

Unit A: Thermochemical Changes – Solutions

Question 1

The given information only pertains to the calorimetric **surroundings** of the reaction.

Given:

$$Q = 400 \text{ kJ}$$

$$\Delta T = T_f - T_i = 55.6^\circ\text{C} - 21.6^\circ\text{C} = 34.0^\circ\text{C}$$

$$c_{\text{H}_2\text{O}} = 4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}} *$$

With the given information, one should see that we just use the formula:

$$Q = m_{\text{H}_2\text{O}} c_{\text{H}_2\text{O}} \Delta T$$

Isolating for the mass of water:

$$m_{\text{H}_2\text{O}} = \frac{Q}{c_{\text{H}_2\text{O}} \Delta T} = \frac{400 \text{ kJ}}{(4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}})(34.0^\circ\text{C})} \doteq \mathbf{2.81 \text{ kg of H}_2\text{O}}$$

*Using the units of $\frac{\text{kJ}}{\text{kg}^\circ\text{C}}$ for $c_{\text{H}_2\text{O}}$ in the formula will allow one to obtain the mass of water in kilograms. These units are **equivalent** to the $\frac{\text{J}}{\text{g}^\circ\text{C}}$ units given in the data booklet.

The correct answer is **2.81**.

Question 2

It is helpful to separate the given information into a surroundings and system table:

Surroundings	System
$Q = ?$	$\Delta_r H_m = -851.5 \frac{\text{kJ}}{\text{mol}} *$
$m_{\text{H}_2\text{O}} = 7.40 \text{ kg}$	$n_{\text{Fe}_2\text{O}_3} = 1.00 \text{ mol}$
$c_{\text{H}_2\text{O}} = 4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}}$	$\Delta_r H = -Q$
$\Delta T = ?$	

*In the balanced chemical equation, the balancing coefficient in front of iron (III) oxide is 1, so the given enthalpy change can be considered as the molar enthalpy change for the reaction.

The Q for the surroundings is simply the reversed sign value of the enthalpy change of the system:

$$Q = 851.5 \text{ kJ}$$

The temperature change can now be obtained by the following formula rearrangement:

$$\Delta T = \frac{Q}{m_{\text{H}_2\text{O}} c_{\text{H}_2\text{O}}} = \frac{851.5 \text{ kJ}}{(7.40 \text{ kg})(4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}})} \doteq 27.5^\circ\text{C}$$

The correct answer is **27.5**.

Question 3

The given information only pertains to the calorimetric **surroundings** of the reaction.

Given:

$$m_{\text{H}_2\text{O}} = 250.0 \text{ g} = 0.2500 \text{ kg} *$$

$$c_{\text{H}_2\text{O}} = 4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}}$$

$$\Delta T = T_f - T_i = 35.0^\circ\text{C} - 20.0^\circ\text{C} = 15.0^\circ\text{C}$$

$$Q = m_{\text{H}_2\text{O}} c_{\text{H}_2\text{O}} \Delta T = (0.2500 \text{ kg}) \left(4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}} \right) (15.0^\circ\text{C}) \doteq 15.7 \text{ kJ}$$

*It is often helpful to convert a gram mass of water surroundings to a mass in kilograms.

Option C is correct! 15.7 kJ

Question 5

When it is said that a reaction “releases” energy, that reaction is **exothermic**. We can express exothermicity using a **negative ΔH** value or an **energy term on the products side** of a thermochemical equation.

Option B shows an energy on the reactants side so it is incorrect.

Options C and D are both exothermic reactions but the energy released in D is less than that of C so option D is incorrect.

Option A shows the formation of 2 moles of NaCl(s) from its elements. The data booklet lists $\Delta_f H_{m_{NaCl(s)}} = -411.2 \frac{kJ}{mol}$. This is the energy released when ONE mole of NaCl(s) is formed. Multiplying this number by 2 gives $-822.4 \frac{kJ}{mol}$. The energy released in the formation of 2 moles of NaCl(s) is less than what is shown in option C. Option A is incorrect.

Option C is correct!

Question 6

The decomposition reaction of a compound can be viewed as the **reverse of the formation** of that compound from its elements.

From the data booklet:

$$\Delta_f H_{m_{PCl_3(l)}} = -319.7 \frac{kJ}{mol}$$

Therefore:

$$\Delta_d H_{m_{PCl_3(l)}} = +319.7 \frac{kJ}{mol}$$

This is the energy required to decompose ONE mole of $PCl_3(l)$. The energy required to decompose 1.5 moles of $PCl_3(l)$ is simply 1.5 times the above decomposition enthalpy value:

$$\Delta_d H_{PCl_3(l)} = n_{PCl_3} \Delta_d H_{m_{PCl_3(l)}} = (1.5 \text{ mol}) \left(319.7 \frac{kJ}{mol} \right) \doteq 480 \text{ kJ}$$

Option A is correct! 480 kJ

Question 7

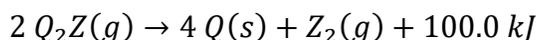
“**Burning**” reactions are always **exothermic**. Any option that has either an energy term on the reactants side or a positive ΔH value can be eliminated.

Options A, C, and D are incorrect.

Option B is correct!

Question 8

Reversing the chemical reaction shown gives:



The energy term in the reverse reaction is on the products side so this reaction is exothermic and its ΔH value should be negative.

Options C and D are incorrect.

The value's magnitude should be the same as the energy term in the reaction.

Option B is incorrect.

Option A is correct! ($\Delta H = -100.0 \text{ kJ}$)

Question 9

The more negative a reaction's ΔH value is, the more energy is released by that reaction.

	Fuel	Molar Enthalpy of Combustion	
As the number of carbons in the chain increases,	Methane	-802.5 kJ/mol	more and more energy is released.
	Ethane	-1 428.4 kJ/mol	
	Butane	-2 657.3 kJ/mol	
	Octane	-5 074.1 kJ/mol	

Option B is correct: "... the greater the amount of energy released during combustion."

Question 10

The given information only pertains to the **chemical system**.

Given:

$$\Delta_c H = -44.7 \text{ kJ}$$

$$m_{C_2H_5OH} = 1.65 \text{ g}$$

To calculate the molar enthalpy of combustion of ethanol, the number of moles of ethanol ($n_{C_2H_5OH}$) must be determined. Simply divide the mass of ethanol by the molar mass of ethanol ($M_{C_2H_5OH}$):

$$n_{C_2H_5OH} = \frac{m_{C_2H_5OH}}{M_{C_2H_5OH}} = \frac{1.65 \text{ g}}{46.08 \frac{\text{g}}{\text{mol}}} \doteq 0.0358073 \text{ mol}$$

The molar enthalpy of combustion can be calculated by dividing the given enthalpy change by the number of moles:

$$\Delta_c H_m = \frac{\Delta_c H}{n_{C_2H_5OH}} = \frac{-44.7 \text{ kJ}}{0.0358073 \text{ mol}} \doteq -1.25 \times 10^3 \frac{\text{kJ}}{\text{mol}}$$

Option D is correct!

