3055 BA SANGAM COLLEGE

PH: 6674003/9264117 E-mail: basangam@connect.com.fj

WORKSHEET 22

School: Ba Sangam College **Year:** 13
Subject: Chemistry **Name:**

Strand	**3- Reactions**
Sub strand	**3.2- Thermo chemistry**
Content Learning Outcome	**-Define calorimetry and identify the two types of calorimeter** **-Define standard heat of reaction, standard heat of formation and standard heat of combustion.** **-Perform calculations based on coffee-cup calorimeter.**

Standard Heat of Reaction (ΔH°)

Standard heat of reaction is the enthalpy change accompanying a reaction in which all the reactants and products are at 25 °C or 298 K and pressure of

1 atm or 760 mmHg or 101.3 kPa and involving moles specified by the coefficients of the equation.

$\Delta H^\circ\ (kJ) = \square H_{f\ products} - \square H_{f\ reactants}$

Thermochemical Equations

It is a balanced stoichiometric chemical equation that includes the enthalpy change and the specific physical states of the reactants and products.

Example

$N_{2(g)} + 2O_{2(g)} \rightarrow 2NO_{2(g)}\ \Delta H^\circ = +33.8\ kJ\ mol^{-1}$

Standard Heat of Formation (ΔH°f)

It is the amount of heat absorbed or released (enthalpy change) when 1 mole of a substance is formed at 25 °C and 1 atm pressure from its

elements in their standard states.

Example: Standard Heat of Formation of Water

$H_{2\ (g)} + ½\ O_{2(g)} \rightarrow H_2O_{(l)}\ \Delta H^\circ f = -285\ kJ\ mol^{-1}$

The formation of 1 mole of water (H2O) from the elements hydrogen and oxygen in their standard states releases 285 kJ of energy.

Standard Heat of Combustion

It is the amount of heat released (enthalpy change) when 1 mole of a substance is completely burned in oxygen under standard conditions.

Example:

$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}\ \Delta H^\circ C = -890\ kJ\ mol^{-1}$

The complete combustion of 1 mole of methane (CH_4) in excess oxygen gas releases 890 kJ of energy.

Hess's Law

Hess's law states that the change in enthalpy for any particular reaction depends only on the nature of the reactants and products and is independent of the number of steps or the pathway taken. It can be used to calculate the overall enthalpy change for a chemical reaction.

Example

Ref: 2019

10. Given below is the equation for the formation of ethanoic acid ($CH_3COOH_{(l)}$).

$$2C_{(s)} + 2H_{2(g)} + O_{2(g)} \longrightarrow CH_3COOH_{(l)} \quad \Delta H^\circ = \mathbf{-484\ kJ\ mol^{-1}}$$

Use **Hess's Law** and the equations given below to show that the **enthalpy change** (ΔH°) for the formation of ethanoic acid ($CH_3COOH_{(l)}$) is **-484 kJ mol⁻¹**.

1. $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \quad \Delta H^\circ = -394\ kJ\ mol^{-1}$
2. $H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(l)} \quad \Delta H^\circ = -286\ kJ\ mol^{-1}$
3. $CH_3COOH_{(l)} + 2O_{2(g)} \longrightarrow 2CO_{2(g)} + 2H_2O_{(l)} \quad \Delta H^\circ = -876\ kJ\ mol^{-1}$

(3 marks)

__Answer__:

1. $2C_{(s)} + \cancel{2O_{2(g)}} \longrightarrow 2CO_{2(g)}$ $\Delta H^\circ = -788\ kJ$] Eq.1 x 2 (1 mark)
2. $2H_{2(g)} + O_{2(g)} \longrightarrow \cancel{2H_2O_{(l)}}$ $\Delta H^\circ = -572\ kJ$] Eq. 2 x 2 (1 mark)
3. $2CO_{2(g)} + \cancel{2H_2O_{(l)}} \longrightarrow CH_3COOH_{(l)} + \cancel{2O_{2(g)}}$ $\Delta H^\circ = +876\ kJ$] Eq. 3 reversed (1mark)

$2C_{(s)} + 2H_{2(g)} + O_{2(g)} \longrightarrow CH_3COOH_{(l)} \quad \Delta H^\circ = -484\ kJ$

(½ mark for each correct equation and its respective energy value)

Ref: 2018

(a) Dinitrogen tetroxide (N_2O_4) can be used in space shuttle propellants. The equation below shows the formation of dinitrogen tetroxide from nitrogen dioxide (NO_2).

$$2NO_{2(g)} \longrightarrow N_2O_{4(g)} \qquad \Delta H^\circ = ?$$

Use Hess's Law and the equations below to calculate the **enthalpy change** (ΔH°) for the formation of dinitrogen tetroxide.

