

**SUGGESTION ANSWER SCHEME  
 COMPILATION PSPM 2<sup>ND</sup> SEMESTER  
 TOPIC: THERMOCHEMISTRY**

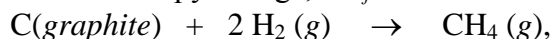
**PSPM 2003/2004 SK027 (1b, 1c & 5b)**

1 (a) State *Hess Law*.

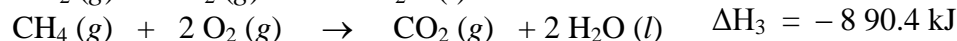
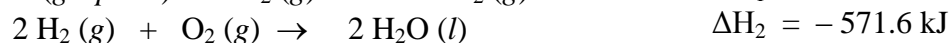
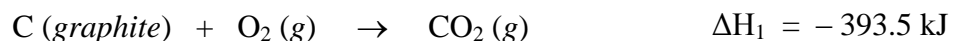
**The change in enthalpy is the same whether the reaction takes place in one step or a series of steps.**

[2 marks]

(b) Determine the enthalpy change,  $\Delta H_f$  for the reaction.



Given:



[3 marks]

$$\begin{aligned} \Delta H_f &= \Sigma \Delta H \text{ product} - \Sigma \Delta H \text{ reactant} \\ &= \Delta H_3 - (\Delta H_1 + \Delta H_2) \\ &= -890.4 - [-393.5 + (-571.6)] \\ &= -74.77 \text{ kJ} \quad @ \quad \text{using algebraic method} \end{aligned}$$

2 Draw the Born-Haber cycle for the formation of magnesium chloride,  $\text{MgCl}_2$  from magnesium metal and chlorine gas. Calculate the enthalpy of formation of  $\text{MgCl}_2$ .

Given:

Heat of sublimation of magnesium,  $\Delta H_1 = +149 \text{ kJ mol}^{-1}$

First ionization energy of magnesium,  $\Delta H_2 = +740 \text{ kJ mol}^{-1}$

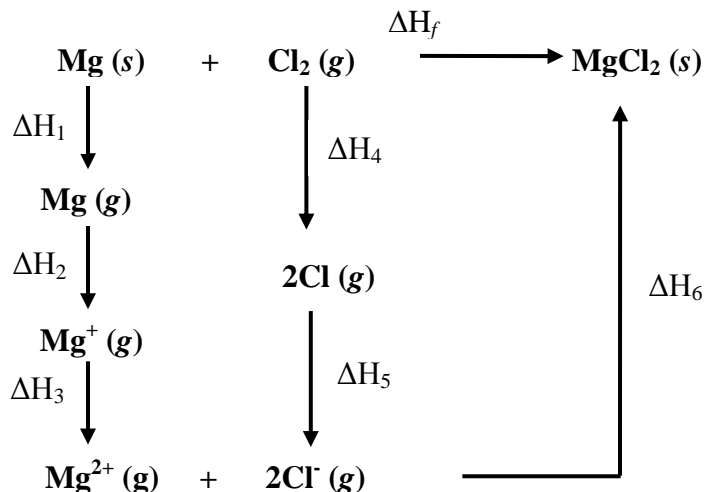
Second ionization energy of magnesium,  $\Delta H_3 = +1456 \text{ kJ mol}^{-1}$

Heat of atomization of chlorine,  $\Delta H_4 = +240 \text{ kJ mol}^{-1}$

Electron affinity of chlorine,  $\Delta H_5 = -369 \text{ kJ mol}^{-1}$

Lattice energy of  $\text{MgCl}_2$ ,  $\Delta H_6 = -3933 \text{ kJ mol}^{-1}$

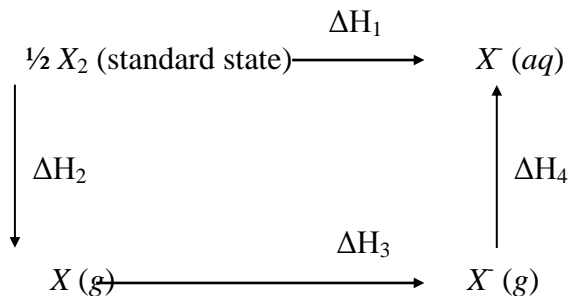
[8 marks]



$$\begin{aligned}\Delta H_f &= \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4 + \Delta H_5 + \Delta H_6 \\ &= 149 + 740 + 1456 + 2(240) + 2(-369) + (-3933) \\ &= -1846 \text{ kJ mol}^{-1}\end{aligned}$$

**PSPM 2004/2005 SK027 (1a, 1b, 6a & 6b)**

- 3 **FIGURE 1** shows an energy cycle for the dissolution of halogen,  $X_2$ .  $X$  represents bromine or iodine.



**FIGURE 1**

- (a) i. What is the phase for bromine and iodine at their standard states?  
Phase for bromine – **liquid**  
Phase for iodine – **solid**
- ii. Name the enthalpy change for  $\Delta H_2$ ,  $\Delta H_3$  and  $\Delta H_4$ .  
 $\Delta H_2 =$  **enthalpy of atomization**  
 $\Delta H_3 =$  **electron affinity**  
 $\Delta H_4 =$  **enthalpy of hydration**

[5 marks]

- (b) **TABLE 1** shows values of  $\Delta H_2$ ,  $\Delta H_3$  and  $\Delta H_4$  of the above data processes for bromine and iodine.