Question 11

Important definitions to know:

Controlled Variable – The conditions we *keep the same* throughout an experiment.

Manipulated Variable – The *changing* pieces of an experiment.

Responding Variable – The values that are *measured* because of the changes.

In this experiment, the type of fuel is *changing* from $CH_4(g)$ to $C_3H_8(g)$. The **type of fuel** is the **manipulated variable**.

Both reactions must be done in the *same* calorimeter. The **type of calorimeter** is the **controlled variable**.

Any changes in the final temperature of the water will be *measured* because of changing the fuel. The **final temperature of the water** is the **responding variable**.

The only option that contains two of the above statements is **option A**.

Note: Options that refer to the “temperature change of the fuel” are incorrect. The fuel is the chemical system. The temperature change of a chemical system cannot be measured because its potential energy is changing, not its thermal or kinetic energy. The temperature change of the surroundings is measured.

Question 12

From the data booklet:

$$\Delta_f H_{m_{\text{AgI}(s)}} = -61.8 \frac{\text{kJ}}{\text{mol}}$$

This is the amount of heat released when ONE mole of AgI(s) is formed from its elements. The amount of heat released when 0.500 mol of AgI(s) is formed is simply half this amount:

$$\Delta_f H_{\text{AgI}(s)} = n_{\text{AgI}} \Delta_f H_{m_{\text{AgI}(s)}} = (0.5 \text{ mol}) \left(-61.8 \frac{\text{kJ}}{\text{mol}} \right) \doteq -30.9 \text{ kJ}$$

The correct answer is **30.9**.

Question 13

Recall that decomposition is the reverse of formation. From the data booklet:

$$\Delta_f H_{m_{\text{H}_2\text{O}(l)}} = -285.8 \frac{\text{kJ}}{\text{mol}}$$

Therefore:

$$\Delta_d H_{m_{\text{H}_2\text{O}(l)}} = +285.8 \frac{\text{kJ}}{\text{mol}}$$

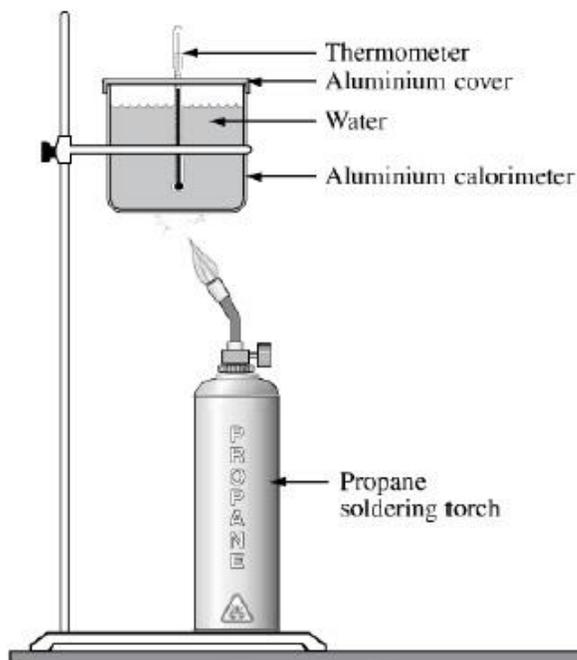
The decomposition of water is endothermic so any options that show energy terms on the products side can be eliminated. Options A and B are incorrect.

The above molar enthalpy for the decomposition of water is the energy required to decompose ONE mole of water. Option C uses the above value as an energy term and it also shows TWO moles of water decomposing. Option C is incorrect as well.

Option D correctly shows the energy term on the correct side and a single mole of water decomposing. **Option D is correct.**

Question 14

A technician performed an experiment to determine the molar enthalpy of combustion of propane in a soldering torch, as represented by the diagram below.



There is an awful lack of insulation in this experimental set up shown. Insulating the entire apparatus would lessen the amount of heat lost from the propane torch to the surrounding air and help to provide a more accurate determination of the molar enthalpy of combustion of propane. Creating an enclosing shield around the apparatus is the only option that provides better insulation, so **option B is correct.**

Question 15

Combustion is exothermic. Any option that shows a negative ΔH can be eliminated. Options B and C are incorrect.

Option D is incorrect because it shows products that would not be produced by combustion.

Option A is correct. The products are complete combustion products and the correct sign is on the ΔH value.

Question 16

It is helpful to separate the given information into a surroundings and system table:

Surroundings	System
$Q = ?$	$\Delta_c H_m = ?$
$m_{H_2O} = 250.0 \text{ g} = 0.2500 \text{ kg}$	$n_{C_{12}H_{22}O_{11}} = 0.0150 \text{ mol}$
$c_{H_2O} = 4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}}$	$\Delta_c H = -Q$
$\Delta T = +55.5^\circ\text{C}$	

The goal is determining $\Delta_c H_m$. As evident from the table, though, Q can be easily calculated:

$$Q = m_{H_2O} c_{H_2O} \Delta T = (0.2500 \text{ kg}) \left(4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}} \right) (55.5^\circ\text{C}) = 58.13625 \text{ kJ}$$

$\Delta_c H_m$ can be calculated through this sequence of operations:

$$\begin{aligned} \Delta_c H &= -Q = -58.13625 \text{ kJ} \\ \Delta_c H_m &= \frac{\Delta_c H}{n_{C_{12}H_{22}O_{11}}} = \frac{-58.13625 \text{ kJ}}{0.0150 \text{ mol}} = -3.88 \times 10^3 \frac{\text{kJ}}{\text{mol}} \end{aligned}$$

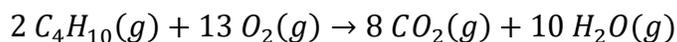
Option D is correct!

Question 17

$\Delta_c H_m$ is negative so this reaction is exothermic. The energy term in a thermochemical equation would be placed on the **products** side. The water in the calorimeter is the calorimetric surroundings. The change in temperature is reflective of the change in **kinetic** energy of the surroundings.

Option C is correct!