1. $NO_{2(g)} \longrightarrow \frac{1}{2} N_{2(g)} + O_{2(g)} \qquad \Delta H^\circ = -44$ kJ
2. $N_2O_{4(g)} \longrightarrow N_{2(g)} + 2O_{2(g)} \qquad \Delta H^\circ = -10$ kJ

(3 marks)

Answer

1. $2NO_{2(g)} \longrightarrow \cancel{N_{2(g)}} + \cancel{2O_{2(g)}}$ (½ mark) **$\Delta H^\circ = -88$ kJ**] Eq.1 x 2 (½ mark)
2. $\cancel{N_{2(g)}} + \cancel{2O_2} \longrightarrow N_2O_4(g)$ (½ mark) **$\Delta H^\circ = +10$ kJ**] Eq. 2 reversed (½ mark)

$2NO_{2(g)} \longrightarrow N_2O_{4(g)}$ $\qquad \Delta H^\circ = -78$ kJ (1 mark)

3, 2½, 2, 1½, 1, ½ or 0

Ref: 2017

(c) The standard enthalpy changes for the formation of aluminium(III) oxide and iron(III) oxide is given below.

I. $2Al_{(s)} + \frac{3}{2}O_{2(g)} \longrightarrow Al_2O_{3(s)} \qquad \Delta H^\circ = -1676$ kJ

II. $2Fe_{(s)} + \frac{3}{2}O_{2(s)} \longrightarrow Fe_2O_{3(s)} \qquad \Delta H^\circ = -824$ kJ

Use Hess's Law and the thermochemical equations above to calculate the **Heat of Reaction** (ΔH°) for the following reaction:

$$Fe_2O_{3(s)} + 2Al_{(s)} \longrightarrow Al_2O_{3(s)} + 2Fe_{(s)} \qquad \Delta H^\circ = ?$$

(3 marks)

Answer and Additional Notes

I. $2Al_{(s)} + \cancel{3/2O_{(g)}} \rightarrow Al_2O_{3(s)}$ $\qquad \Delta H^\circ = -1676$ kJ (½ mark)

Reversed

$Fe_2O_{3(s)} \rightarrow 2Fe_{(s)} + \cancel{3/2O_{2(g)}}$ (½ mark) $\qquad$ **$\Delta H^\circ = +824$ kJ** (½ mark)

$Fe_2O_{3(s)} + 2Al_{(s)} \rightarrow Al_2O_{3(s)} + 2Fe_{(s)}$ $\qquad \Delta H^\circ = -852$kJ (½ mark)

- Equation I - remains unchanged
- Equation II - reversed

Exercise – Pg: 135

1. In aqueous solution, potassium hydroxide reacts with hydrochloric acid as follows:

$$KOH_{(aq)} + HCl_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$$

In an experiment, 50 mL of 0.5 mol L^{-1} solution of KOH was rapidly mixed with 50 mL of 0.5 mol L^{-1} HCl in a polystyrene foam beaker.

Results:	
Initial temperature of $KOH_{(aq)}$	= 19.6 °C
Initial temperature of $HCl_{(aq)}$	= 19.6 °C
Final temperature of mixture	= 23.1 °C

a. State with a reason, whether the reaction is exothermic or endothermic.
b. Explain why the solutions were mixed rapidly.
c. Explain why a polystyrene foam beaker was used instead of a glass beaker.
d. Calculate the enthalpy change of this reaction in kJ mol^{-1}. (Assume specific heat capacity of the solution is the same as that of water, which is 4.18 J/g.°C).

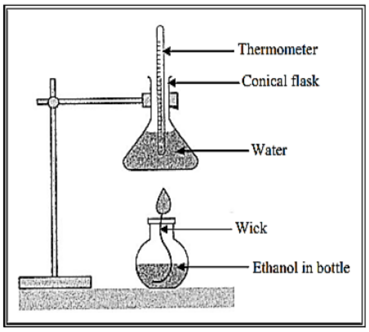
2. An experiment was carried out to measure the heat of solution of hydrated magnesium sulphate. The result of the experiment is recorded in the table below.

Mass of salt (g)	15.5
Mass of water (g)	100
Initial temperature (°C)	24
Final temperature (°C)	22.2

[M ($MgSO_{4(s)}$) = 120 g mol^{-1}; M ($MgSO_4.7H_2O$) = 246 g mol^{-1}]
Specific heat capacity of water = 4.18 J/g.°C.

a. Calculate the heat of solution of hydrated salt in kJ mol^{-1}.
b. Is the dissolving of hydrated magnesium sulphate an endothermic or exothermic process?
c. Give a reason for your answer to part (b) above.