	Bromine (kJ mol <sup>-1</sup> )	Iodine (kJ mol <sup>-1</sup> )
$\Delta H_2$	+ 112	+ 107
$\Delta H_3$	- 342	- 295
$\Delta H_4$	- 351	- 307

- i. Calculate  $\Delta H_1$  values for bromine and iodine.  
**Hess Law,  $\Delta H_1 = \Delta H_2 + \Delta H_3 + \Delta H_4$**

$$\begin{aligned}\text{For bromine, } \Delta H_1 &= 112 + (-342) + (-351) \\ &= -581 \text{ kJ mol}^{-1}\end{aligned}$$

$$\begin{aligned}\text{For iodine, } \Delta H_1 &= 107 + (-295) + (-307) \\ &= -495 \text{ kJ mol}^{-1}\end{aligned}$$

ii. What can be deduced from b(i)?

$\Delta H_1$  Bromine is more exothermic than  $\Delta H_1$  Iodine.

This indicates that  $\text{Br}^-$  (aq) is formed more easily than  $\text{I}^-$  (aq) or bromine dissolves more easily than iodine.

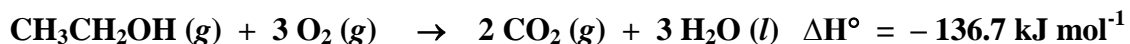
[5 marks]

- 4 (a) Define standard enthalpy of formation,  $\Delta H_f^\circ$ .  
**The heat change when one mole of a compound is formed from its elements in their stable states under standard condition (25 °C and 1 atm).**

The combustion of 4.6 g of ethanol has evolved 136.7 kJ heat at 25 °C and 1 atm.

Calculate the standard enthalpy of formation of ethanol.

$$[\Delta H_f^\circ \text{CO}_2(g) = -393.5 \text{ kJ mol}^{-1} \text{ and } \Delta H_f^\circ \text{H}_2\text{O}(l) = -285.8 \text{ kJ mol}^{-1}]$$



Mr ethanol = 46 g/mol

$$\begin{aligned}n_{\text{ethanol}} &= 4.6 \text{ g} / 46 \text{ g mol}^{-1} \\ &= 0.1 \text{ mol}\end{aligned}$$

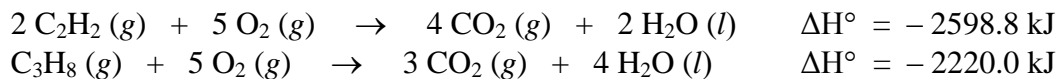
0.1 mol  $\text{CH}_3\text{CH}_2\text{OH}$  released 136.7 kJ

$$\begin{aligned}1.0 \text{ mol } \text{CH}_3\text{CH}_2\text{OH} \text{ released} &= -136.7 \text{ kJ} / 0.1 \\ &= -1367 \text{ kJ mol}^{-1}\end{aligned}$$

$$\begin{aligned}\Delta H^\circ_{\text{rxn}} &= \Sigma \Delta H \text{ product} - \Sigma \Delta H \text{ reactant} \\ -1367 &= [2(-393.5) + 3(-285.8)] - \Delta H_f^\circ \text{CH}_3\text{CH}_2\text{OH} \\ \Delta H_f^\circ \text{CH}_3\text{CH}_2\text{OH} &= -277.4 \text{ kJ mol}^{-1}\end{aligned}$$

[9 marks]

- (b) The following equations show the combustion reactions of ethyne and propane.



A mixture of 5.0 mol gases consisting of ethyne and propane is burnt and 7647.8 kJ of heat is evolved. What is the mole ratio of ethyne to propane in the mixture?

Let:

$$\begin{aligned}x &= \text{moles of } \text{C}_2\text{H}_2 \\ y &= \text{moles of } \text{C}_3\text{H}_8\end{aligned}$$

Total mole:

$$\begin{aligned}x + y &= 5 \\x &= 5 - y \quad \text{----- (1)}\end{aligned}$$

$$\begin{aligned}\text{Enthalpy of combustion 1 mol C}_2\text{H}_2 &= -2598/2 \\&= 1299.4 \text{ kJ}\end{aligned}$$

$$\text{Enthalpy of combustion 1 mol C}_3\text{H}_8 = -2220.0 \text{ kJ}$$

Total heat released:

$$1299.4x + 2220.0y = 7647.8 \quad \text{----- (2)}$$

Solve the simultaneous equation:

$$\begin{aligned}x &= 3.75 \\y &= 1.25\end{aligned}$$

$$\begin{aligned}\therefore \text{Mole ratio n C}_2\text{H}_2 : \text{n C}_3\text{H}_8 &= 3.75 : 1.25 \\&= 3 : 1\end{aligned}$$

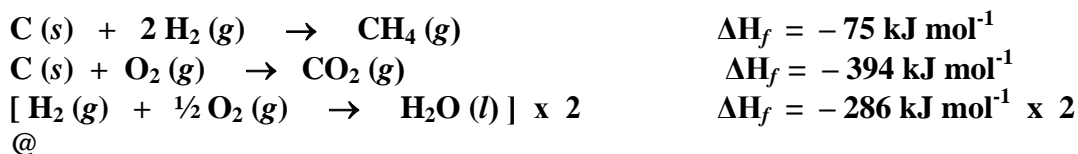
[6 marks]

**PSPM 2005/2006 SK027 (1a & 1b)**

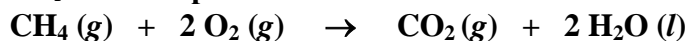
- 5 (a) i. Define standard enthalpy of formation.  
**Energy/heat/enthalpy change when 1 mole of substance is formed from its elements in their stable state under standard condition (1 atm and 25°C).**
- ii. Calculate the heat of combustion of methane using standard heats of formation below:

$$\begin{aligned}\Delta H_f \text{ CO}_2 (g) &= -394 \text{ kJ mol}^{-1} \\ \Delta H_f \text{ H}_2\text{O} (l) &= -286 \text{ kJ mol}^{-1} \\ \Delta H_f \text{ CH}_4 (g) &= -75 \text{ kJ mol}^{-1}\end{aligned}$$