Question 18



The enthalpy change for this reaction can be determined using the modified Hess' Law where molar enthalpies of formation of compounds can be obtained from the data booklet:

$$\Delta_c H^\circ = \sum n \Delta_f H_m(\text{products}) - \sum n \Delta_f H_m(\text{reactants})$$

$$\sum n \Delta_f H_m(\text{products}) = (8 \text{ mol}) \left(-393.5 \frac{\text{kJ}}{\text{mol}} \right) + (10 \text{ mol}) \left(-241.8 \frac{\text{kJ}}{\text{mol}} \right) = -5566.0 \text{ kJ}$$

$$\sum n \Delta_f H_m(\text{reactants}) = (2 \text{ mol}) \left(-125.7 \frac{\text{kJ}}{\text{mol}} \right) + (13 \text{ mol}) \left(0 \frac{\text{kJ}}{\text{mol}} \right) = -251.4 \text{ kJ}$$

$$\Delta_c H^\circ = (-5566.0 \text{ kJ}) - (-251.4 \text{ kJ}) = -5314.6 \text{ kJ}$$

The multiplication of each compound's molar enthalpy of formation by its respective balancing coefficient is summed up for products and reactants. The molar enthalpy of formation of elements are always zero. It is helpful to calculate the product sum and the reactant sum separately before doing the final subtraction.

Note how the reaction shows 2 moles of butane reacting. The question requires calculating the **molar enthalpy** of combustion of butane which is the enthalpy change associated with ONE mole of butane reacting. The molar enthalpy will be half of our calculated value:

$$\Delta_c H_m = \frac{-5314.6 \text{ kJ}}{2 \text{ mol}} = -2657.3 \frac{\text{kJ}}{\text{mol}}$$

Option A is correct!

Question 19

It is helpful to separate the given information into a surroundings and system table:

Surroundings	System
$Q = ?$	$\Delta_{\text{sol}}H_m = ?$
$m_{\text{H}_2\text{O}} = 175.45 \text{ g} - 25.45 \text{ g} = 150.00 \text{ g}$ $= 0.15000 \text{ kg}$	$m_{\text{NH}_4\text{NO}_3} = 1.68 \text{ g}$
$c_{\text{H}_2\text{O}} = 4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}}$	$\Delta_{\text{sol}}H = -Q$
$\Delta T = 20.98^\circ\text{C} - 22.30^\circ\text{C} = -1.32^\circ\text{C}$	

The goal is determining $\Delta_{\text{sol}}H_m$. As evident from the table, though, Q can be easily calculated:

$$Q = m_{\text{H}_2\text{O}}c_{\text{H}_2\text{O}}\Delta T = (0.15000 \text{ kg})\left(4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}}\right)(-1.32^\circ\text{C}) = -0.82962 \text{ kJ}$$

$\Delta_{\text{sol}}H$ is opposite in sign to Q :

$$\Delta_{\text{sol}}H = -Q = -(-0.82962 \text{ kJ}) = +0.82962 \text{ kJ}$$

The moles of NH_4NO_3 ($n_{\text{NH}_4\text{NO}_3}$) must be calculated from the mass of NH_4NO_3 by dividing it by the molar mass of NH_4NO_3 ($M_{\text{NH}_4\text{NO}_3}$):

$$n_{\text{NH}_4\text{NO}_3} = \frac{m_{\text{NH}_4\text{NO}_3}}{M_{\text{NH}_4\text{NO}_3}} = \frac{1.68 \text{ g}}{80.06 \frac{\text{g}}{\text{mol}}} \doteq 0.020984 \text{ mol}$$

$\Delta_{\text{sol}}H_m$ can be calculated by dividing $\Delta_{\text{sol}}H$ by the moles of NH_4NO_3 :

$$\Delta_{\text{sol}}H_m = \frac{\Delta_{\text{sol}}H}{n_{\text{NH}_4\text{NO}_3}} = \frac{0.82962 \text{ kJ}}{0.020984 \text{ mol}} \doteq 39.5 \frac{\text{kJ}}{\text{mol}}$$

The correct answer is **39.5**.

Question 21

The main point is that the theoretical value for the molar enthalpy of combustion for methanol is higher than the one she obtained by experiment. This **must** be interpreted as her observing **less heat released** than what the theoretical value states.

$$\text{Experimental } |\Delta_c H_m| < \text{Theoretical } |\Delta_c H_m|$$

It is therefore best to deal with the absolute values of the enthalpy changes, not considering the negative sign in front of the ΔH values.

$$|\Delta_c H| = |Q|$$

One can go through each option and use numerical examples to determine which gives a lower $|\Delta_c H_m|$ value than the theoretical value for the molar enthalpy of combustion for methanol.

- A) Incorrect. A lower mass of methanol in the experiment means fewer moles of methanol reacted resulting in a higher experimental $|\Delta_c H_m|$ value.

Theoretical:

$$n_{CH_3OH} \rightarrow n_{CH_3OH}$$
$$|\Delta_c H_m| = \left| \frac{\Delta_c H}{n_{CH_3OH}} \right| = \left| \frac{-63.7 \text{ kJ}}{0.100 \text{ mol}} \right| = 637 \frac{\text{kJ}}{\text{mol}}$$

Experimental:

$$m_{\text{lower}} \rightarrow n_{\text{lower}}$$
$$|\Delta_c H_m| = \left| \frac{\Delta_c H}{n_{\text{lower}}} \right| = \left| \frac{-63.7 \text{ kJ}}{0.0800 \text{ mol}} \right| \doteq 796 \frac{\text{kJ}}{\text{mol}}$$

$$\text{Experimental } |\Delta_c H_m| > \text{Theoretical } |\Delta_c H_m|$$

- B) Incorrect. A higher final temperature in the experiment would widen the temperature change resulting in a higher experimental $|\Delta_c H|$ value.

Theoretical:

$$|\Delta_c H| = |m_{H_2O} c_{H_2O} (T_f - T_i)| = \left| (1.00 \text{ kg}) \left(4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}} \right) (35.2^\circ\text{C} - 20.0^\circ\text{C}) \right| \doteq 63.7 \text{ kJ}$$

Experimental:

$$|\Delta_c H| = |m_{H_2O} c_{H_2O} (T_f - T_i)| = \left| (1.00 \text{ kg}) \left(4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}} \right) (40.0^\circ\text{C} - 20.0^\circ\text{C}) \right| = 83.8 \text{ kJ}$$

$$\text{Experimental } |\Delta_c H| > \text{Theoretical } |\Delta_c H|$$

C) Incorrect. A lower initial temperature in the experiment would also widen the temperature change resulting in a higher experimental $|\Delta_c H|$ value.