3. A group of students measured the heat of combustion of ethanol using the equipment shown below.



The correct value for the heat of combustion of ethanol is 1.367 kJ mol^{-1}. The students' results show a heat of combustion of 0.254 kJ mol^{-1}.

a. Provide a reason for the difference in the value.
b. Suggest one way of improving the set up for a more accurate result.

Exercise – Pg: 140

1. Consider the following enthalpy data:

$\Delta H^o_c\ (H_{2(g)})$ = -282 kJ mol^{-1}
$\Delta H^o_f\ (H_2O_{(l)})$ = -286 kJ mol^{-1}

i. Explain fully what is meant by:
$\Delta H^o_c\ (H_{2(g)})$ = -282 kJ mol^{-1}
ii. Account for the fact that the two values of $\Delta H^o_f\ (H_2O_{(l)})$ and $\Delta H^o_c\ (H_{2(g)})$ are equal.
iii. What are the standard conditions for thermochemical measurements?

2. Consider the equations shown below:

$H_{(g)} + F_{(g)} \rightarrow HF_{(g)}$ ΔH^o = -568 kJ
$H_{2(g)} \rightarrow 2H_{(g)}$ ΔH^o = +436 kJ
$H_{2(g)} + F_{2(g)} \rightarrow 2HF_{(g)}$ ΔH^o = -221 kJ

Show how the above equations can be manipulated to give ΔH^o for the reaction: $F_{2(g)} \rightarrow 2F_{(g)}$

3. Consider the equations shown below:

$Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)}$ ΔH^o = -26.74 kJ
$CO_{(g)} + \frac{1}{2}O_{2(g)} \rightarrow CO_{2(g)}$ ΔH^o = -283.0 kJ

Use the thermochemical equations given above to calculate the value of ΔH^o for the reaction: $2Fe_{(s)} + 1\frac{1}{2}O_{2(g)} \rightarrow Fe_2O_{3(s)}$

4. Consider the following enthalpy data:

$\frac{1}{2}O_{2(g)} + H_{2(g)} \rightarrow H_2O_{(g)}$ ΔH^o = -285.8 kJ mol^{-1}
$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ ΔH^o = -393.5 kJ mol^{-1}
$CH_3OH_{(l)} + {}^3/_2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$ ΔH^o = -726.6 kJ mol^{-1}

Using the above information, calculate the heat of formation for methanol.

6. Consider the following enthalpy data:

$S_{(s)} + {}^3/_2O_{2(g)} \rightarrow SO_{3(g)}$ ΔH^o = -395.2 kJ mol^{-1}
$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$ ΔH^o = -198.2 kJ mol^{-1}

Calculate the enthalpy change for the formation of sulphur dioxide and state whether it is an exothermic or endothermic reaction.

7. Calculate the standard heat of reaction for the following reaction:

$2NH_{3(g)} + 3O_{2(g)} \rightarrow 2NO_{2(g)} + 2H_2O_{(g)}$

Given the following values of standard heat of formation:

$NH_{3(g)}$ ΔH^o_f = -46 kJ mol^{-1}
$NO_{2(g)}$ ΔH^o_f = +34 kJ mol^{-1}
$H_2O_{(g)}$ ΔH^o_f = -286 kJ mol^{-1}
$O_{2(g)}$ ΔH^o_f = 0 kJ mol^{-1}

8. When brandy is poured on a Christmas pudding and ignited, a pale blue flame is seen and the following reaction occurs:

$C_2H_5OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(g)}$

Given the following values of standard heat of formation, calculate the heat of combustion of ethanol.

$C_2H_5OH_{(l)}$ ΔH^o_f = -277 kJ mol^{-1}
$CO_{2(g)}$ ΔH^o_f = -393.5 kJ mol^{-1}
$H_2O_{(g)}$ ΔH^o_f = -286 kJ mol^{-1}

5. Consider the following enthalpy data:

$2C_{(s)} + H_{2(g)} \rightarrow C_2H_{2(g)}$ ΔH^o = +227 kJ mol^{-1}
$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ ΔH^o = -393.5 kJ mol^{-1}
$H_{2(g)} + ½O_{2(g)} \rightarrow H_2O_{(g)}$ ΔH^o = -286 kJ mol^{-1}

Using the above information, calculate the standard enthalpy of combustion of ethyne.