[5 marks]

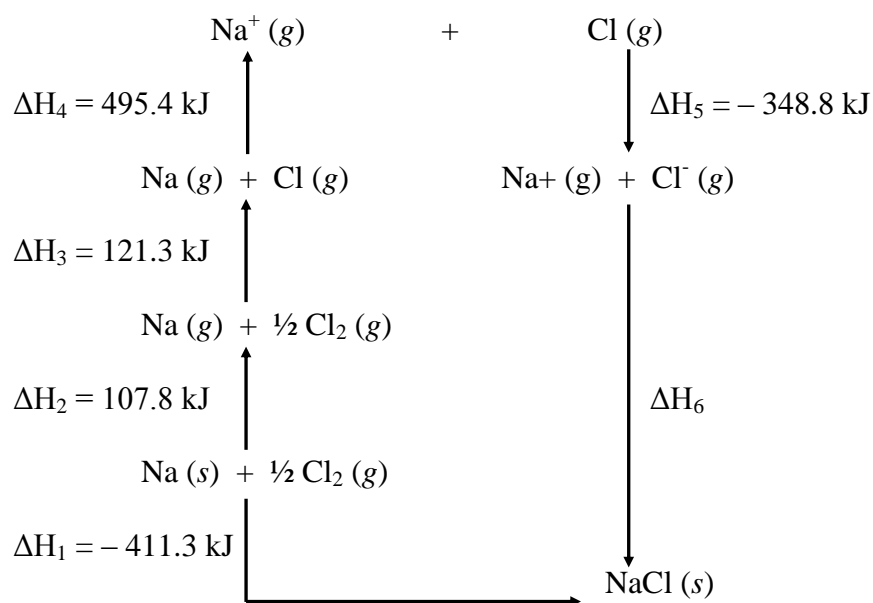


$$\Delta H_f = \Sigma \Delta H \text{ product} - \Sigma \Delta H \text{ reactant}$$



$$\begin{aligned}\Delta H_c \text{ CH}_4 (g) &= -394 + (-572) + 75 \\ &= -891 \text{ kJ mol}^{-1}\end{aligned}$$

- (b) i. Name the enthalpies  $\Delta H_2$ ,  $\Delta H_4$  and  $\Delta H_5$  in the following Born-Haber cycle for sodium chloride, NaCl.  
 $\Delta H_2$  = **atomization energy of sodium**  
 $\Delta H_4$  = **ionization energy of sodium**  
 $\Delta H_5$  = **electron affinity of chlorine**
- ii. Calculate the lattice energy,  $\Delta H_6$  for NaCl.



$$\Delta H_1 = \Delta H_2 + \Delta H_3 + \Delta H_4 + \Delta H_5 + \Delta H_6$$

$$\Delta H_6 = \Delta H_1 - (\Delta H_2 + \Delta H_3 + \Delta H_4 + \Delta H_5) \quad @$$

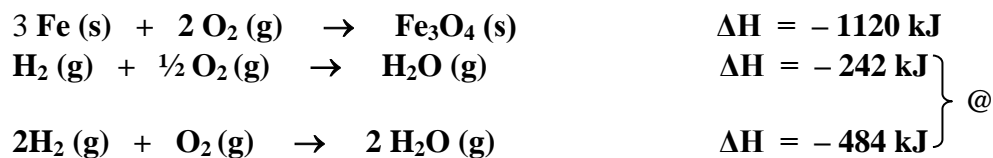
$$= (-411.3 - 107.8 - 121.3 - 495.4 - 348.8) \text{ kJ}$$

$$= -787 \text{ kJ or kJ mol}^{-1}$$

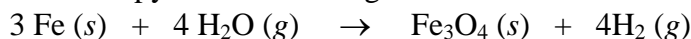
[5 marks]

**PSPM 2006/2007 SK027 (1a & 1b)**

- 6 (a) The enthalpies of formation of  $\text{Fe}_3\text{O}_4 (s)$  and  $\text{H}_2\text{O} (g)$  are  $-1120 \text{ kJ mol}^{-1}$  and  $-242 \text{ kJ mol}^{-1}$  respectively.
- i. Write the thermochemical equations for the formation of both  $\text{Fe}_3\text{O}_4 (s)$  and  $\text{H}_2\text{O} (g)$



ii. Calculate the enthalpy of the following reaction:



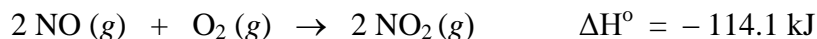
$$\begin{aligned}\Delta H_f &= \Sigma \Delta H \text{ product} - \Sigma \Delta H \text{ reactant} \\ &= [-1120 + 4(0)] - [3(0) + 4(-242)] \\ &= -152 \text{ kJ}\end{aligned}$$

iii. Which of the reaction involved in (i) and (ii) is most easily be made reversible? Give your reason.

**Reaction :  $3 \text{Fe} (s) + 4 \text{H}_2\text{O} (g) \rightarrow \text{Fe}_3\text{O}_4 (s) + 4\text{H}_2 (g)$  @  
reaction in (ii) because less exothermic**

[6 marks]

(b) Calculate the heat released when 0.75 g of nitrogen dioxide,  $\text{NO}_2$  is produced in the thermochemical reaction below.