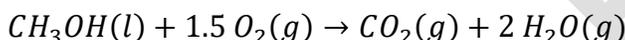
Experimental:

$$|\Delta_c H| = |m_{H_2O} c_{H_2O} (T_f - T_i)| = \left| (1.00 \text{ kg}) \left(4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}} \right) (35.2^\circ\text{C} - 15.2^\circ\text{C}) \right| = 83.8 \text{ kJ}$$

$$\text{Experimental } |\Delta_c H| > \text{Theoretical } |\Delta_c H|$$

D) **Correct!** Incomplete combustion in the experiment would **release less heat** than complete combustion. Here are modified Hess' Law calculations using molar enthalpies of formation to demonstrate how:

Theoretical:



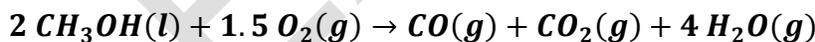
$$\Delta_c H^\circ = \sum n \Delta_f H_m(\text{products}) - \sum n \Delta_f H_m(\text{reactants})$$

$$\sum n \Delta_f H_m(\text{products}) = (1 \text{ mol}) \left(-393.5 \frac{\text{kJ}}{\text{mol}} \right) + (2 \text{ mol}) \left(-241.8 \frac{\text{kJ}}{\text{mol}} \right) = -877.1 \text{ kJ}$$

$$\sum n \Delta_f H_m(\text{reactants}) = (1 \text{ mol}) \left(-239.2 \frac{\text{kJ}}{\text{mol}} \right) + (1.5 \text{ mol}) \left(0 \frac{\text{kJ}}{\text{mol}} \right) = -239.2 \text{ kJ}$$

$$|\Delta_c H_m| = \left| \frac{(-877.1 \text{ kJ}) - (-239.2 \text{ kJ})}{1 \text{ mol CH}_3\text{OH}(l)} \right| = 637.9 \frac{\text{kJ}}{\text{mol}}$$

Experimental:



$$\Delta_{ic} H^\circ = \sum n \Delta_f H_m(\text{products}) - \sum n \Delta_f H_m(\text{reactants})$$

$$\sum n \Delta_f H_m(\text{products})$$

$$= (1 \text{ mol}) \left(-110.5 \frac{\text{kJ}}{\text{mol}} \right) + (1 \text{ mol}) \left(-393.5 \frac{\text{kJ}}{\text{mol}} \right) + (4 \text{ mol}) \left(-241.8 \frac{\text{kJ}}{\text{mol}} \right) \\ = -1471.2 \text{ kJ}$$

$$\sum n \Delta_f H_m(\text{reactants}) = (2 \text{ mol}) \left(-239.2 \frac{\text{kJ}}{\text{mol}} \right) + (1.5 \text{ mol}) \left(0 \frac{\text{kJ}}{\text{mol}} \right) = -478.4 \text{ kJ}$$

$$\Delta_{ic} H^\circ = (-1471.2 \text{ kJ}) - (-478.4 \text{ kJ}) = -992.8 \text{ kJ}$$

$$|\Delta_{ic} H_m| = \left| \frac{-992.8 \text{ kJ}}{2 \text{ mol CH}_3\text{OH}(l)} \right| = 496.4 \frac{\text{kJ}}{\text{mol}}$$

$$\text{Experimental } |\Delta_{ic} H_m| < \text{Theoretical } |\Delta_c H_m|$$

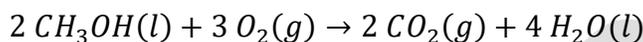
Question 22

The reaction is exothermic based on the negative ΔH value shown. Any option that states that energy is “absorbed” is incorrect. Therefore, options A and B are incorrect.

Option D is incorrect because $N_2H_4(l)$ is consumed, not produced, as the reaction states it.

Option C is correct! Energy is “released” as expected from an exothermic reaction, and $HNO_3(aq)$ is consumed as stated in the reaction.

Question 25



The enthalpy change for this reaction can be determined using the modified Hess’ Law where molar enthalpies of formation of compounds can be obtained from the data booklet.

NOTE: Liquid water is produced in this reaction as opposed to water vapour which is usually produced in a combustion reaction. Also, TWO MOLES of methanol are reacting.

$$\Delta_c H^\circ = \sum n \Delta_f H_m(\text{products}) - \sum n \Delta_f H_m(\text{reactants})$$

$$\sum n \Delta_f H_m(\text{products}) = (2 \text{ mol}) \left(-393.5 \frac{\text{kJ}}{\text{mol}} \right) + (4 \text{ mol}) \left(-285.8 \frac{\text{kJ}}{\text{mol}} \right) = -1930.2 \text{ kJ}$$

$$\sum n \Delta_f H_m(\text{reactants}) = (2 \text{ mol}) \left(-239.2 \frac{\text{kJ}}{\text{mol}} \right) + (3 \text{ mol}) \left(0 \frac{\text{kJ}}{\text{mol}} \right) = -478.4 \text{ kJ}$$

Because two moles of methanol are reacting as shown in the reaction, we must divide the $\Delta_c H^\circ$ result by 2 moles to get the molar enthalpy of combustion for the above reaction.

$$\Delta_c H_m = \frac{(-1930.2 \text{ kJ}) - (-478.4 \text{ kJ})}{2 \text{ mol } CH_3OH(l)} = -725.9 \frac{\text{kJ}}{\text{mol}}$$

Option A is correct!

Question 26

Consider the following points:

- 1) The ΔH° value shown in the reaction is the enthalpy change associated with TWO moles of methanol burning.
- 2) The number of moles that need to be burned to give an enthalpy change of -9.00×10^4 kJ must be determined.
- 3) This mole amount can then be converted into a mass amount in kilograms using the molar mass of methanol.

$$(-9.00 \times 10^4 \text{ kJ}) \left(\frac{2 \text{ mol } CH_3OH(l)}{-1275.8 \text{ kJ}} \right) \left(\frac{32.05 \text{ g}}{1 \text{ mol } CH_3OH(l)} \right) \left(\frac{1 \text{ kg}}{1000 \text{ g}} \right) \doteq 4.52 \text{ kg}$$

The correct answer is **4.52**.

