



**CHEMISTRY ONLINE**  
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# CHEMISTRY

## Physical Chemistry

Level & Board	AQA (A-LEVEL)
TOPIC:	ENERGETICS
PAPER TYPE:	SOLUTION - 1 (PAPER 1)
TOTAL QUESTIONS	10
TOTAL MARKS	68

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## Energetics - 1

1.

(a)

Standard enthalpy of combustion is the enthalpy change when one mole of a substance burns completely in oxygen. All substances should be in their standard states. It is denoted as  $\Delta H^\circ C$ .

(2)

(b)

Find the heat absorbed by the water using the formula:

$$q = mc\Delta T$$

Where:

$$m = \text{mass of water} = 150 \text{ g}$$

$$c = \text{specific heat capacity of water} = 4.18 \text{ J K}^{-1} \text{ g}^{-1}$$

$$q = m c \Delta T$$

$$= 150 \times 4.18 \times 13.9$$

$$= 8715.3 \text{ J}$$

$$\text{No. of moles of (propan-1-ol)} = 0.497/60.0$$

$$= 0.00828 \text{ mol}$$

$$\Delta H = -8.7153/0.00828$$

$$= -1050 \text{ kJ mol}^{-1}$$

(3)

(c)

The experiment might yield a less exothermic enthalpy of combustion due to factors like incomplete combustion, fuel evaporation, or deviations from standard experimental conditions.

(1)

2.

(a)

The term "enthalpy change" refers to the heat (energy) change that occurs at constant pressure. It is represented as  $\Delta H$

It represents the difference in enthalpy between the products and the reactants of a chemical reaction, measured at constant pressure.

(1)

(b)

To improve the experimental method and apparatus for determining the enthalpy change of the reaction between calcium carbonate and hydrochloric acid and to obtain more accurate data, consider the following improvements:

#### **Apparatus**

**Use a burette/pipette (instead of a measuring cylinder)**

- Burettes or pipettes allow for more precise volume measurements compared to measuring cylinders.

**Use a polystyrene cup (instead of a beaker) / insulate beaker**

- A polystyrene cup provides better insulation and minimizes heat exchange with the surroundings during the reaction. This reduces heat loss or gain, ensuring a more controlled environment for the reaction to occur and more accurate temperature measurements.

**Reweigh the watchglass after adding the solid:**

- Weighing the watch glass after adding the solid calcium carbonate helps in accurately determining the mass of the reactant used in the reaction.

**Use powdered solid:**

- Using powdered solid instead of solid chunks increases the surface area available for the reaction to occur, enhancing the reaction rate and making it more efficient.

#### **Temperature Measurements**

**Measure and record the initial temperature of the solution for a few minutes before addition**

- Recording the initial temperature over a few minutes allows for stability assessment. It helps identify any fluctuations in temperature before the reaction begins, ensuring the starting temperature is accurately established.
- Regular and frequent temperature measurements after the addition of calcium carbonate help capture the temperature change throughout the reaction. This comprehensive data collection allows for the observation of the reaction progress and the identification of the maximum temperature reached.

#### **Temperature Determination**

**Plot a graph of temperature against time**

- Plotting a graph of temperature against time helps visualize the temperature changes during the reaction. It provides a clear idea of how the temperature evolves over time, aiding in the interpretation of the reaction kinetics.

**Extrapolate to the point of addition:**

- Extrapolating the graph to the point of addition helps in determining the initial temperature of the reactants before the reaction began.

This extrapolation provides a more accurate starting point for calculating the temperature change ( $\Delta T$ ).

**Determine  $\Delta T$  at the point of addition:**

- By accurately determining the temperature change ( $\Delta T$ ) at the point of addition using the extrapolated data, it becomes possible to calculate the enthalpy change ( $\Delta H$ ) of the reaction more precisely.

Implementing these improvements in apparatus and data collection methods allows for more precise measurements, a more controlled environment, and a more comprehensive understanding of the temperature changes during the reaction.