**From equation:**

$$\begin{aligned}2 \text{ mol of NO}_2 &\approx -114.1 \text{ kJ} \\ 1 \text{ mol of NO}_2 &= -57.05 \text{ kJ}\end{aligned}$$

$$\text{Mr NO}_2 = 46 \text{ g/mol}$$

$$\begin{aligned}n \text{ NO}_2 &= 0.75 \text{ g} / 46 \text{ g mol}^{-1} \\ &= 0.016 \text{ mol}\end{aligned}$$

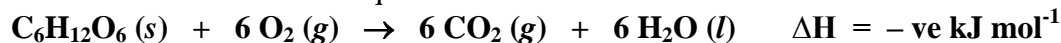
$$\begin{aligned}Q_{rxn} &= 0.016 \times -57.05 \text{ kJ} \\ &= -0.9128 \text{ kJ}\end{aligned}$$

[4 marks]

### PSPM 2007/2008 SK027 (1a & 1b)

7 (a) The enthalpy of combustion of fructose,  $\text{C}_6\text{H}_{12}\text{O}_6$ , is  $21.2 \text{ kJ mol}^{-1}$ . An amount of 2.63 g of  $\text{C}_6\text{H}_{12}\text{O}_6$  was completely combusted in a bomb calorimeter at  $25.0^\circ\text{C}$ .

i. Write the thermochemical equation for the reaction.



ii. Calculate the final temperature if the calorimeter contains 225.0 mL of water.

$$\begin{aligned}n_{\text{fructose}} &= \frac{2.63 \text{ g}}{180.18 \text{ g mol}^{-1}} \\ &= 0.0146 \text{ mol}\end{aligned}$$

**From equation:**

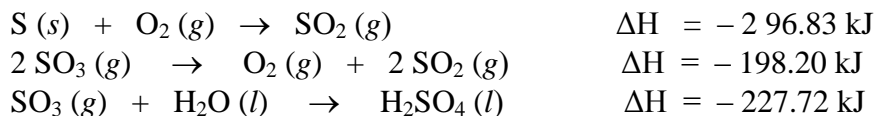
$$\begin{aligned}1 \text{ mol C}_6\text{H}_{12}\text{O}_6 &\cong -21.2 \text{ kJ} \\ 0.0146 \text{ mol C}_6\text{H}_{12}\text{O}_6 &\cong -0.3096 \text{ kJ}\end{aligned}$$

$$\begin{aligned}
 Q_{rxn} &= mc\Delta T \\
 \Delta T &= \frac{0.3096 \text{ mol} \times 1000}{225 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}} \\
 &= 0.33 \text{ }^\circ\text{C}
 \end{aligned}$$

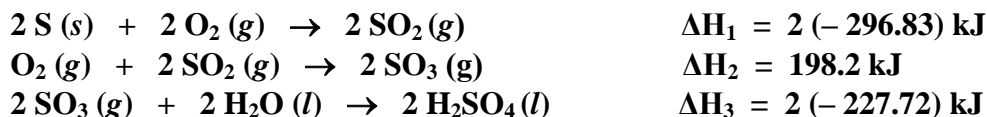
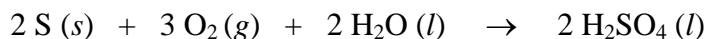
$$\begin{aligned}
 T_{final} &= 25.0 + 0.33 \\
 &= 25.33 \text{ }^\circ\text{C}
 \end{aligned}$$

[5 marks]

(b) The industrial process for the production of sulphuric acid has three steps.



Using the above data, calculate the enthalpy of the reaction,



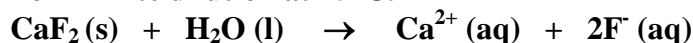
$$\begin{aligned}
 \Delta H &= \Delta H_1 + \Delta H_2 + \Delta H_3 \\
 &= -850.9 \text{ kJ}
 \end{aligned}$$

[5 marks]

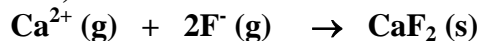
### PSPM 2008/2009 SK027 (5)

- 8 Calcium fluoride,  $\text{CaF}_2$ , is a stable compound in the solid state. The enthalpy of solution is affected by the lattice energy and the hydration energy of the ionic solid. Using  $\text{CaF}_2$  as an example, define standard enthalpy of solution and lattice energy.

**Standard enthalpy of solution is the heat change when 1 mole of  $\text{CaF}_2$  dissolves in water to form a solution of infinite dilution at  $25^\circ\text{C}$ .**

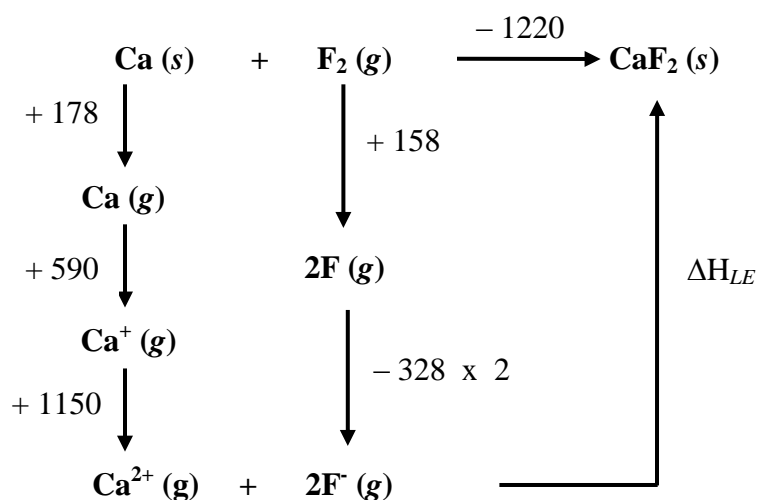


**Lattice energy is the heat released/evolved when 1 mole of an ionic solid  $\text{CaF}_2$  is formed from its gaseous ions,  $\text{Ca}^{2+}$  and  $\text{F}^-$ .**



Construct a Born-Haber cycle for  $\text{CaF}_2$  in the solid state by using the data below. Calculate the lattice energy of  $\text{CaF}_2$ .