Question 27

It is helpful to separate the given information into a surroundings and system table:

Surroundings	System
$Q = ?$	$\Delta_c H_m = \frac{-1275.8 \text{ kJ}}{2 \text{ mol CH}_3\text{OH}(l)} = -637.9 \frac{\text{kJ}}{\text{mol}}$
$m_{\text{H}_2\text{O}} = 250.0 \text{ g} = 0.2500 \text{ kg}$	$n_{\text{CH}_3\text{OH}} = ?$
$c_{\text{H}_2\text{O}} = 4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}}$	$\Delta_c H = -Q$
$\Delta T = 35.0^\circ\text{C} - 20.0^\circ\text{C} = 15.0^\circ\text{C}$	

The goal is determining the moles of methanol needed to cause that temperature change in the surrounding water ($n_{\text{CH}_3\text{OH}}$). As evident from the table, though, Q can be easily calculated:

$$Q = m_{\text{H}_2\text{O}} c_{\text{H}_2\text{O}} \Delta T = (0.2500 \text{ kg}) \left(4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}} \right) (15.0^\circ\text{C}) = 15.7125 \text{ kJ}$$

$\Delta_c H$ is opposite in sign to Q :

$$\Delta_c H = -Q = -15.7125 \text{ kJ}$$

To figure out $n_{\text{CH}_3\text{OH}}$, $\Delta_c H$ is simply divided by $\Delta_c H_m$:

$$n_{\text{CH}_3\text{OH}} = \frac{\Delta_c H}{\Delta_c H_m} = \frac{-15.7125 \text{ kJ}}{-637.9 \frac{\text{kJ}}{\text{mol}}} \doteq 2.46 \times 10^{-2} \text{ mol}$$

Option C is correct!

Question 34

In the potential energy diagram, it is seen that **potential energy (E_p) decreases** going from reactants to products. This is an **exothermic** reaction!

Options C and D are incorrect. Both options state this reaction as endothermic.

Option B is incorrect. The energy term for an exothermic reaction would not be placed on the reactants side of a thermochemical equation.

Option A is correct! The **energy term** would be placed on the **products side!**

Question 36

What is given in this diagram are the enthalpy changes that occur when the given compounds are formed from their elements.

Each of the formation enthalpies in this diagram go below the zero line. One should conclude that these formation enthalpies are negative.

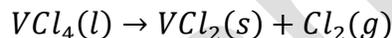
The quantity of each substance shown in the diagram is 1 mole (no balancing coefficient present), so the enthalpies given are **molar** formation enthalpies ($\Delta_f H_m$) measured in $\frac{kJ}{mol}$.

$$\Delta_f H_{m_{VCl_2(s)}} = -452.0 \frac{kJ}{mol}$$

$$\Delta_f H_{m_{VCl_4(l)}} = -569.4 \frac{kJ}{mol}$$

$$\Delta_f H_{m_{VCl_3(s)}} = -580.7 \frac{kJ}{mol}$$

The reaction of concern is:



The enthalpy change associated with this reaction can now be determined using a modified Hess' Law calculation:

$$\Delta_r H^\circ = \sum n \Delta_f H_m(\text{products}) - \sum n \Delta_f H_m(\text{reactants})$$

$$\sum n \Delta_f H_m(\text{products}) = (1 \text{ mol})(\Delta_f H_m[VCl_2(s)]) = -452.0 \text{ kJ} *$$

$$\sum n \Delta_f H_m(\text{reactants}) = \Delta_f H_m[VCl_4(l)] = -569.4 \text{ kJ}$$

$$\Delta_r H^\circ = -452.0 \text{ kJ} - (-569.4 \text{ kJ}) = 117.4 \text{ kJ}$$

*Remember that the $\Delta_f H_{m_{Cl_2(g)}} = 0$ and doesn't need to be included in the $\sum n \Delta_f H_m(\text{reactants})$.

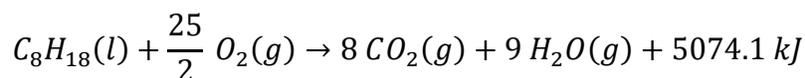
$\Delta_r H^\circ = 117.4 \text{ kJ}$ is associated with ONE MOLE of $VCl_4(l)$ reacting as seen in the balanced chemical equation.

Finally, the enthalpy change associated with 0.350 moles of $VCl_4(l)$ reacting needs to be determined:

$$\left(\frac{117.4 \text{ kJ}}{1 \text{ mol } VCl_4(l)} \right) (0.350 \text{ mol } VCl_4(l)) \doteq \mathbf{41.1 \text{ kJ}}$$

The correct answer is **41.1**.

Question 37



The above balanced chemical equation shows ONE MOLE of $C_8H_{18}(l)$ reacting. There is also an **energy term** on the **products side** of the reaction meaning this reaction is **exothermic!**

Options A and D are incorrect. The potential energy diagrams show an increase in potential energy one would expect from an endothermic reaction.

Option B is incorrect. Notice how the balancing coefficients have doubled. The ΔH value has not.

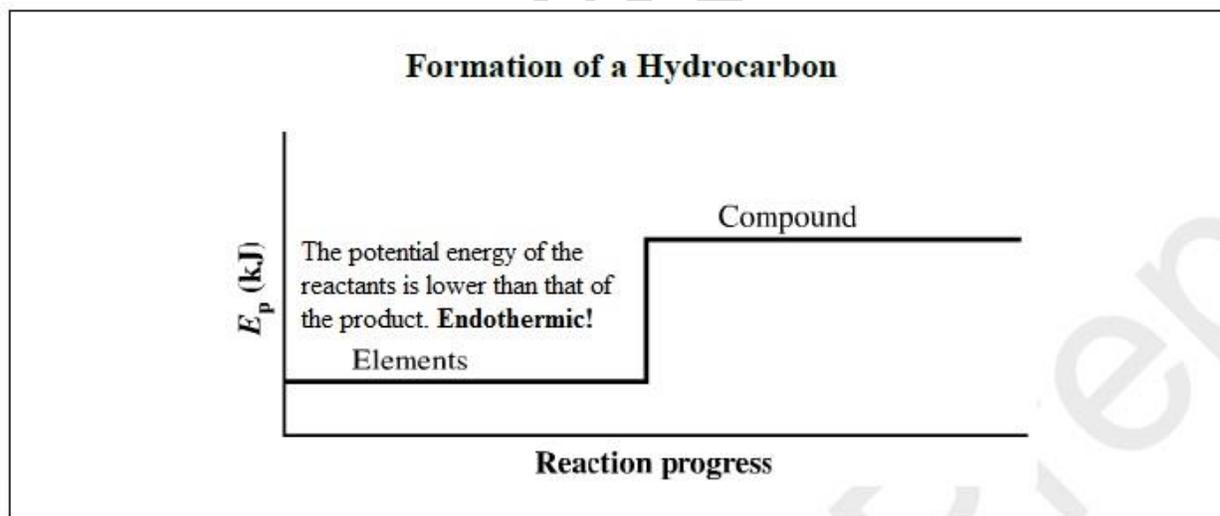
Option C is correct! The diagram shows a decrease in potential energy. The balancing coefficients have doubled and the ΔH value has doubled accordingly.

Question 38

A reaction that releases a large quantity of energy is **exothermic**. Options C and D are incorrect.

