Chapter 5 – Answers to end of chapter questions

1  a \[ Q = m \times c \times \Delta T = 400 \times 4.2 \times 10 = 16800 \text{ J} = 16.8 \text{ kJ} \]

b Number of moles ethanol used = 0.02.
Energy produced per mole = 840 kJ.

c The experiment is not carried out under standard conditions (i.e. 298 K, 1 atm.)

d Any 3 of the following:

• Heat loss to surroundings
• Heat loss to calorimeter and thermometer
• Incomplete combustion
• Evaporation of ethanol when not burning
• Water not stirred
• Loss of heat from water

2  a \[
\text{TiO}_2(s) + 2\text{C}(s) \rightarrow \text{Ti}(s) + 2\text{CO}(g)
\]

b

\[
\text{TiO}_2(s) + 2\text{C}(s) \xrightarrow{\Delta H_f} \text{Ti}(s) + 2\text{CO}(g)
\]

\[
\begin{align*}
-940 \text{kJ} \\
2 \times -110 \text{kJ}
\end{align*}
\]

\[
\Delta H_f = -(940) + (2 \times -110) \text{kJ mol}^{-1}
= +940 - 220 = +720 \text{kJ mol}^{-1}
\]

3  a Number of moles of butane \[= \frac{1.12}{22.4} = 0.05 \]

Heat liberated by burning butane \[= 0.05 \times 3000 \times \frac{3}{4} = 112.5 \text{ kJ} \]

.: If volume of water which they could boil is \[x \text{ cm}^3\],

\[x \times 4.2 \times 80 = 112500\]

\[\Rightarrow x = 335 \text{ cm}^3\]
b Other assumptions:
• 1 mole of butane at 0°C and 1 atm occupies 22.4 dm³.
• The butane is completely burnt.
• The mass of 1 cm³ of water is 1.0 g.
• The specific heat capacity of water = 4.2 J g⁻¹ K⁻¹.
• The water boils at 100°C.

4 a i (3 × 12) + (8 × 1) = 44
   ii (4 × 12) + (10 × 1) = 58

b i −2220 kJ mol⁻¹ ÷ 44 g mol⁻¹ = 50.5 kJ/g
   ii −2880 kJ mol⁻¹ ÷ 58 g mol⁻¹ = 49.7 kJ/g

c Butane and propane transfer similar amounts of heat on burning 1 g.

d Propane has a lower boiling point, so remains in the gas state at lower temperatures than butane.

5 a The double bond is shorter and stronger. This is because there are two pairs of electrons to pull the oxygen atoms close together in a double bond, and only one pair of electrons to pull the oxygen atoms close together in a single bond.

b The triple bond is shorter and stronger. This is because there are three pairs of electrons to pull the carbon atoms close together in a triple bond, and only one pair of electrons to pull the carbon atoms close together in a single bond.

c (triple bond energy) – (single bond energy) = 838 kJ mol⁻¹ – 350 kJ mol⁻¹
   = 488 kJ mol⁻¹

Assuming that each π bond is of equal strength, then the energy required to break one π bond = 488 kJ mol⁻¹ ÷ 2 = 244 kJ mol⁻¹.

6 a ΔH (step 1) = + 413 + 158 − 565 = + 6 kJ;  ΔH (step 2) = −495 kJ.

b Yes, step 1 with a small endothermic value will be relatively slow compared to step 2 with a large exothermic value.

c CH₄ + Cl₂ → CH₃• + HCl + Cl•  ΔH = + 413 + 242 − 431 = +224 kJ

   CH₃• + Cl• → CH₃Cl  ΔH = −339 kJ

d Yes

e The activation energy for step 1 is much lower for CH₄/F₂ than for CH₄/Cl₂. The energy in sunlight is required to provide part of the activation energy for the CH₄/Cl₂ reaction.