	$\Delta H^\circ$ ( $\text{kJ mol}^{-1}$ )
First ionization energy of Ca	+ 590
Second ionization energy of Ca	+ 1150
Standard enthalpy of atomization of Ca	+ 178
Bond energy of F-F bond	+ 158
Electron affinity of F	- 328
Standard enthalpy of formation of $\text{CaF}_2$	- 1220



**Hess's Law:**

$$178 + 590 + 1150 + (-328 \times 2) + x = -1220 \\
 x = -2640 \text{ kJ mol}^{-1}$$

What would you expect the lattice energy of calcium chloride,  $\text{CaCl}_2$ , compared to that of  $\text{CaF}_2$ ? Explain your answer.

In terms of solubility in water,  $\text{CaF}_2$  is less soluble than  $\text{CaCl}_2$ . Explain this observation.

**The lattice energy of  $\text{CaCl}_2$  is expected to be lower (less exothermic or more endothermic) than that of  $\text{CaF}_2$ . This is because the ionic radius of  $\text{Cl}^-$  ion is larger than  $\text{F}^-$  ion. Hence the ionic bond of  $\text{CaCl}_2$  is weaker and the solubility of  $\text{CaF}_2$  is less than that of  $\text{CaCl}_2$ .**

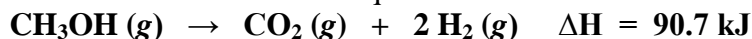
[15 marks]



**PSPM 2009/2010 SK027 (1)**

9 At a constant pressure, the decomposition of 1.0 mol of methanol, CH<sub>3</sub>OH, in the aqueous phase requires 90.7 kJ of heat to produce carbon monoxide and hydrogen gases.

(a) Write the thermochemical equation for the reaction.



(b) Calculate the amount of heat required when 1.60 kg of gaseous methanol decomposes.

$$\begin{aligned} \text{no of mol of CH}_3\text{OH} &= 1.6 \times 10^3 \text{ g} / 32 \text{ g mol}^{-1} \\ &= 50 \text{ mol} \end{aligned}$$

$$1.0 \text{ mol of CH}_3\text{OH required } 90.7 \text{ kJ}$$

$$50.0 \text{ mol of CH}_3\text{OH required } 90.7 \times 50 = 4535 \text{ kJ}$$

(c) The enthalpy change for a given sample of methanol for the above reaction is 64.7 kJ. Calculate the mass of hydrogen gas produced.

$$90.7 \text{ kJ of heat produced from } 1.0 \text{ mol of CH}_3\text{OH}$$

$$64.7 \text{ kJ of heat required } (64.7 \text{ kJ} / 90.7 \text{ kJ}) \times 1.0 \text{ mol} = 0.7133 \text{ mol CH}_3\text{OH}$$

$$1.0 \text{ mol CH}_3\text{OH} \approx 2.0 \text{ mol H}_2$$

$$\begin{aligned} \text{no of mol of H}_2 &= 2.0 \times 0.7133 \text{ mol} \\ &= 1.4266 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{mass of H}_2 &= 1.466 \text{ mol} \times 2 \text{ g mol}^{-1} \\ &= 2.85 \text{ g} \end{aligned}$$

(d) Calculate the amount of heat released when 32.0 g of carbon monoxide reacts completely with hydrogen to form methanol under constant pressure.



$$\begin{aligned} \text{no of mol of CO} &= 32.0 \text{ g} / 28 \text{ g mol}^{-1} \\ &= 1.143 \text{ mol} \end{aligned}$$

$$\begin{aligned} \therefore \text{Amount of heat released} &= (1.143 \text{ mol} / 1.0 \text{ mol}) \times -90.7 \text{ kJ} \\ &= -103.670 \text{ kJ} \end{aligned}$$

[10 marks]

**PSPM 2010/2011 SK027 (1a & 1b)**

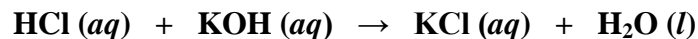
- 10 (a) The thermochemical processes involved in the formation of lithium fluoride, LiF, are given below. Write a complete thermochemical equation for each process.

Process	Enthalpy, $\Delta H$ (kJ mol <sup>-1</sup> )
Atomisation of lithium	+ 155
Atomisation of fluorine gas	+ 150
Ionisation of lithium atom	+ 520
Electron affinity for fluorine atom	- 382
Formation of lithium fluoride	- 594



[5 marks]

- (b) 100.00 cm<sup>3</sup> of 2.0 mol dm<sup>-3</sup> hydrochloric acid and 100.00 cm<sup>3</sup> potassium hydroxide solution, both at initial temperature of 30.0 °C are mixed in a calorimeter. The maximum temperature of the solution is 41.0 °C. Calculate the enthalpy of neutralization for the reaction.