In an exothermic reaction, the **reactants have more potential energy** than the products. **Option B is correct!**

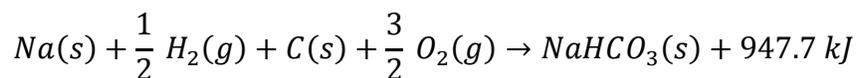
Question 39



The chemical potential energy diagram is for an **endothermic** formation of a compound from its elements. The **potential energy of the reactants is lower** than that of the product. The compound being formed should have a **positive** molar enthalpy of formation ($\Delta_f H^\circ$).

The only compound option with a positive $\Delta_f H^\circ$ is **option B, ethene**.

Question 40



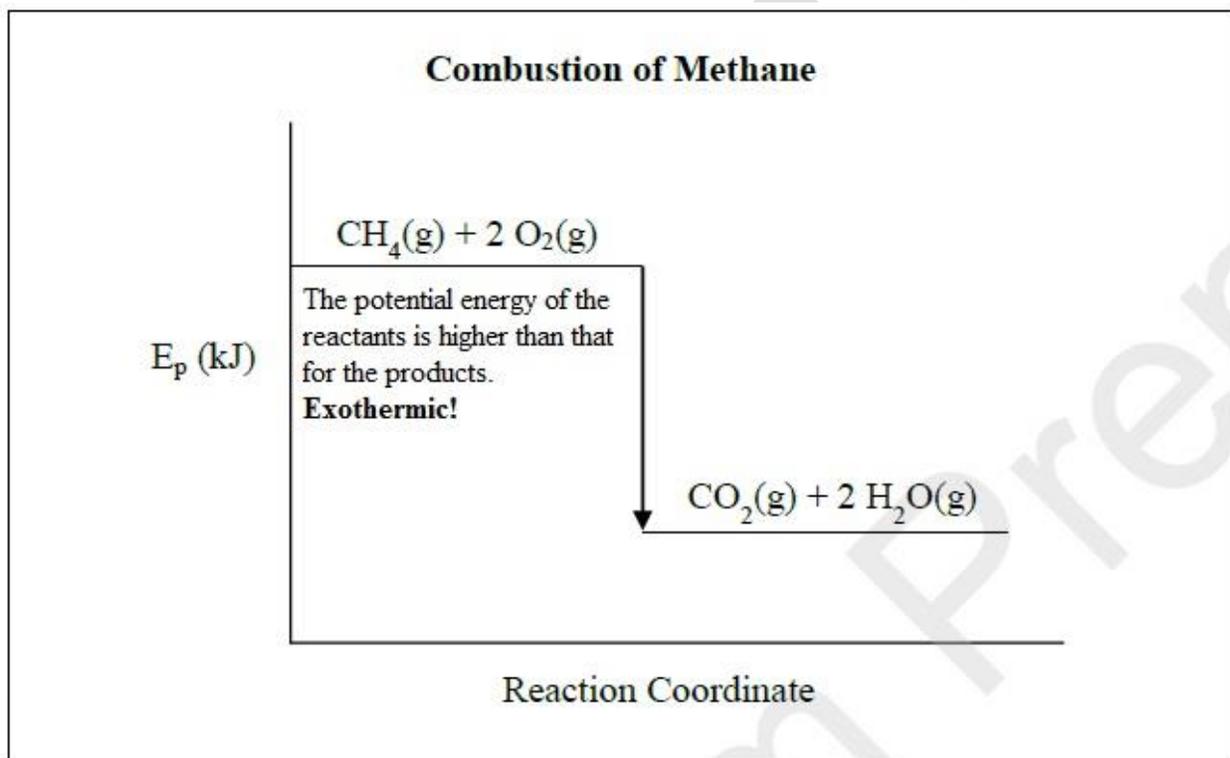
The balanced chemical equation has an **energy term on the products side** of the reaction. This is an **exothermic reaction!**

Options A and B are incorrect. It is in endothermic reactions that energy is absorbed from the surroundings.

Option D is incorrect. Kinetic energy changes of the system are not considered when looking at a chemical potential energy diagram.

Option C is correct! Energy is **released to** the surroundings because the bonds in the products contain **less potential** energy than the reactants.

Question 41



Option C is correct! An exothermic reaction's enthalpy change is represented by a negative ΔH value.

Question 42

During photosynthesis, the energy absorbed by the plant or microorganism originates from the Sun (i).

It is **chemical (ii)** energy that is stored in chemical bonds.

Option B is correct!

Question 44

Cellular Respiration:



The Combustion of Octane:



The product that the two reactions have in common that is the only option listed is $CO_2(g)$.

Option C is correct!

Question 45

Option B is incorrect. Cellular respiration does not require sunlight.

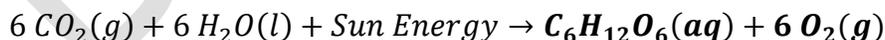
Option C is incorrect. Combustion is NOT endothermic.

Option D is incorrect. Respiration is NOT endothermic.

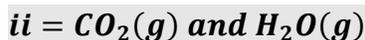
Option A is correct! Both combustion and cellular respiration produce $CO_2(g)$.

Question 47

Photosynthesis:



The complete combustion of octane, a hydrocarbon:



Note that in hydrocarbon combustion, one of the products is $H_2O(g)$, NOT $H_2O(l)$ as in cellular respiration.

Option C is correct!

Question 48

The combustion of methanol is an **exothermic reaction!**

Options A and D are incorrect. They show the chemical potential energy diagrams of endothermic reactions with the potential energies of the reactants being lower than the potential energy of the reactants.

Option B is incorrect. The enthalpy change shown is associated with the combustion of one mole of methanol. The balanced chemical equation shows two moles of methanol combusting. Note: See the solution for question 21.

Option C is correct! The chemical potential energy diagram shows the potential energy of the reactants being higher than the products. A modified Hess' Law calculation can show that the enthalpy change is correct for the combustion of two moles of methanol:

$$\Delta_c H^\circ = \sum n \Delta_f H_{m(\text{products})} - \sum n \Delta_f H_{m(\text{reactants})}$$

$$\sum n \Delta_f H_{m(\text{products})} = (2 \text{ mol}) \left(-393.5 \frac{\text{kJ}}{\text{mol}} \right) + (4 \text{ mol}) \left(-285.8 \frac{\text{kJ}}{\text{mol}} \right) = -1930.2 \text{ kJ}$$

$$\sum n \Delta_f H_{m(\text{reactants})} = (2 \text{ mol}) \left(-239.2 \frac{\text{kJ}}{\text{mol}} \right) + (3 \text{ mol}) \left(0 \frac{\text{kJ}}{\text{mol}} \right) = -478.4 \text{ kJ}$$

$$\Delta_c H^\circ = (-1930.2 \text{ kJ}) - (-478.4 \text{ kJ}) = -1451.8 \text{ kJ}$$