7 a Energy transferred on breaking bonds in reactants
   = +(413 + 242) = +655 kJ mol⁻¹
Energy transferred on making bonds in products
= −[338 + 431] = −769 kJ mol⁻¹
Overall energy change = +654 kJ mol⁻¹ − 769 kJ mol⁻¹ = −114 kJ mol⁻¹

b Enthalpy change of reaction = \( \Delta H_f^0 \) (products) − \( \Delta H_f^0 \) (reactants)
= \((-82.0 + -92.3)\) kJ mol⁻¹ − \((-74.8)\) kJ mol⁻¹
= −99.5 kJ mol⁻¹

c Some of the bond energies given are average values – they are not necessarily the same as the bond energies in the given compounds.

8 a The lattice energy of an ionic compound, \( \Delta H_{\text{latt}} \), is the enthalpy change of formation for one mole of the substance from gaseous ions under standard conditions.
i.e. \( X^+(g) + Y^-(g) \rightarrow X^+Y^- (s) \)
The hydration energy of an ionic compound, \( \Delta H_{\text{hyd}} \), is the enthalpy change when gaseous ions making up one mole of the substance are solvated by water molecules to form an infinitely dilute solution.
\( X^+(g) + Y^-(g) + (aq) \rightarrow X^+(aq) + Y^-(aq) \)
The enthalpy change of solution, \( \Delta H_{\text{soln}} \), of an ionic compound is the enthalpy change when one mole of the substance dissolves in water to form an infinitely dilute solution.
\( X^+Y^- (s) + (aq) \rightarrow X^+(aq) + Y^- (aq) \)
b
\[ X^+(g) + Y^-(g) + (aq) \]
\[ \Delta H_{\text{latt}} \]
\[ \Delta H_{\text{hyd}} \]
\[ X^+(aq) + Y^-(aq) \]
\[ \Delta H_{\text{soln}} \]
c \( \Delta H_{\text{latt}} = \Delta H_{\text{hyd}} − \Delta H_{\text{soln}} \)
\( \Rightarrow -642 = \Delta H_{\text{hyd}} − 21 \)
\( \Delta H_{\text{hyd}} = -621 \) kJ mol⁻¹

9 a i The (first) ionisation energy of an element is the energy required to remove one electron from each atom in a mole of gaseous atoms under standard conditions.

ii The atomisation energy of an element is the enthalpy change when one mole of gaseous atoms is formed from the element under standard conditions.
b See section 5.13 for full labelling.
(All values on the cycle are in kJ mol⁻¹)
c \[ \Delta H_{\text{latt}}(\text{KBr(s)}) = +342 - 420 - 112 - 90 - 392 = -672 \text{ kJ mol}^{-1} \]

d From KF to KI, the lattice energy gets less exothermic. As the ionic radius of the halide ion increases, its charge density decreases. Its attraction to the positive K\(^+\) ion also decreases, so the lattice energy gets less exothermic.

10 a \[ \Delta H_{\text{latt}}(\text{CaO(s)}) = -790 + 141 - 249 - 1100 - 590 - 177 - 636 \]
\[ = -3401 \text{ kJ mol}^{-1} \]

b \[ \Delta H_{i_1}(\text{Ca}) \] involves the removal of one electron from neutral atoms. \[ \Delta H_{i_2}(\text{Ca}) \] involves the removal of a second electron from positive Ca\(^+\) ions. The attraction of negative electrons for Ca\(^+\) ions makes \[ \Delta H_{i_2} \] more endothermic.

c O atoms readily gain electrons. The first electron is being added to neutral O atoms and this process, \[ \Delta H_e(O) \], is exothermic. However, the second electron must be added to a negative O\(^-\) ion. The repulsion between the O\(^-\) ion and an approaching electron means that this process, \[ \Delta H_e(O^-) \], will be endothermic.

d The first ionisation energy of Mg will be more endothermic than that for Ca. The atomic radius of Mg is smaller than that of Ca and the screening effect of inner shells of electrons will be less for Mg than Ca. Both these factors will increase the attraction of the nucleus for the outermost electron. However, the nuclear charge in Mg is less than that in Ca which would cause the attraction of the outermost electron to be less. Overall, the distance and shielding factors outweigh the nuclear charge factor so the outermost electron is held more strongly by Mg than Ca atoms, so the first ionisation energy is more endothermic for Mg.