Chemistry in society

Homework exercise

Equilibrium, Hess Law, Enthalpy
Equilibrium

Exercise 1

1) A reaction can be described as

\[ A + B \rightleftharpoons C + D \]

a) Write an equation to show the forwards reaction

b) Write an equation to show the reverse reaction

2) What happens to:

a) the concentration of reactants and products
b) the rate of the forwards and reverse reactions

when a reaction reaches equilibrium?

3) What does the term 'dynamic equilibrium' mean?

4) Define Le Chatelier's principle

5) What will happen to a reaction at equilibrium when:

a) The pressure is increased
b) The temperature is increased
c) The concentration of a reactant is increased
d) A substance is added which reacts with and removes a reactant

6) Use Le Chatelier's principle to explain each of your answers to question 5.

7) What effect will addition of a catalyst have on:

a) the rate of the forward reaction
b) the rate of the reverse reaction
c) the time taken to reach equilibrium
d) the position of the equilibrium
Equilibrium

Exercise 2

1) Chemical reactions are in a state of dynamic equilibrium only when:
   A  The rate of the forward reaction equals that of the backward reaction.
   B  The concentration of reactants and products are equal.
   C  The activation energies of the forward and backward reactions are equal.
   D  The reaction involves zero enthalpy change.

2) Increasing the pressure on the equilibrium shown below:
   \[ \text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g}) \quad \Delta H = +58\text{kJ} \]
   Will cause the position of the equilibrium to
   A  move to the product side
   B  move to the reactant side
   C  remain unchanged
   D  increase

3) Decreasing the temperature in the reaction shown below
   \[ \text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g}) \quad \Delta H = +58\text{kJ} \]
   Will cause the position of the equilibrium to
   A  move to the product side
   B  move to the reactant side
   C  remain unchanged
   D  increase
Questions 4 and 5 refer to the following equilibrium which exists in bromine water.

\[ \text{Br}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons 2\text{H}^+(\text{aq}) + \text{BrO}^-(\text{aq}) + \text{Br}^-(\text{aq}) \]

4) Addition of which of the following substances would move the equilibrium position to the right?

A. potassium nitrate  
B. sodium bromide  
C. sulphuric acid  
D. sodium hydroxide

5) Addition of which of the following substances would increase the pH of the equilibrium mixture?

A. potassium nitrate  
B. sodium bromide  
C. bromine  
D. sodium chloride

6) Methanol can be prepared from synthesis gas as follows:

\[ \text{CO}(\text{g}) + 2\text{H}_2(\text{g}) \rightleftharpoons \text{CH}_3\text{OH}(\text{g}) \quad \Delta H = -91\text{kJ} \]

The formation of methanol is favoured by:

A. high pressure and low temperature  
B. high pressure and high temperature  
C. low pressure and low temperature  
D. low pressure and high temperature
7) In the Haber process nitrogen and hydrogen are converted into ammonia in an exothermic equilibrium reaction. The operating conditions are 250 atmospheres pressure and a temperature of about 500°C in the process.

a) Write a balanced equation for the reaction.

b) The highest percentage yield of ammonia would be obtained at a low operating temperature. Explain, therefore, the use of a temperature of about 500°C in the process.

c) The reaction is not allowed to reach equilibrium.

i) State how this is done.

ii) Explain why this is done.

8) Iodine dissolves only slightly in water, the process being endothermic. With excess iodine present, the following equilibrium is set up.

\[ \text{I}_2(s) + \text{aq} \rightleftharpoons \text{I}_2(\text{aq}) \quad \Delta H \text{ positive} \]

The concentration of dissolved iodine was measured over a period of time. The graph below was obtained as the iodine dissolved.

a) Copy the graph and add a curve to show how the iodine concentration would change with time if the measurements were repeated at a higher temperature.
b) The dissolved iodine reacts with water as follows:

\[ \text{I}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons 2\text{H}^+(aq) + \text{I}^-(aq) + \text{IO}^-(aq) \]

i) Copy and complete the table to show the effect on the equilibrium of adding each of the solids.