Question 49

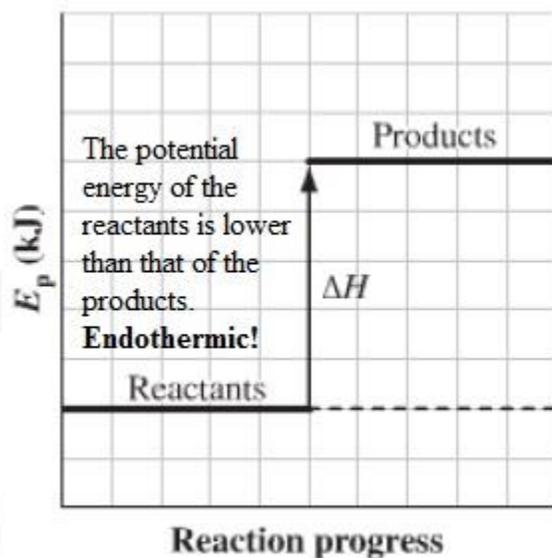
Option A is incorrect. Energy is a product in the chemical equation for an exothermic reaction.

Option B is incorrect. The temperature of the surroundings increases for an exothermic reaction.

Option C is incorrect. The enthalpy change for an exothermic reaction is negative.

Option D is correct! For an endothermic reaction, the potential energy of the products is greater than the potential energy of the reactants.

Question 50



This is the chemical potential energy diagram of an **endothermic reaction!** Correct options will have a **positive ΔH°** or an **energy term on the reactants side** of their chemical equations.

They are: **2, 4, 6, and 7.**

Question 51

Photosynthesis:



The reactants are $\text{CO}_2(g)$ (3) and $\text{H}_2\text{O}(l)$ (4). Note how the reactant water is in liquid form. In **living systems**, **water** as a reactant or product is in **liquid** form.

The complete combustion of octane, a hydrocarbon:



The products are $\text{CO}_2(g)$ (3) and $\text{H}_2\text{O}(g)$ (5). Note how the water product is in gaseous form. In **complete hydrocarbon combustion**, the **water** product is a **gas** unless otherwise specified.

The correct response is **3435.**

Question 52

The most thermally stable compound is the one with the most negative molar enthalpy of formation. Of the compounds that are listed, the one with the most negative molar enthalpy of formation is $\text{SiF}_4(g)$ at -1617 kJ/mol .

Option A is correct. The most thermally stable compound of the ones listed is $\text{SiF}_4(g)$.

Question 53

The original source of energy for ALL chemical processes on Earth is the Sun!

Option A is correct!

Question 54

To determine the enthalpy change per mole of carbon dioxide gas one must take each reaction's enthalpy change and divide it by its corresponding mole amount of $\text{CO}_2(\text{g})$:

$$1) \frac{5074.1 \text{ kJ}}{8 \text{ mol CO}_2(\text{g})} = 634.3 \frac{\text{kJ}}{\text{mol CO}_2(\text{g})}$$

$$2) \frac{2043.9 \text{ kJ}}{3 \text{ mol CO}_2(\text{g})} = 681.3 \frac{\text{kJ}}{\text{mol CO}_2(\text{g})}$$

$$3) \frac{1234.8 \text{ kJ}}{2 \text{ mol CO}_2(\text{g})} = 617.4 \frac{\text{kJ}}{\text{mol CO}_2(\text{g})}$$

$$4) \frac{802.5 \text{ kJ}}{1 \text{ mol CO}_2(\text{g})} = 802.5 \frac{\text{kJ}}{\text{mol CO}_2(\text{g})}$$

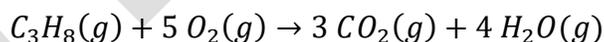
Ordering the results from smallest to largest one gets the correct answer of **3124** (NOT 3142).

Question 55

With no other pieces of information given, this enthalpy change can be calculated by the modified Hess' Law using molar enthalpies of formation from the data booklet where:

$$\Delta_c H^\circ = \sum n \Delta_f H_m(\text{products}) - \sum n \Delta_f H_m(\text{reactants})$$

The balanced chemical equation for the complete combustion of propane ($\text{C}_3\text{H}_8(\text{g})$) is:



In an open system, remember that it is water vapour that is formed, NOT liquid water. Also, note how in the balanced chemical equation, ONE MOLE of propane is shown to be reacting. After the calculations are carried out, $\Delta_c H^\circ = \Delta_c H_m$.

$$\sum n \Delta_f H_m(\text{products}) = (3 \text{ mol}) \left(-393.5 \frac{\text{kJ}}{\text{mol}} \right) + (4 \text{ mol}) \left(-241.8 \frac{\text{kJ}}{\text{mol}} \right) = -2147.7 \text{ kJ}$$

$$\sum n \Delta_f H_m(\text{reactants}) = (1 \text{ mol}) \left(-103.8 \frac{\text{kJ}}{\text{mol}} \right) + (5 \text{ mol}) \left(0 \frac{\text{kJ}}{\text{mol}} \right) = -103.8 \text{ kJ}$$

$$\Delta_c H_m = \frac{\Delta_c H^\circ}{n_{\text{C}_3\text{H}_8}} = \frac{-2147.7 \text{ kJ} - (-103.8 \text{ kJ})}{1 \text{ mol C}_3\text{H}_8} = \mathbf{-2043.9 \frac{\text{kJ}}{\text{mol}}}$$

Option B is correct!

Question 56

Remember:

Photosynthesis is endothermic. Cellular respiration is exothermic.

Option A is incorrect. The products of photosynthesis do not have less potential energy than the reactants. This would imply that photosynthesis is exothermic. The reactants of cellular respiration do not have do not have less potential energy than the products. This would imply that cellular respiration is endothermic.

Option B is incorrect. Photosynthesis does not trap energy in the form of glucose; it traps sunlight energy. Cellular respiration does not absorb energy; it releases energy.

Option C is incorrect. Photosynthesis is not exothermic. Cellular respiration is not endothermic.

Option D is correct! Photosynthesis requires energy, and cellular respiration produces energy.

Question 58

The exothermic reaction(s) should show either an energy term on the products side of the reaction equation or a negative ΔH value. Of the options listed, the only two that show either of these are **II** and **III**.

Option D is correct!

Question 59

This question requires a Hess' Law calculation where the intermediate reaction equations need to be reversed and/or multiplied by numbers. Remember that the same operations need to be carried out on their respective ΔH° values.