$$\begin{aligned} \text{Heat of reaction} &= m_{\text{soln}}c_{\text{soln}}\Delta T \\ &= 200.0 \text{ g} \times 4.184 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1} \times 11.0 \text{ }^\circ\text{C} \\ &= 9204.8 \text{ J} \end{aligned}$$

$$\begin{aligned} \text{No. of mol of acid} &= (2.0 \text{ mol dm}^{-3} \times 100) / 1000 \\ &= 0.2 \text{ mol} \end{aligned}$$

$$\begin{aligned} 0.2 \text{ mol produced} &= 9204.8 \text{ J} \\ 1.0 \text{ mol produced} &= (1.0 / 0.2) \times 9204.8 \text{ J} \\ &= 46024 \text{ J mol}^{-1} \\ &= 46.024 \text{ kJ mol}^{-1} \end{aligned}$$

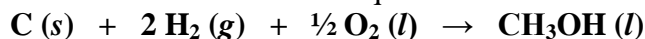
$$\therefore \text{Enthalpy of neutralization} = - 46.024 \text{ kJ mol}^{-1}$$

[5 marks]

**PSPM 2010/2011 TK027 (2a & 2b)**

11 (a) Standard enthalpies of formation for H<sub>2</sub>O, CO<sub>2</sub> and methanol, CH<sub>3</sub>OH are – 285.8 kJ mol<sup>-1</sup>, – 393.5 kJ mol<sup>-1</sup> and – 238.6 kJ mol<sup>-1</sup> respectively.

i. Write the thermochemical equation for the formation of CH<sub>3</sub>OH.



ii. Calculate standard enthalpy of combustion for CH<sub>3</sub>OH.

$$\begin{aligned} \Delta H^\circ_{\text{rxn}} &= \sum \Delta H^\circ_f \text{(product)} - \sum \Delta H^\circ_f \text{(reactant)} \\ &= [(-393.5) + 2(-285.8)] - [(-238.6) + (0)] \\ &= -726.5 \text{ kJ mol}^{-1} \quad @ \quad \text{algebraic method} \end{aligned}$$

(b) An amount of 120.0 mL of coffee in a well-insulated cup at 82 °C is too hot to drink. What volume of cold fresh milk at 15 °C need to be added to the coffee in order to achieve a temperature of 65°C? Assume specific heat capacities and densities of coffee and milk are the same as water.

$$\begin{aligned} Q_{\text{cold milk}} &= Q_{\text{hot coffee}} \\ \text{heat absorb by cold milk} &= \text{heat release from hot coffee} \\ m_{\text{milk}} \times c \times \Delta T &= m_{\text{coffee}} \times c \times \Delta T \\ m_{\text{milk}} \times 4.18 \times (65 - 15) &= 120 \times 4.18 (65 - 82) \\ m_{\text{milk}} &= 40.8 \text{ g} \\ \therefore \text{Volume of cold milk} &= 40.8 \text{ mL} \end{aligned}$$

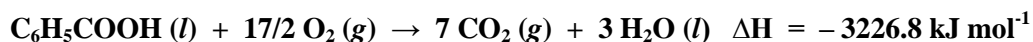
[6 marks]

**PSPM 2011/2012 SK026 (2 & 5b)**

12 (a) i. Define standard enthalpy of combustion.

**Heat released when 1 mole of substance/compound is completely combusted in excess oxygen at standard condition (25 °C and 1 atm).**

ii. The enthalpy of combustion of benzoic acid is – 3226.8 kJ mol<sup>-1</sup>. When 3.2 g benzoic acid, C<sub>6</sub>H<sub>5</sub>COOH is completely combusted in a bomb calorimeter containing 2.0 kg of water, the temperature of the water increased by 3.8 °C. Calculate the heat capacity of the calorimeter.



$$\begin{aligned} \text{no of mol of C}_6\text{H}_5\text{COOH} &= 3.2 \text{ g} / 122 \text{ g mol}^{-1} \\ &= 0.0262 \text{ mol} \end{aligned}$$

$$\begin{aligned} 1.0 \text{ mol C}_6\text{H}_5\text{COOH} &\text{ released } 3226.7 \text{ kJ} \\ 0.0262 \text{ mol C}_6\text{H}_5\text{COOH} &\text{ released } = \frac{0.0262 \text{ mol}}{1.0 \text{ mol}} \times -3226.7 \text{ kJ} \\ &= -84.63 \text{ kJ} \end{aligned}$$

$$Q_{\text{rxn}} = Q_{\text{soln}} + Q_{\text{cal}}$$

$$84.63 = [2000 \text{ g} \times 4.184 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1} \times 3.8] + [C_{\text{cal}} \times 3.8]$$

$$C_{\text{cal}} = 1.39 \times 10^4 \text{ J }^{\circ}\text{C}^{-1}$$

[6 marks]

**PSPM 2011/2012 TK025 (2 & 5b)**

- 13 To determine the enthalpy change of reaction between zinc and copper (II) sulphate, a student placed 100.0 cm<sup>3</sup> of 0.200 mol dm<sup>-3</sup> copper (II) sulphate solution into a polystyrene beaker. After two minutes, 1.20 g of zinc powder was added. The solution was stirred and the temperature recorded every half minute. The results obtained were then plotted to give the graph as shown in **FIGURE 1**.