<table>
<thead>
<tr>
<th>Solid</th>
<th>Effect on equilibrium position</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium iodide</td>
<td></td>
</tr>
<tr>
<td>Potassium sulphate</td>
<td></td>
</tr>
</tbody>
</table>

ii) Why does the position of the equilibrium move to the right when solid potassium hydroxide is added?

9) Consider the following equilibrium:

\[ \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \quad \Delta H \text{ positive} \]

(pale yellow) (dark brown)

What would be **seen** if the equilibrium mixture was:

a) placed in a freezing mixture

b) compressed?
Exercise 3

1) \[ 2\text{NO}_2(g) \rightleftharpoons \text{N}_2\text{O}_4(g) \quad \Delta H = +56\text{kJ} \]

Which two conditions favour the decomposition of NO\(_2\)?

A. low temperature, high pressure
B. high temperature, low pressure
C. low temperature, low pressure
D. high temperature, high pressure

2) Which of the following is likely to apply to the use of a catalyst in a chemical reaction?

<table>
<thead>
<tr>
<th>Rate of forward reaction</th>
<th>Rate of reverse reaction</th>
<th>Position of equilibrium</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>Increased</td>
<td>Unchanged</td>
</tr>
<tr>
<td>B</td>
<td>Increased</td>
<td>Unchanged</td>
</tr>
<tr>
<td>C</td>
<td>Increased</td>
<td>Decreased</td>
</tr>
<tr>
<td>D</td>
<td>unchanged</td>
<td>unchanged</td>
</tr>
</tbody>
</table>

3) Consider the following equilibrium:

\[ \text{N}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}(g) \quad \Delta H = +180\text{kJ} \]

How would the equilibrium concentration of nitrogen oxide be affected by:

a) increasing the temperature
b) decreasing the pressure
c) decreasing the concentration of oxygen
4) When chlorine is dissolved in water, the following equilibrium is set up.

\[ \text{Cl}_2(g) + \text{H}_2\text{O}(l) \rightleftharpoons 2\text{H}^+(aq) + \text{ClO}^-(aq) + \text{Cl}^-(aq) \]

The efficiency of a chlorine bleach depends on a high concentration of \( \text{ClO}^- \) ions. Various substances can be added to achieve this, e.g. sodium carbonate and silver nitrate.

a) Explain why each of these compounds should increase the concentration of \( \text{ClO}^- \) ions.

b) Why is silver nitrate unlikely to be used in practice.

c) Explain what effect adding potassium hydroxide solution would have on the position of the equilibrium.

5) Soda water is made by dissolving carbon dioxide in water, under pressure.

\[ \text{CO}_2(g) \rightleftharpoons \text{CO}_2(aq) \]

a) When the stopper is taken off a bottle of soda water, the carbon dioxide gas escapes. Explain why the drink eventually goes completely flat.

b) This graph shows the solubility of carbon dioxide in water at different temperatures.
What does the graph indicate about the enthalpy of solution of carbon dioxide in water?

c) When all of the carbon dioxide is removed from one litre of soda water at 0°C, the gas is found to occupy 1.7 litres. Use the information in the graph to calculate the molar volume of carbon dioxide at this temperature.

6) In the contact process, sulphur trioxide is formed from sulphur dioxide and oxygen in a reversible reaction. The equation for the reaction is:

\[ 2\text{SO}_2(g) + \overset{\rightleftharpoons}{\text{O}_2(g)} \rightarrow 2\text{SO}_3(g) \]

The forward reaction is exothermic.

a) One litre of sulphur dioxide and one litre of oxygen were mixed under certain conditions and allowed to reach equilibrium. Analysis of the equilibrium mixture showed that 60% conversion of the sulphur dioxide had taken place.

Calculate the volume of each gas present in the equilibrium mixture.

b) What effect, if any, will there be on the equilibrium position if:

i) the temperature is increased
ii) a catalyst is used

iii) the pressure is increased?
1) What name is given to the device which measures the energy given out by a fuel?

2) Copy the passage and delete the incorrect phrases.

   The burning of a fuel is an endothermic / exothermic reaction in which energy is given out / taken in. This means the products have more / less energy than the reactants.