(6)

(c)

Given:

- Volume of each solution =  $50.0 \text{ cm}^3$
- Concentration of hydrochloric acid (HCl) =  $0.500 \text{ mol dm}^{-3}$
- Concentration of sodium hydroxide (NaOH) =  $0.500 \text{ mol dm}^{-3}$
- Initial temperature of each solution =  $18.5 \text{ }^\circ\text{C}$

Determine the number of moles of each reactant:

The volume of each solution is  $50.0 \text{ cm}^3$

so the number of moles ( $n$ ) can be calculated using the formula:

$$n = \text{concentration} \times \text{volume}$$

No. of moles (HCl) or (NaOH)

$$= 50 \times 0.500 / 1000$$

$$= 0.025 \text{ moles}$$

$$q = -\Delta H \times n$$

$$= 57.1 \times 0.025$$

$$= 1.4275 \text{ kJ}$$

$$\Delta T = q/mc$$

$$\Delta T = (1.4275 \times 1000) / (100 \times 4.18)$$

$$= 3.4 \text{ }^\circ\text{C}$$

$$\text{Final Temperature} = 18.5 + 3.4 = 21.9 \text{ }^\circ\text{C}$$

(5)

(d)

To reduce the percentage uncertainty in the temperature change without changing the apparatus, increasing the concentration of the solutions would be a good approach.

(1)

3. (D)

(1)

4. (B)

(1)

5.

(a)

$$q = C_{cal} \cdot \Delta T$$

Given:

- Mass of hexane ( $m$ ) = 2.00 g
- Temperature change ( $\Delta T$ ) = 12.4 °C

First, we need to convert the mass of hexane to moles. Given the molar mass of hexane as 86.0g/mol, we can find the number of moles:

$$\begin{aligned} \text{Moles of hexane} &= 2/86 \\ &= 0.0233 \text{ mol} \end{aligned}$$

Now, we'll use the information that 1.00 mol of hexane releases 4154 kJ of energy to calculate the energy released by the combustion of 0.0233mol of hexane:

$$\text{Energy released} = \text{Number of moles} \times \text{Energy released per mole}$$

$$\begin{aligned} q &= 4154 \times 0.0233 \\ &= 96.6 \text{ kJ} \end{aligned}$$

$$\begin{aligned} C_{cal} &= q/\Delta T \\ &= 96.6 \text{ kJ}/12.4 \\ &= 7.79 \text{ (kJ K}^{-1}\text{)} \end{aligned}$$

(3)

(b)

Heat change calculated from the bomb calorimeter experiment is not an enthalpy change.

Because pressure is not constant in bomb calorimeter as enthalpy change requires constant pressure.

(1)

(c)

$$\begin{aligned} \text{Percentage Uncertainty} &= (\text{Uncertainty} / \text{Measurement}) \times 100\% \\ &= (0.2 / 12.2) \times 100\% = \\ &= 1.64\% \end{aligned}$$

To decrease the percentage uncertainty while using the same thermometer use bigger mass of fuel.

5(2)

6.

(a)

Given:

Initial temperature of the water = 19.1°C

Mass of water = 50.0 g

Specific heat capacity of water = 4.18 J K<sup>-1</sup> g<sup>-1</sup>Enthalpy of combustion of cyclohexane = -1216 kJ mol<sup>-1</sup>

Initial mass of spirit burner and cyclohexane = 192.730 g

Final mass of spirit burner and cyclohexane = 192.100 g

Relative molecular mass (Molecular weight) of cyclohexane = 84.0 g/mol

Calculate the mass of cyclohexane burned:

Initial mass of cyclohexane

= Initial mass of spirit burner and cyclohexane - Final mass of spirit burner and cyclohexane

Initial mass of cyclohexane = 192.730 g - 192.100 g

Initial mass of cyclohexane = 0.630 g

Determine the number of moles of cyclohexane burned:

Number of moles of cyclohexane = Mass / Molar mass

Number of moles of cyclohexane = 0.630 g / 84.0

Number of moles of cyclohexane = 0.0075 mol

Heat released = Enthalpy of combustion × Number of moles of cyclohexane

Heat released = 1216 × 1000 J/mol × 0.0075 mol

Heat released = 9120 J

Heat released = mcΔT

ΔT = Heat released / mc

9120 / 50 × 4.18

= 43.6°C

Final temperature = Initial temperature + ΔT

= 19.1 + 43.6

= 62.7°C

(4)



(b)

The smaller than expected temperature rise in the experiment might be due to:

- **Heat loss to surroundings:** Poor insulation or the absence of a lid leads to heat escaping.
- **Incomplete combustion:** Not all cyclohexane burns, producing less heat.
- **Evaporation:** Cyclohexane might evaporate before complete combustion, absorbing energy.

(1)

(c)

Given:

- $\Delta_c H^\circ$  for  $C(s) = -394 \text{ kJ/mol}$
- $\Delta_c H^\circ$  for  $H_2(g) = -286 \text{ kJ/mol}$
- $\Delta_c H^\circ$  for  $C_6H_{12}(l) = -3920 \text{ kJ/mol}$

So,

$$\Delta H = [6 \times (-394)] + [6 \times (-286)] + (-3920)$$

$$\Delta H = [-2364] + [-1716] - 3920$$

$$\Delta H = -4080 + 3920$$

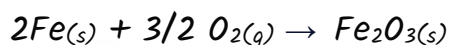
$$H = -160 \text{ kJ mol}^{-1}$$

(3)

7.

