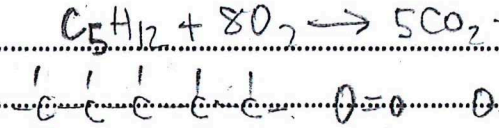
1. An equation for the combustion of pentane is given below:

$$C_5H_{12}(l) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$$

(a) Determine the standard enthalpy change, $\Delta H^\ominus$, for this reaction using **section 11** of the data booklet. **Show your working.**

[3]

$C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$



BROKEN: 4 × C–C; 12 × C–H; 8 × O=O

MADE: 10 × C=O; 12 × H–O

BROKEN = (4×346) + (12×414) + (8×498) = 10336 ✓

MADE = (10×804) + (12×463) = 13596 ✓

BROKEN – MADE = 10336 – 13596 = –3260 (kJ mol^{-1}) ✓ allow ecf.

(b) Calculate the standard enthalpy change, $\Delta H^\ominus$, for this reaction using **section 12** of the data booklet. **Show your working.**

[3]

$C_5H_{12} + 8O_2 \rightarrow (5)CO_2 + (6)H_2O$

ELEMENTS

$\Delta H_f(CO_2) = -393.5$

$\Delta H_f(H_2O) = -285.8$

$\Delta H_f(C_5H_{12}) = -173.$

→ from data booklet

$\Delta H = -(-173) + 5(-393.5) + 6(-285.8)$ ✓ ✓

$= 173 - 1967.5 - 1714.8$

$= -3509.3$ ✓ allow ecf

(c) State and briefly explain whether the method in (a) or the method in (b) above is likely to be the most accurate determination of $\Delta H^\ominus$, for this reaction.

[1]

method (b) AND because bond enthalpy values are measured in gaseous state (and pentane is liquid) and are averages. (or) (both not required.)

2. Copper has a relatively low specific heat capacity. A 50.0g sample of copper rises in temperature by 52.0°C when it absorbs 1000J of energy.

(a) Determine the specific heat capacity of copper in J g^{-1} K^{-1} using section 1 of the data booklet. Give your answer to three significant figures.

[2]

$q = mc\Delta T$ $\quad c = \frac{q}{m.\Delta T} = \frac{1000}{50 \times 52} = \frac{1000}{2600}$ ✓

$c = 0.3846153$

$= 0.385$ ($J g^{-1} K^{-1}$) ✓ 3 sig figs

allow ecf

(b) 0.840g of iron powder was added to 40.0cm^3 of copper sulphate solution in a calorimeter. The copper sulphate was in excess. The maximum temperature rise of the solution was 15.0°C.

$$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$$

(i) Assuming that the heat released was absorbed only by the solution, calculate the enthalpy change, ΔH, for this reaction. Use sections 1 and 2 of the data booklet.

[3]

$q = mc\Delta T = 40.0 \times 4.18 \times 15.0 = 2508$ J ✓

moles of iron $= \frac{0.840}{55.85} = 0.0150402$ ✓

$\Delta H = -\frac{2508}{0.0150402} = -166752$ J mol^{-1} ✓

(exothermic) $= -167$ kJ mol^{-1}

(ii) State another assumption that you made in (b)(i).

[1]