3) Using the equation \( \Delta h = cm \Delta T \) calculate the energy given out in the following examples. Show working.

   a) Methane burns and raises the temperature of 200cm\(^3\) of water from 20°C to 30.5°C.
   b) Propanol is burned causing the temperature of 500cm\(^3\) of water to increase by 23°C.
   c) A fuel burns, increasing the temperature of 150cm\(^3\) of water by 12°C.
   d) The temperature of 50cm\(^3\) of water increases from 25°C to 29°C when propane burns.

4) When butanol burns the temperature of 250cm\(^3\) of water increases by 16°C. The energy released is:

   A. 16.7 kJ  
   B. 167 kJ  
   C. 16720 kJ  
   D. 1.67 kJ

5) A can containing 0.1kg of water at 21°C is heated to 29°C when a burner containing an alcohol is lit underneath it. Calculate the energy given out.

6) The burning of methane is used to raise the temperature of 0.25L of water from 18.5°C to 28.5°C. Calculate the energy given out.

7) Propane is burned to heat 250cm\(^3\) of water from 22.5°C to 31°C. Calculate the energy given out.
Calculating Enthalpy of combustion

Exercise 5

1) Why is it important to predict the enthalpy change of an industrial reaction?

2) When 0.58g of butane (C\textsubscript{4}H\textsubscript{10}) was burned, the heat produced raised the temperature of 200cm\textsuperscript{3} of water from 20°C to 54.5°C. Calculate the enthalpy of combustion of butane.

3) 0.02 moles of a hydrocarbon is burned completely in air and the heat produced is used to heat 0.1kg of water from 20.5°C to 29.5°C. Calculate the enthalpy of combustion of the fuel.

4) A can containing 0.1kg of water at 21°C is heated to 29°C when a burner containing alcohol is lit underneath it. If 0.02 moles of the alcohol is burned in the process, calculate its enthalpy of combustion.

5) 0.32g of methanol CH\textsubscript{3}OH is burned in a spirit burner which is used to heat up 0.2kg of water from 19.5°C to 27.5°C. Calculate the enthalpy of combustion of methanol.

6) The burning of 0.2g of methane, CH\textsubscript{4}, is used to raise the temperature of 0.25kg of water from 18.5°C to 28.5°C. Calculate the enthalpy of combustion of methane.

7) A burner containing ethanol, C\textsubscript{2}H\textsubscript{5}OH, is used to heat up 0.4kg of water from 21°C to 37°C. In the process, 0.92g of ethanol is burned. Calculate the enthalpy of combustion of ethanol.

8) 0.22g of propane, C\textsubscript{3}H\textsubscript{8}, is burned to heat 0.25kg of water by 10°C. Calculate the enthalpy of combustion of propane.

9) A gas burner containing butane, C\textsubscript{4}H\textsubscript{10}, is used to heat 0.15kg of water from 22.5°C to 31°C. 0.116g of butane is burned in the process. Calculate the enthalpy of combustion of butane.

Re-arranging \( \Delta h = cm\Delta T \)
1) The enthalpy of combustion of ethanol, \( \text{C}_2\text{H}_5\text{OH} \), is \(-1367\text{kJ mol}^{-1}\). If a spirit burner containing ethanol was used to heat a can of water, what mass of ethanol would raise the temperature of 300cm\(^3\) of water by 10°C?

2) The enthalpy of combustion of propanol is \(-2020\text{kJ mol}^{-1}\). A burner containing propanol, \( \text{C}_3\text{H}_7\text{OH} \), is used to heat up 200cm\(^3\) of water. What mass of propanol would require to be burned to produce a temperature rise of 13.5°C?

3) A bunsen burner uses methane, \( \text{CH}_4 \), which has an enthalpy of combustion of \(-891\text{kJ mol}^{-1}\). If 0.4g of methane were completely burned to heat a can containing 500cm\(^3\) of water, what would be the maximum temperature rise which would be produced?

4) The enthalpy of combustion of methanol was calculated experimentally as \(-669\text{kJ mol}^{-1}\). 0.02moles of methanol are used to heat 300cm\(^3\) of water.
   a) Calculate the temperature rise of the water.
   b) Explain why the experimentally derived enthalpy of combustion is significantly lower than the value quoted in the databook.