(a)

The enthalpy change of formation for iron(III) oxide ( $Fe_2O_3$ ) can be represented by the balanced chemical equation:



(1)

(b)

$$\Delta H_{\text{reaction}} = -19 \text{ kJ/mol}$$

$$\text{Enthalpy of formation of } \Delta_f H_{Fe_2O_3} = -822 \text{ kJ/mol}$$

$$\text{Enthalpy of formation of } \Delta_f H_{CO} = -111 \text{ kJ/mol}$$

For the enthalpy of formation of  $\Delta_f H_{CO_2}$

Using the equation:

$$(3 \times \Delta_f H_{CO_2}) = -19 + (-822) + 3(-111) - 0$$

$$(3 \times \Delta_f H_{CO_2}) = -19 - 822 - 333$$

$$(3 \times \Delta_f H_{CO_2}) = -1174 \text{ kJ/mol}$$

$$\Delta_f H_{CO_2} = -1174/3 = -391 \text{ kJ/mol}$$

(3)

(c)

Given:

$$\Delta H_1 = \Delta H_{(N_2 \rightarrow 2N)} = +944 \text{ kJ/mol}$$

$$\Delta H_2 = \Delta H_{(H_2 \rightarrow 2H)} = +436 \text{ kJ/mol}$$

$$\Delta H_{(\text{formation of ammonia})} = -92 \text{ kJ/mol}$$

To calculate the bond enthalpy for the N-H bond in ammonia ( $\text{NH}_3$ ), we use Hess's Law:

$$\Delta H_{(\text{formation of ammonia})} = \Delta H_1 + \Delta H_2 + \Delta H_{(N_2 + 3H_2 \rightarrow 2NH_3)}$$

$$-92 \text{ kJ/mol} = 944 \text{ kJ/mol} + (3 \times 436 \text{ kJ/mol}) + \Delta H_{(N_2 + 3H_2 \rightarrow 2NH_3)}$$

For  $\Delta H_{(N_2 + 3H_2 \rightarrow 2NH_3)}$ :

$$\Delta H_{(N_2 + 3H_2 \rightarrow 2NH_3)} = -92 \text{ kJ/mol} - 944 \text{ kJ/mol} - (3 \times 436 \text{ kJ/mol})$$

$$\Delta H_{(N_2 + 3H_2 \rightarrow 2NH_3)} = -92 \text{ kJ/mol} - 944 \text{ kJ/mol} - 1308 \text{ kJ/mol}$$

$$\Delta H_{(N_2 + 3H_2 \rightarrow 2NH_3)} = -2344 \text{ kJ/mol}$$

To find the energy for one N-H bond:

$$\text{Energy for one N-H bond} = -2344/6 \text{ kJ/mol}$$

$$\text{N-H bond} = -391 \text{ kJ/mol}$$

(3)

(d)

The bond enthalpy that is calculated is different from the mean bond enthalpy quoted in a data book ( $388 \text{ kJ mol}^{-1}$ ) because data book value is derived from (a number of) different compounds (not just different  $\text{NH}_3$  molecules)

(1)

8.

(a)

Given:

$$\text{Volume of aqueous solution of zinc sulfate} = 50 \text{ cm}^3 = 50 \text{ cm}^3 = 50 \times 10^{-3} \text{ dm}^3$$

$$\text{Concentration of aqueous zinc sulfate} = 1.00 \text{ mol dm}^{-3}$$

$$\text{Mass of magnesium used} = 2.08 \text{ g}$$

$$\text{Initial temperature} = 23.9^\circ\text{C}$$

$$\text{Maximum temperature} = 61.2^\circ\text{C}$$

$$\text{Specific heat capacity of the solution} = 4.18 \text{ J/(g}^\circ\text{C)}$$

$$\text{Density of the solution} = 1.00 \text{ g/cm}^3$$

Determine the number of moles of magnesium used:

$$\text{Molar mass of magnesium (Mg)} = 24.305 \text{ g/mol}$$

$$\text{Number of moles of Mg} = \text{Mass} / \text{Molar mass}$$

$$= 2.08 \text{ g} / 24.305 \text{ g/mol}$$



$$= 0.0856 \text{ mol}$$

Calculate the number of moles of zinc sulfate used:

