Answers to exam-style questions

Topic 5

1 B

2 C

3 B

4 D

5 C

6 B

7 D

8 D

9 B

10 A

11 **a** The average amount of energy required to break one mole of covalent bonds in a gaseous molecule under standard conditions. 'Average' refers to the fact that the bond enthalpy is different in different molecules and therefore the value quoted is the average amount of energy to break a particular bond in a range of molecules. **[2]**

b

Bond broken	Bond energy / $kJ\,mol^{-1}$	No. bonds	Total energy / $kJ\,mol^{-1}$
C–H	412	5	2060
C–C	348	1	348
C–O	360	1	360
O–H	463	1	463
O=O	496	3	1488
Total energy to break all bonds			4719

Bond made	Bond energy / $kJ\,mol^{-1}$	No. bonds	Total energy / $kJ\,mol^{-1}$
C=O	743	4	2972
O–H	463	6	2778
Total energy released when bonds made			5750

$\Delta H_r = \Sigma$(bonds broken) $- \Sigma$(bonds made)
$\Delta H_r = 4719 - 5750 = -1031\,kJ\,mol^{-1}$ **[3]**

c Two of the substances involved are in the liquid state, but bond enthalpies can be used only when all substances are in the gaseous state. When substances are not in the liquid state, energy must also be supplied to overcome intermolecular forces, or energy is released when intermolecular forces are formed. **[2]**

d
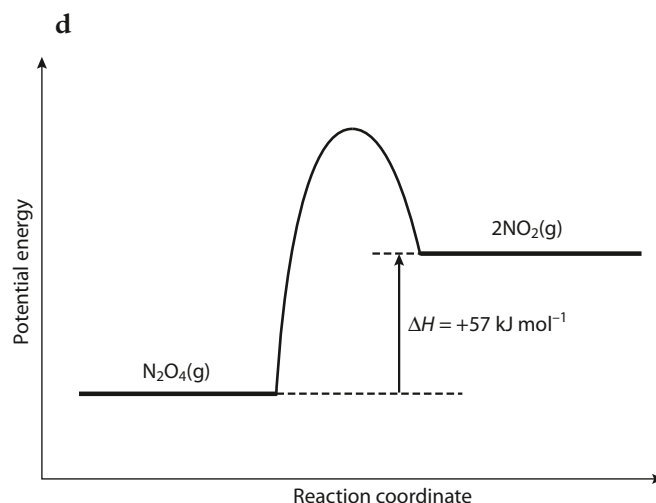
N$_2$O$_4$ is more stable because it is at a lower energy level. **[3]**

12 **a** The enthalpy change when one mole of a substance is formed from its elements in their standard states under standard conditions (100 kPa pressure). **[2]**

b $4H_2(g) + 3C(s) + \frac{1}{2}O_2(g) \rightarrow C_3H_7OH(l)$ **[2]**
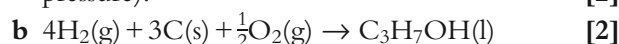
c $\Delta H = \Sigma \Delta H_f \text{ (products)} - \Sigma \Delta H_f \text{ (reactants)}$
$= [(3 \times -394) + (4 \times -286)] - [-316]$
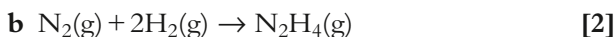$= (-1182 + -1144) + 316$
$= -2010\,kJ\,mol^{-1}$ **[3]**

d The equation for the complete combustion of propan–1–ol is:
$C_3H_7OH(l) + 4\frac{1}{2}O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$
To calculate the entropy change we use:
$\Delta S = \Sigma S\text{(products)} - \Sigma S\text{(reactants)}$
$\Delta S^{\ominus} = [(3 \times 214) + (4 \times 69.9)] - [(1 \times 196.6) + (4.5 \times 205)]$
$\Delta S^{\ominus} = -197.5\,J\,K^{-1}\,mol^{-1}$
The sign of ΔS is negative because the reaction involves a decrease in the number of moles of gas (4.5 on the left–hand side but only 3 on the right–hand side) and therefore the energy is less spread out in the products. **[3]**

e For the complete combustion of propan–1–ol:
$\Delta H^{\ominus} = -2010\,kJ\,mol^{-1}$
$\Delta S^{\ominus} = -197.5\,J\,K^{-1}\,mol^{-1}$
At 298 K:
$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$
$\Delta G^{\ominus} = -2010 - 298 \times \left(\frac{-197.5}{1000}\right)$
Therefore, $\Delta G^{\ominus} = -1951\,kJ\,mol^{-1}$. $\Delta G^{\ominus}$ is negative, so the reaction is spontaneous at 25 °C. **[3]**