5) The enthalpy of combustion of ethanol, \( \text{C}_2\text{H}_5\text{OH} \), is \(-1367\text{kJ mol}^{-1}\). If a spirit burner containing ethanol was used to heat a can of water, what temperature rise would occur when 0.5g of ethanol are used to heat 100cm\(^3\) of water.

6) The enthalpy of combustion of methanol is \(-727\text{kJ mol}^{-1}\). A burner containing methanol, \( \text{CH}_3\text{OH} \), is used to heat up 400cm\(^3\) of water. What temperature rise would be produced in the water if 0.64g of methanol were completely burned?

7) An experiment was carried out using an ethanol burner.
   a) Write the equation for the complete combustion of ethanol, \( \text{C}_2\text{H}_5\text{OH}(l) \) and write down the enthalpy of combustion from your data book.
   b) If the apparatus used is 50% efficient, calculate the mass of ethanol burned when 80cm\(^3\) of water rises in temperature from 20.5°C to 31.0°C.
1) The boiling point of pentane is 36°C. Which equation illustrates the enthalpy of combustion of pentane?

A  \( C_5H_{12}(l) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(l) \)
B  \( C_5H_{12}(l) + 5.5O_2(g) \rightarrow 5CO(g) + 6H_2O(l) \)
C  \( C_5H_{12}(g) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g) \)
D  \( C_5H_{12}(l) + 5O_2(g) \rightarrow 5CO_2(g) + 6H_2(g) \)

2) When 3.6g of glucose, \( C_6H_{12}O_6 \), was burned, 56kJ of energy was released. From this data, what is the enthalpy of combustion of glucose in kJ mol\(^{-1}\)?

A  -15.6
B  +15.6
C  -2800
D  +2800

3) The enthalpy of combustion of benzene (\( C_6H_6 \)) is -3268kJ mol\(^{-1}\). When 1.8g of benzene is completely burned the enthalpy change is:

A  -32.7kJ
B  -75.4kJ
C  -32.7kJ
D  -65.4kJ

4) \( 5N_2O_4(l) + 4CH_3NHNH_2(l) \rightarrow 4CO_2(g) + 12H_2O(l) + 9N_2(g) \quad \Delta H = -5116kJ \)
   The energy released when 2 moles of each reactant are mixed and ignited is:

A  2046kJ
B  2558kJ
C  4093kJ
D  5116kJ
5) Aluminium reacts with oxygen to form aluminium oxide.

\[ 2\text{Al}(s) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{Al}_2\text{O}_3(s) \quad \Delta H = -1670\text{kJ mol}^{-1} \]

What is the enthalpy of combustion of aluminium in kJ mol\(^{-1}\)?

A. -835  
B. -1113  
C. -1670  
D. +1670

6) The enthalpies of combustion of some alcohols are shown in the table below:

<table>
<thead>
<tr>
<th>Name of alcohol</th>
<th>Enthalpy of combustion /kJ mol(^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methanol</td>
<td>-727</td>
</tr>
<tr>
<td>Ethanol</td>
<td>-1367</td>
</tr>
<tr>
<td>Propan-1-ol</td>
<td>-2020</td>
</tr>
</tbody>
</table>

a) Using this data predict the enthalpy of combustion of butan-1-ol in kJ mol\(^{-1}\).

b) A value for the combustion of butan-2-ol, \(\text{C}_4\text{H}_9\text{OH}\), can be determined experimentally using the apparatus shown.

Mass of butan-2-ol burned = 1.0g  
Temperature rise of water = 40°C

Use these results to calculate the enthalpy of combustion of butan-2-ol in kJ mol\(^{-1}\).
7) A calorimeter like the one shown below can be used to measure the enthalpy of combustion of ethanol. The ethanol is ignited and burns completely in the oxygen gas. The heat energy released in the reaction is taken in by the water

- **a)** Explain why the sample is stirred.
- **b)** The value for the enthalpy of combustion of ethanol obtained by the calorimeter method is higher than the value obtained by the typical school method.
  