Number of moles of zinc sulfate = Concentration  $\times$  Volume

$$\text{Number of moles of zinc sulfate} = 1.00 \text{ mol/dm}^3 \times 50 \times 10^{-3} \text{ dm}^3$$

$$= 0.050 \text{ mol}$$

So, Mg is in excess

Calculate the heat transferred using the formula  $Q = mc\Delta T$

$$Q = 50.0 \text{ g} \times 4.18 \text{ J/(g}^\circ\text{C)} \times 37.3^\circ\text{C}$$

$$= 7795.7 \text{ J}$$

$$Q = 7795.7 \text{ J}$$

$$= 7.795 \text{ kJ}$$

Energy released per mole is:

$$= 7.795 / 0.05 \text{ kJ mol}^{-1}$$

$$\Delta H = -156 \text{ kJ mol}^{-1}$$

(6)

(b)

The large difference between the student's experimental value and the data book value is due to heat loss (from the apparatus would mean the experimental value is smaller / lower / less exothermic)

(3)

(c)

For improving the experiment aim to enhance accuracy and precision in determining the enthalpy of reaction between magnesium and aqueous zinc sulfate.

**Improved Insulation**

**Insulating the Beaker or Using a Polystyrene Cup or Lid**

Insulating the reaction vessel helps in minimizing heat loss to the surroundings, ensuring that the observed temperature change is more reflective of the reaction's heat output.

**Improved Temperature Recording**

**Recording the Temperature Before Adding the Metal**

Capturing the initial temperature accurately before introducing the magnesium allows for a better and more accurate determination of the temperature change during the reaction.

**Recording Temperature Values at Regular Time Intervals**

Collecting temperature data at regular intervals throughout the reaction provides a comprehensive dataset to analyze and plot temperature changes over time.

**Plotting Temperature Results Against Time on a Graph**

Plotting the temperature-time graph gives the reaction's progress, aiding in identifying the maximum temperature and understanding the reaction kinetics.

Extrapolating the cooling curve back to the moment of metal addition helps estimate the initial maximum temperature reached, providing a theoretical value even if the exact moment wasn't captured.

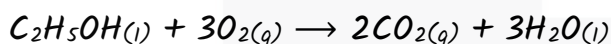
**Overall Impact:**

These improvements contribute to enhancing the precision and accuracy of the experimental data collected during the reaction. Insulation reduces heat loss, better temperature recording captures more comprehensive data, and improved analysis methods help estimate theoretical values more effectively.

(6)

9.

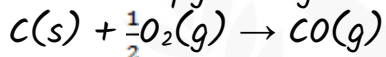
(a)



(2)

(b)

The enthalpy change for



represented as Standard enthalpy of formation.

It is difficult to prevent C reacting with O<sub>2</sub> to form some CO<sub>2</sub>

That is why this enthalpy change would be difficult to determine directly.

(2)

(c)

$$\Delta H = \sum \Delta H_c \text{ reactants} - \sum \Delta H_c \text{ products}$$

$$\Delta H = -393 - (-283)$$

$$\Delta H = -110 \text{ (kJ mol}^{-1}\text{)}$$

(2)

(d)

As XeF<sub>4</sub> has 4 bonds so:

$$4 (\text{Xe-F}) = 252 + (2 \times 158)$$

$$= 568$$

$$\text{Xe-F} = 568 / 4$$

$$\text{Xe-F} = 142 \text{ (kJ mol}^{-1}\text{)}$$

(3)

(e)

The value calculated above differs because mean bond enthalpy found by taking an average for Xe-F in a range of compounds.

(1)

10.

Using graph lines to calculate the temperature change at the 4th minute  
Is  $17^{\circ}\text{C}$

$$Q = mc\Delta T$$

$$Q = 10 \times 4.18 \times 17^{\circ}\text{C}$$

$$Q = 710.6 \text{ J}$$

$$\text{Mass of magnesium used} = 240 \text{ mg} = 0.24 \text{ g}$$

$$\text{Volume of aqueous trichloroethanoic acid used} = 10.0 \text{ cm}^3 = 10.0 \text{ g}$$

$$\text{Specific heat capacity of water} = 4.18 \text{ J/g}^{\circ}\text{C}$$

$$\text{Heat absorbed by the reaction} = 710.6 \text{ J}$$

$$\text{Moles of Mg} = 0.24 / 24.3$$

$$= 0.00988 \text{ mol}$$

$$\Delta H = 710.6 / 0.00988$$

$$= 71923.07 \text{ J mol}^{-1}$$

For  $\text{kJ mol}^{-1}$  dividing by 1000

$$\Delta H = -72 \text{ kJ mol}^{-1}$$

(7)

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I am Sorry !!!!!



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