density of solution is 1.00 g cm^{-3} or

specific heat capacity same as water or

reaction goes to completion or

iron does not react with other substances/water

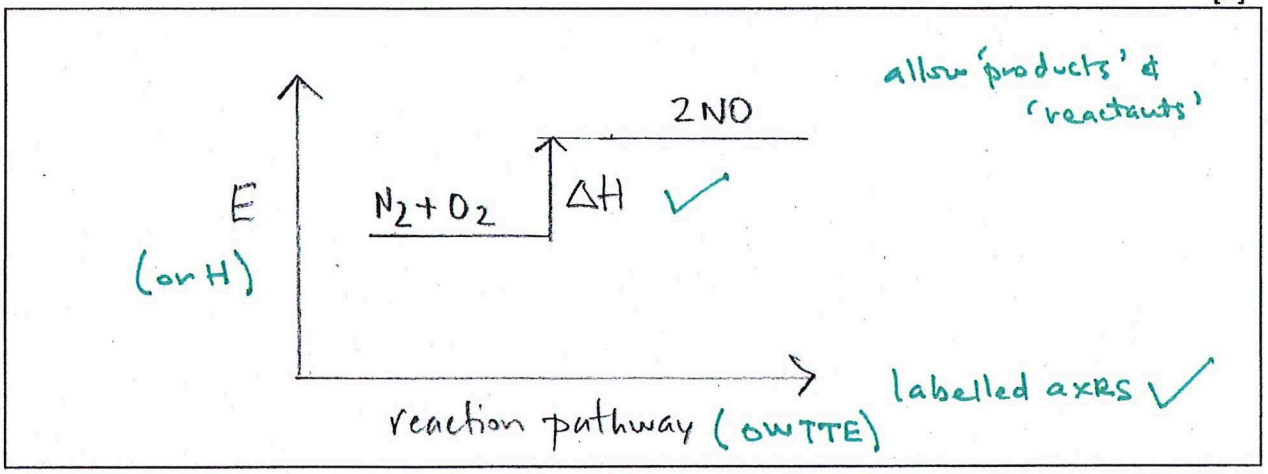
any one ✓

3. (a) The reaction below is an endothermic reaction.

$$N_2(g) + O_2(g) \rightarrow 2NO(g) \qquad \Delta H = +181kJ$$

(i) Sketch a labelled potential energy profile for this reaction, label the enthalpy change, ΔH.

[2]



(ii) Given that the enthalpy change, ΔH, for the reaction as shown above in (a) is +181kJ, use section 11 of the data booklet to calculate the bond enthalpy of the bond in NO(g).

[2]

$N_2 + O_2 \rightarrow 2NO$

$N \equiv N \quad O=O$

ΔH = BROKEN − MADE $\quad +181 = (945+498) - (2NO)$ ✓

$2NO = 1443 - 181$

$= 1262$ allow ecf

Bond Enthalpy NO = 631 ($kJ\,mol^{-1}$) ✓

4. Sulfuric acid is produced in the Contact Process. The first two steps are the reactions:

I	$S(s) + O_2(g) \rightarrow SO_2(g)$	$\Delta H^\ominus = -297kJ$
II	$SO_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons SO_3(g)$	$\Delta H^\ominus = -98kJ$

(a) SO_3 is a solid at just below room temperature. If $SO_3(s)$ was the product in reaction II, instead of $SO_3(g)$, would the ΔH for reaction II be more or less negative? Explain your answer.

[2]

gas to solid is an exothermic process ✓, therefore ΔH would be more negative ✓

(b) Write the equation for the standard enthalpy of formation of $SO_3(g)$.

[2]

$S(s) + 1\frac{1}{2}O_2(g) \rightarrow SO_3(g)$ equation ✓

state symbols ✓

allow $\frac{1}{8}S_8(s)$ (or multiples)

(c) Using the $\Delta H^\ominus$ values given for reactions I and II above, calculate the $\Delta H^\ominus$ for the standard enthalpy of formation of $SO_3(g)$.

[2]

$S(s) + 1\frac{1}{2}O_2(g) \rightarrow SO_3(g)$

(−297) (−98)

$SO_2 + \frac{1}{2}O_2$

$\Delta H = -297 + -98$ ✓

$= -395 \text{ kJ mol}^{-1}$ ✓ allow ecf.

5. Given the enthalpy changes below, and average bond enthalpy for Cl–Cl bond of +242kJ mol^{-1}, calculate the average bond enthalpy for the C–Cl bond. **Show your working.** *

$C(s) + 2Cl_2(g) \rightarrow CCl_4(g)$ $\Delta H^\ominus = -96 \text{kJ mol}^{-1}$

$C(s) \rightarrow C(g)$ $\Delta H^\ominus = +715 \text{kJ mol}^{-1}$

[3]

$CCl_4(g) \rightarrow C(g) + 4Cl(g)$ ← 4 × bond enthalpy C–Cl

(−96) (+715) 2(+242)

$C(s) + 2Cl_2(g) \rightarrow C(g) + 2Cl_2(g)$

$4 \times C\text{–}Cl = -(-96) + 715 + 2(+242)$ ✓ 324 with working scores (3)

$= +1295$ ✓ 323.75 scores (3)

$1 \times C\text{–}Cl = 323.75 \text{ kJ mol}^{-1}$ ✓ * (it's in data booklet!)

$= 324$ (3 sig figs) 324 alone scores zero

Total Marks 26 (39 minutes)