  *One reason for this is that more heat is lost to the surroundings in the school method. Give one other reason for the value being higher with the calorimeter method.*

- **c)** In one experiment the burning of 0.980g of ethanol resulted in the temperature of 400cm³ of water rising from 14.2°C to 31.6°C.
  
  Use this information to calculate the enthalpy of combustion of ethanol.
  
  *Show your working clearly.*
**Hess Law – given equations**

Exercise 8

1) Define Hess Law

2) The enthalpy of formation of ethyne:

\[ 2C(s) + H_2(g) \rightarrow C_2H_2(g) \]

Can be calculated using the enthalpies of combustion of carbon, hydrogen and ethyne.

Given that:

\[ 2C(s) + 2O_2(g) \rightarrow 2CO_2(g) \quad \Delta H = -788kJ \]
\[ H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -286kJ \]
\[ 2CO_2(g) + H_2O(l) \rightarrow C_2H_2(g) + 2\frac{1}{2}O_2(g) \quad \Delta H = +1300kJ \]

Calculate the enthalpy of formation of ethyne.

3) Methanol can be formed from hydrogen, carbon and oxygen as follows:

\[ C(s) + 2H_2(g) + \frac{1}{2}O_2(g) \rightarrow CH_3OH(l) \]

Use the enthalpies of combustion of carbon, hydrogen and methanol to calculate the enthalpy change for this reaction.

\[ C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -394kJ \]
\[ 2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \quad \Delta H = -573kJ \]
\[ CH_3OH(l) + 1.5 O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -727kJ \]

4) The simplest silicon hydride has the formula SiH₄. It is a gas at room temperature. The equation for the complete combustion of this compound is:

\[ SiH_4(g) + 2O_2(g) \rightarrow SiO_2(s) + 2H_2O(l) \]

Calculate the enthalpy change for this reaction using the following data:

\[ Si(s) + 2H_2(g) \rightarrow SiH_4(g) \quad \Delta H = +34kJ \]
\[ Si(s) + O_2(g) \rightarrow SiO_2(s) \quad \Delta H = -911kJ \]
\[ H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -286kJ \]
5) The equation for the combustion of methane, CH₄, is given below:

\[ CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \]

Calculate the enthalpy of combustion of methane using the data below:

\[ C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -394kJ \]
\[ H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -286kJ \]
\[ C(s) + 2H_2(g) \rightarrow CH_4(g) \quad \Delta H = -75kJ \]

6) The enthalpy of formation of ethane, C₂H₆, is represented by the following equation:

\[ 2C(s) + 3H_2(g) \rightarrow C_2H_6(g) \]

Calculate the enthalpy of combustion of ethane using the data below:

\[ C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -394kJ \]
\[ H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -286kJ \]
\[ C_2H_6(g) + 3\frac{1}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H = -1560kJ \]

7) The equation for the combustion of ethanol, C₂H₅OH, is given below:

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

Calculate the enthalpy of combustion of methane using the data below:

\[ C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -394kJ \]
\[ H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -286kJ \]
\[ 2C(s) + 3H_2(g) + \frac{1}{2}O_2(g) \rightarrow C_2H_5OH(l) \quad \Delta H = -278kJ \]
1) Methane can be converted into ethyne at very high temperatures (about 1500°C). The equation for the reaction is:

\[2\text{CH}_4(\text{g}) \rightarrow \text{C}_2\text{H}_2(\text{g}) + 3\text{H}_2(\text{g})\]

Use the enthalpies of combustion from the SQA data book to calculate the enthalpy change for this reaction.

2) Use appropriate enthalpies of combustion from the SQA data book to calculate the enthalpy changes for the following reactions.

   a) \[3\text{C}(\text{s}) + 4\text{H}_2(\text{g}) \rightarrow \text{C}_3\text{H}_8(\text{g})\] (propane)

   b) \[2\text{C}(\text{s}) + 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CH}_3\text{COOH}(\text{l})\] (ethanoic acid)

3) The formation of methanoic acid, HCOOH, is represented by the following equation:

\[\text{C}(\text{s}) + \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{HCOOH}(\text{l})\]

Calculate \(\Delta H\) for the above reaction using the enthalpies of combustion of carbon, hydrogen and methanoic acid.