13 a

$$\begin{array}{cc} H & H \\ \ddot{N}\!:\!\ddot{N} \\ H & H \end{array}$$

[1]

b $N_2(g) + 2H_2(g) \rightarrow N_2H_4(g)$ **[2]**

c

Bond broken	Bond energy / kJ mol^{-1}	No. bonds	Total energy / kJ mol^{-1}
N≡N	944	1	944
H–H	436	2	872
Total energy to break all bonds			1816

Bond made	Bond energy / kJ mol^{-1}	No. bonds	Total energy / kJ mol^{-1}
N–N	163	1	163
N–H	388	4	1552
Total energy released when bonds made			1715

$\Delta H = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds made})$
$\Delta H = 1816 - 1715 = 101\,\text{kJ}\,\text{mol}^{-1}$ **[3]**

d We have been given enthalpy change of formation data, so we can use the equation:
$\Delta H_r = \Sigma\Delta H_f(\text{products}) - \Sigma\Delta H_f(\text{reactants})$
The enthalpy change of formation of elements in their standard states is zero. Substituting known values we get:
$-622 = (2 \times -286) - \Delta H_f(N_2H_4(l))$
Rearranging this we get:
$\Delta H_f(N_2H_4(l)) = 50\,\text{kJ}\,\text{mol}^{-1}$ **[3]**

e We have calculated ΔH_f values in parts **c** and **d** and can use these in the equation:
$\Delta H = \Sigma\Delta H_f(\text{products}) - \Sigma\Delta H_f(\text{reactants})$
$\Delta H = 101 - 50 = 51\,\text{kJ}\,\text{mol}^{-1}$ **[2]**

14 a Lattice enthalpy is the enthalpy change when one mole of an ionic compound is broken apart into its constituent gaseous ions under standard conditions. **[2]**
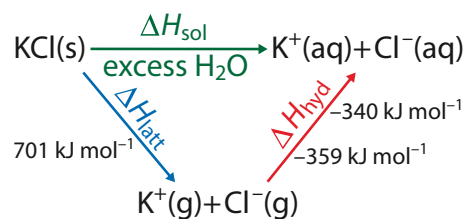
b i $KCl(s) \rightarrow K^+(g) + Cl^-(g)$
ii $Cl(g) + e^- \rightarrow Cl^-(g)$
iii $K(g) \rightarrow K^+(g) + e^-$ **[3]**

c
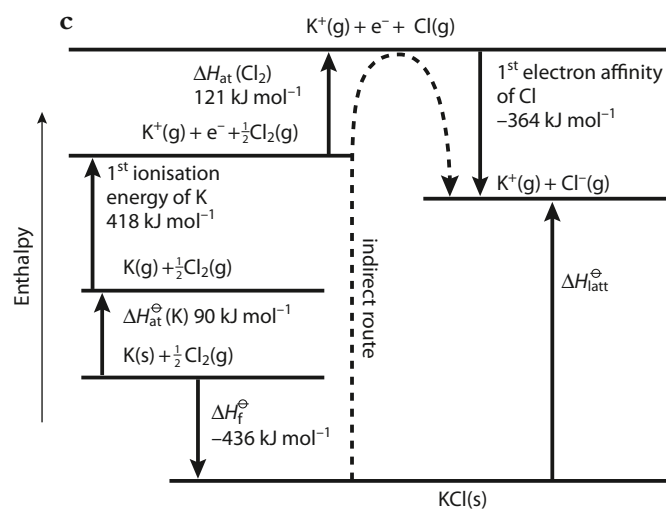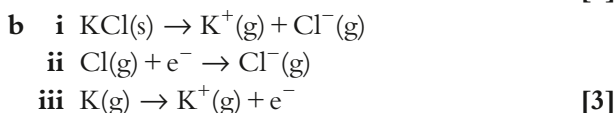
The enthalpy change for the indirect route is the same as that for the direct route:
$\Delta H_{latt} = -(-436) + 90 + 418 + 121 - 364$
$\qquad = 701\,\text{kJ}\,\text{mol}^{-1}$ **[5]**

d Calcium chloride contains 2+ and 1− ions, whereas potassium chloride contains 1+ and 1− ions. The electrostatic attraction between more highly charged ions is greater, so there are stronger electrostatic attractions in the calcium chloride lattice and more energy is required to break the lattice apart. The Ca^{2+} ion is also smaller than the K^+ ion, and this will also lead to the electrostatic attraction between the Ca^{2+} ion and the Cl^- ion being greater. **[2]**

e Using the lattice enthalpy from part **c**, a cycle can be constructed:

$\Delta H_{sol} = 701 - 340 - 359 = 2\,\text{kJ}\,\text{mol}^{-1}$ **[2]**