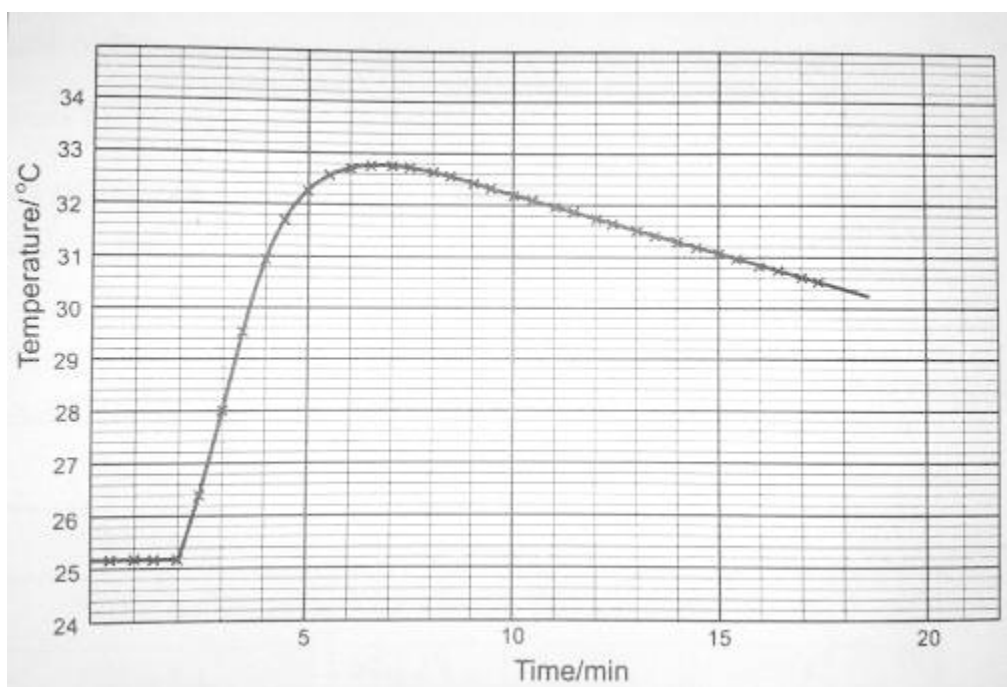


FIGURE 1

- (a) Write an equation for the reaction taking place in the polystyrene beaker.  
 **$\text{Zn (s)} + \text{CuSO}_4 \text{ (aq)} \rightarrow \text{ZnSO}_4 \text{ (aq)} + \text{Cu (s)}$**   
 [2 marks]
- (b) What is the highest temperature? Explain what is happening in the system at this point.  
**32.80 °C. The reaction reacted vigorously and released heat; and the rate of reaction increases/maximum at this point.**  
 [2 marks]

- (c) Calculate how much heat was evolved during the reaction in this experiment.

**Heat released = Heat absorbed**

$$\begin{aligned} Q_{\text{reaction}} &= Q_{\text{solution}} \\ &= m_{\text{soln}}c\Delta T \\ &= 100 \text{ g} \times 4.18 \text{ Jg}^{-1}\text{K}^{-1} \times (305.8 - 298.2) \text{ K} \\ &= 3176.8 \text{ J} @ \\ &= 3.177 \text{ kJ} \end{aligned}$$

[2 marks]

- (d) Calculate the enthalpy change for this reaction. If the accepted value for the enthalpy change of reaction is  $-218 \text{ kJ mol}^{-1}$ , suggest two reasons for these differences.

$$\begin{aligned} \text{No of mole of Zn} &= 1.2 \text{ g} / 65.4 \text{ g/mol} \\ &= 1.835 \times 10^{-2} \text{ mol (limiting reagent)} \end{aligned}$$

$$\begin{aligned} \text{No of mole of CuSO}_4 &= 0.2 \text{ M} \times 0.1 \text{ L} \\ &= 2.0 \times 10^{-2} \text{ mol (excess reagent)} \end{aligned}$$

$$\begin{aligned} 1.835 \times 10^{-2} \text{ mol released } &3.177 \text{ kJ} \\ \therefore 1.0 \text{ mol released} &= (1.0 / 1.835 \times 10^{-2}) \text{ mol} \times 3.177 \text{ kJ} \\ &= 173.1 \text{ kJ mol}^{-1} \end{aligned}$$

**Reasons:**

- Heat lost to the surrounding
- Reaction do not complete
- All reactant not pure

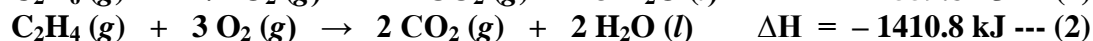
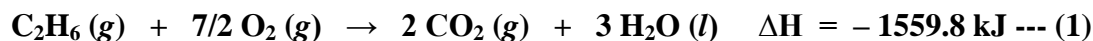
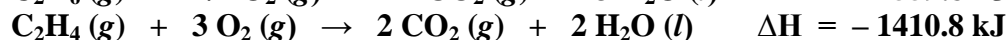
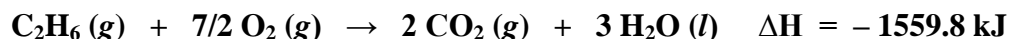
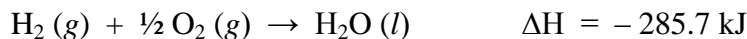
[Specific heat capacity of solution =  $4.18 \text{ Jg}^{-1}\text{K}^{-1}$ ]

[4 marks]

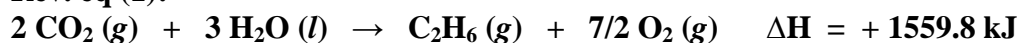
- 14 The heat evolved on combustion of 1.0 mol  $\text{C}_2\text{H}_6$  (g) and 1.0 mol  $\text{C}_2\text{H}_4$  (g) are 1559.8 kJ and 1410.8 kJ respectively. Write the balanced equation of the combustion reactions and calculate enthalpy of hydrogenation,  $x$  for the reaction below.