4) The fermentation of glucose, \(\text{C}_6\text{H}_{12}\text{O}_6\), to ethanol and carbon dioxide can be represented by the equation below:

\[\text{C}_6\text{H}_{12}\text{O}_6(\text{s}) \rightarrow 2\text{C}_2\text{H}_5\text{OH}(\text{l}) + 2\text{CO}_2(\text{g})\]

Calculate the \(\Delta H\) for this reaction using the enthalpies of combustion of ethanol and glucose. (The enthalpy of combustion of glucose is \(-2813\text{kJ mol}^{-1}\)).

5) Use the enthalpies of combustion of ethyne, \(\text{C}_2\text{H}_2\), ethane, \(\text{C}_2\text{H}_6\) and hydrogen to calculate the \(\Delta H\) for the complete hydrogenation of ethyne given by the equation below:

\[\text{C}_2\text{H}_2(\text{g}) + 2\text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})\]
**Hess Law past paper questions**

**Exercise 10**

1) The formation of hydrogen peroxide can be shown as:

\[
\text{H}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow \text{H}_2\text{O}_2\text{(l)}
\]

If \[\text{H}_2\text{O}_2\text{(l)} \rightarrow \text{H}_2\text{O}\text{(l)} + \frac{1}{2}\text{O}_2\text{(g)}\] \[\Delta H = -52\text{kJ mol}^{-1}\]

\[\text{H}_2\text{(g)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{H}_2\text{O}\text{(l)}\] \[\Delta H = -286\text{kJ mol}^{-1}\]

What is the enthalpy of formation of hydrogen peroxide (H\(_2\text{O}_2\))?

A. \(-234\text{kJ mol}^{-1}\)
B. \(+234\text{kJ mol}^{-1}\)
C. \(-338\text{kJ mol}^{-1}\)
D. \(+338\text{kJ mol}^{-1}\)

2) The enthalpies of combustion of C(s), H\(_2\text{(g)}\) and \(\text{C}_4\text{H}_9\text{OH(l)}\) (in kJ mol\(^{-1}\)) are as follows:

\[\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}\] \[\Delta H = a\]

\[\text{H}_2\text{(g)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{H}_2\text{O}\text{(l)}\] \[\Delta H = b\]

\[\text{C}_4\text{H}_9\text{OH(l)} + 6\text{O}_2\text{(g)} \rightarrow 4\text{CO}_2\text{(g)} + 5\text{H}_2\text{O(l)}\] \[\Delta H = c\]

The enthalpy of formation of butanol is:

\[4\text{C(s)} + 5\text{H}_2\text{(g)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{C}_4\text{H}_9\text{OH(l)}\]

Which of these equations could be used to calculate the enthalpy of formation of butanol?

A. \(4a + 5b - c\)
B. \(2a + 10b - c\)
C. \(c - 4a - 5b\)
D. \(2a + 5b + c\)
3) Consider the reaction:

\[ C \rightarrow \frac{1}{2}A + 1\frac{1}{2}B \quad \Delta H = +48kJ \]

What is the \( \Delta H \) for the reaction \( A + 3B \rightarrow 2C \)?

A. -96kJ
B. -48kJ
C. +96kJ
D. +48kJ

4) Given these equations:

\[
\begin{align*}
Mg(s) + 2H^+(aq) &\rightarrow Mg^{2+}(aq) + H_2(g) \quad \Delta H = a \\
Zn(s) + 2H^+(aq) &\rightarrow Zn^{2+}(aq) + H_2(g) \quad \Delta H = b \\
Mg(s) + Zn^{2+}(aq) &\rightarrow Mg^{2+}(aq) + Zn(s) \quad \Delta H = c
\end{align*}
\]

Then according to Hess’s law:

A. \( c = a + b \)
B. \( c = a - b \)
C. \( c = b - a \)
D. \( c = -b - a \)

5) Diborane is used as rocket fuel.

The equation for the combustion of diborane is shown below:

\[ B_2H_6(g) + 3O_2(g) \rightarrow B_2O_3(s) + 3H_2O(l) \]

a) Calculate the enthalpy of combustion of the gas diborane \( B_2H_6 \) using the following data:

\[
\begin{align*}
2B(s) + 3H_2(g) &\rightarrow B_2H_6(g) \quad \Delta H = +36kJ \text{ mol}^{-1} \\
H_2(g) + \frac{1}{2}O_2(g) &\rightarrow H_2O(l) \quad \Delta H = -286kJ \text{ mol}^{-1} \\
2B(s) + 1\frac{1}{2}O_2(g) &\rightarrow B_2O_3(s) \quad \Delta H = -1274kJ \text{ mol}^{-1}
\end{align*}
\]
b) Diborane can be used to manufacture pentaborane (B₅H₉).

Pentaborane was also considered for use as a rocket fuel because its enthalpy of combustion is -9037kJ mol⁻¹.

Calculate the energy released, in kJ, when 1 kilogram of pentaborane is completely burned.

6) The equation for the enthalpy of formation of propan-1-ol is:

\[ 3C(s) + 4H₂(g) + \frac{1}{2}O₂(g) \rightarrow C₃H₇OH(l) \]

Use information on enthalpies of combustion from the data book to calculate the enthalpy of formation of propan-1-ol.

Show your working clearly.

7) Mobile phones are being developed that can be powered by methanol.

Methanol can be made in a two stage process.

In the first stage, methane is reacted with steam to produce a mixture of hydrogen and carbon monoxide.

\[ CH₄(g) + H₂O(g) \rightleftharpoons CO(g) + 3H₂(g) \]

a) Give the name for the mixture of carbon monoxide and hydrogen which is produced.

b) Use the data below to calculate the enthalpy change, in kJ mol⁻¹, for the forward reaction.

\[ CO(g) + \frac{1}{2}O₂(g) = CO₂(g) \quad \Delta H = -283kJ \text{ mol}^{-1} \]

\[ H₂(g) + \frac{1}{2}O₂(g) = H₂O(g) \quad \Delta H = -242kJ \text{ mol}^{-1} \]

\[ CH₄(g) + 2O₂(g) = CO₂(g) + 2H₂O(g) \quad \Delta H = -803kJ \text{ mol}^{-1} \]
**Bond enthalpy**

**Exercise 11**

1) Use the bond enthalpies quoted in the databook to answer the following questions:
   a) Which halogen has the strongest bond?
   b) Which hydrogen-halogen bond is the easiest to break?
   c) Explain why nitrogen has a higher bond enthalpy than oxygen

2) Calculate the energy required (in kJ) to break 4g of hydrogen gas into atoms.

3) Explain why a C-H bond enthalpy is called a mean bond enthalpy whilst a H-H bond enthalpy is called a molar bond enthalpy.

4) Use bond enthalpies to calculate the enthalpy changes for the following reactions:
   a) \( \text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2\text{HCl}(g) \)
   b) \( 2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g) \)
   c) \( \text{H}_2(g) + \text{F}_2(g) \rightarrow 2\text{HF}(g) \)
   d) \( \text{C}(g) + 2\text{H}_2(g) \rightarrow \text{CH}_4(g) \)
   e) \( \text{HCl}(g) + \frac{1}{2} \text{F}_2(g) \rightarrow \text{HF}(g) + \frac{1}{2} \text{Cl}_2(g) \)

5) Use the bond enthalpies quoted below to calculate the enthalpy change for this reaction:
   \( \text{CO}(g) + \text{H}_2\text{O}(g) \rightarrow \text{CO}_2(g) + \text{H}_2(g) \)

<table>
<thead>
<tr>
<th>Bond</th>
<th>Enthalpy (kJ mol(^{-1}))</th>
</tr>
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<tbody>
<tr>
<td>C-O in carbon monoxide</td>
<td>+358</td>
</tr>
<tr>
<td>C=O in carbon dioxide</td>
<td>+798</td>
</tr>
<tr>
<td>O-H</td>
<td>+458</td>
</tr>
<tr>
<td>H-H</td>
<td>+432</td>
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