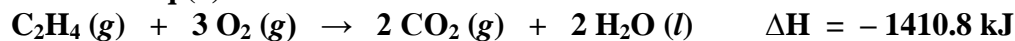
Given:



**Rev. eq (1):**



Maintain eq (2) :



Maintain eq (3)



[5 marks]

**PSPM 2012/2013 SK026**

- 15 Calculate the enthalpy of hydration of  $\text{Cl}^-$  in the dissolution process of  $\text{LiCl}$  in water using energy cycle method.

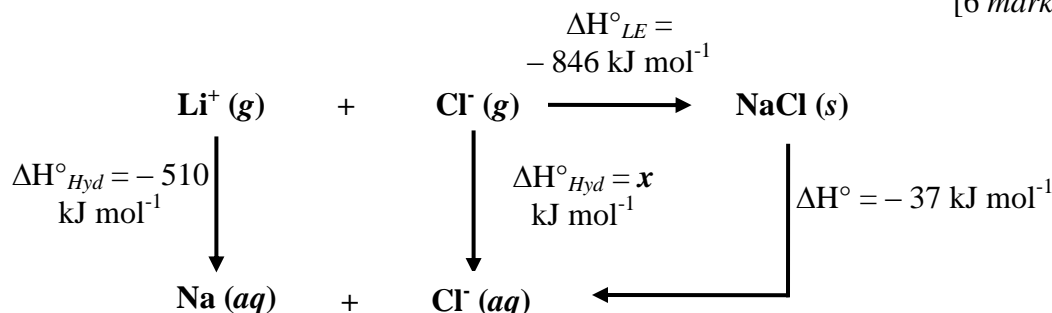
Given:

Lattice energy of  $\text{LiCl} = -846 \text{ kJ mol}^{-1}$

Enthalpy of solution of  $\text{LiCl}$  in water =  $-37 \text{ kJ mol}^{-1}$

Enthalpy of hydration of  $\text{Li}^+ = -510 \text{ kJ mol}^{-1}$

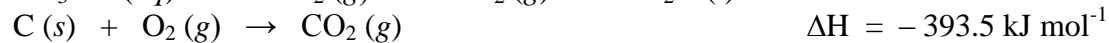
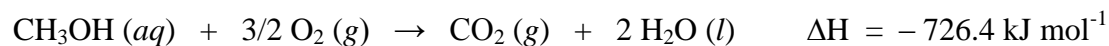
[6 marks]



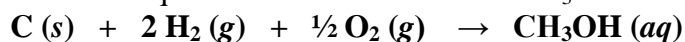
$$\begin{aligned}
 \Delta H^\circ_{LE} \text{ LiCl} &= \Delta H^\circ_{Hyd} \text{ Li}^+ + \Delta H^\circ_{Hyd} \text{ Cl}^- + \Delta H^\circ_{Soln} \\
 -846 &= (-510) + x + (+37) \\
 \Delta H^\circ_{Hyd} \text{ Cl}^- &= -373 \text{ kJ mol}^{-1}
 \end{aligned}$$

**PSPM 2012/2013 TK025 (2a & 2b)**

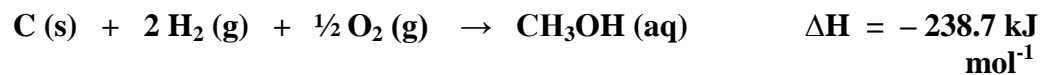
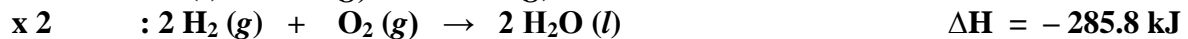
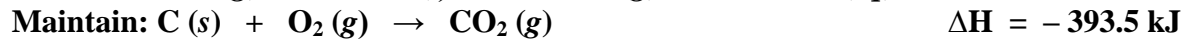
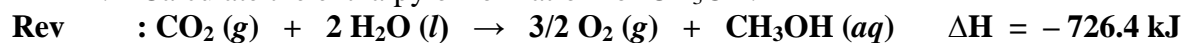
- 16 (a) Given the enthalpies of reactions below:



- i. Write the equation for the formation of  $\text{CH}_3\text{OH}$  from its elements.



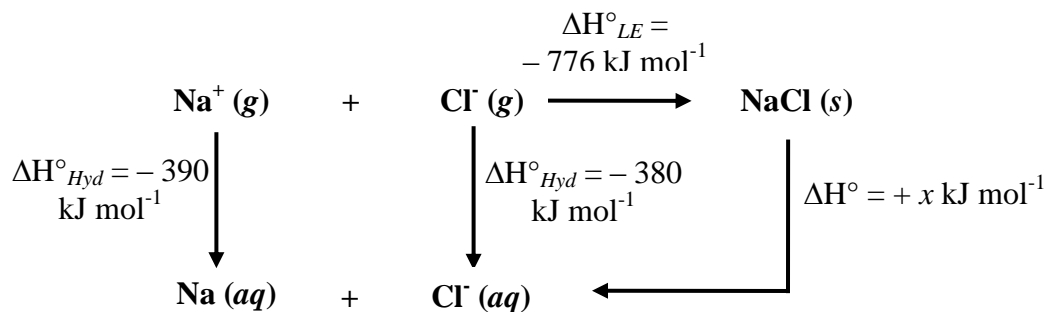
ii. Calculate the enthalpy of formation for CH<sub>3</sub>OH.



[5 marks]

(b) Construct an energy cycle diagram from the following data:

Lattice energy for NaCl	- 776 kJ mol <sup>-1</sup>
Hydration energy for Na <sup>+</sup>	- 390 kJ mol <sup>-1</sup>
Hydration energy for Cl <sup>-</sup>	- 380 kJ mol <sup>-1</sup>



[5 marks]