The first equation needs to be multiplied by 3:



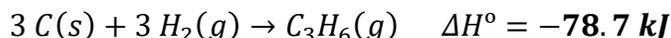
The second equation's direction needs to be reversed:



The third equation needs to be multiplied by 1.5:



Summing up the reactions while cancelling equivalent entities on the left and right as well as summing up the ΔH° values will give the overall reaction and a ΔH° of:



The correct answer is **78.7**.

Question 60



The ΔH° can simply be calculated using molar enthalpies of formation:

$$\Delta H^\circ = \sum n \Delta_f H_m(\text{products}) - \sum n \Delta_f H_m(\text{reactants})$$

$$\sum n \Delta_f H_m(\text{products}) = (2 \text{ mol}) \left(-436.5 \frac{\text{kJ}}{\text{mol}} \right) + (3 \text{ mol}) \left(0 \frac{\text{kJ}}{\text{mol}} \right) = -873.0 \text{ kJ}$$

$$\sum n \Delta_f H_m(\text{reactants}) = (2 \text{ mol}) \left(-397.7 \frac{\text{kJ}}{\text{mol}} \right) = -795.4 \text{ kJ}$$

$$\Delta H^\circ = (-873.0 \text{ kJ}) - (-795.4 \text{ kJ}) = -77.6 \text{ kJ}$$

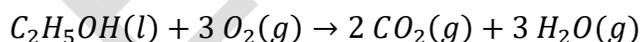
The correct answer is **77.6**.

Question 61

If the compounds **require energy** during their formation from their constituent elements, their **molar enthalpies of formation** should be **positive**. Looking up molar enthalpies of formation ($\Delta_f H^\circ$) in the data booklet of the options listed, the only compounds with a positive $\Delta_f H^\circ$ are: ethene (**2**), benzene (**3**), ethyne (**5**), and nitrogen dioxide (**6**). Remember that the options are to be listed in numerical order as numbered in the question diagram.

The correct answer is **2356**.

Question 63



The above reaction is the complete combustion of ethanol ($\text{C}_2\text{H}_5\text{OH}(l)$).

Combustion reactions are always **exothermic** (*ii* = "**exothermic**").

In exothermic reactions, energy is **released to** the surroundings (*i* = "**released to**").

Option A is correct!

Question 64

The molar enthalpy for the combustion of ethanol is calculated using molar enthalpies of formation:

$$\Delta_c H^\circ = \sum n \Delta_f H_{m(\text{products})} - \sum n \Delta_f H_{m(\text{reactants})}$$

$$\sum n \Delta_f H_{m(\text{products})} = (2 \text{ mol}) \left(-393.5 \frac{\text{kJ}}{\text{mol}} \right) + (3 \text{ mol}) \left(-241.8 \frac{\text{kJ}}{\text{mol}} \right) = -1512.4 \text{ kJ}$$

$$\sum n \Delta_f H_{m(\text{reactants})} = (1 \text{ mol}) \left(-277.6 \frac{\text{kJ}}{\text{mol}} \right) + (3 \text{ mol}) \left(0 \frac{\text{kJ}}{\text{mol}} \right) = -277.6 \text{ kJ}$$

$$\Delta_c H_m = \frac{\Delta_c H^\circ}{n_{\text{C}_2\text{H}_5\text{OH}}} = \frac{-1512.4 \text{ kJ} - (-277.6 \text{ kJ})}{1 \text{ mol C}_2\text{H}_5\text{OH}} = -1234.8 \frac{\text{kJ}}{\text{mol}}$$

Through the molar mass of ethanol ($M_{\text{C}_2\text{H}_5\text{OH}} = 46.08 \frac{\text{g}}{\text{mol}}$), the enthalpy change associated with the combustion of 1.50 g of ethanol can be determined:

$$(1.50 \text{ g C}_2\text{H}_5\text{OH}) \left(\frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.08 \text{ g C}_2\text{H}_5\text{OH}} \right) \left(\frac{-1234.8 \text{ kJ}}{1 \text{ mol C}_2\text{H}_5\text{OH}} \right) \doteq -40.2 \text{ kJ}$$

Therefore, **40.2 kJ** of energy is released when 1.50 g of ethanol is burned!

Question 67



This question requires a calculation of the reaction's ΔH° using molar enthalpies of formation to determine if the reaction is exothermic or endothermic. The molar enthalpies of formation for $\text{KO}_2(\text{s})$ and $\text{K}_2\text{CO}_3(\text{s})$ cannot be looked up in the data booklet but are given as:

$$\Delta_f H^\circ_{\text{KO}_2(\text{s})} = -284.9 \text{ kJ/mol}$$

$$\Delta_f H^\circ_{\text{K}_2\text{CO}_3(\text{s})} = -1152.7 \text{ kJ/mol}$$

$$\Delta H^\circ = \sum n \Delta_f H_{m(\text{products})} - \sum n \Delta_f H_{m(\text{reactants})}$$

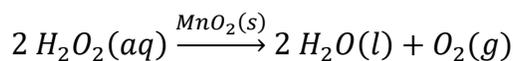
$$\sum n \Delta_f H_{m(\text{products})} = (2 \text{ mol}) \left(-1152.7 \frac{\text{kJ}}{\text{mol}} \right) + (3 \text{ mol}) \left(0 \frac{\text{kJ}}{\text{mol}} \right) = -2305.4 \text{ kJ}$$

$$\sum n \Delta_f H_{m(\text{reactants})} = (4 \text{ mol}) \left(-284.9 \frac{\text{kJ}}{\text{mol}} \right) + (2 \text{ mol}) \left(-393.5 \frac{\text{kJ}}{\text{mol}} \right) = -1926.6 \text{ kJ}$$

$$\Delta H^\circ = (-2305.4 \text{ kJ}) - (-1926.6 \text{ kJ}) = -378.8 \text{ kJ}$$

This reaction is **exothermic** with a **negative ΔH value**. **Option C is correct!**

Question 68



MnO₂(s) placed above the reaction arrow means that it acts as a **catalyst** for this reaction.

Option A is incorrect. Catalysts do not increase the activation energy; in fact, they lower the activation energy.

Options B and C are incorrect. The presence of a catalyst does not change the relative potential energies of the reactants and the products, nor the net amount of energy released or absorbed (ΔH°) by a reaction.

Option D is correct! Catalysts increase the rate of reaction by lowering the reaction's activation energy through providing an alternative pathway for the reaction.

Question 69

Remember, catalysts increase the rate of reaction by lowering the reaction's activation energy through providing an alternative pathway for the reaction. **Option 1 is correct.**

Catalysts do not change the relative potential energies of the reactants and the products, nor the ΔH° of a reaction. Options 2 and 4 are incorrect. **Option 3 is also correct.**

Therefore, **option B is correct!**

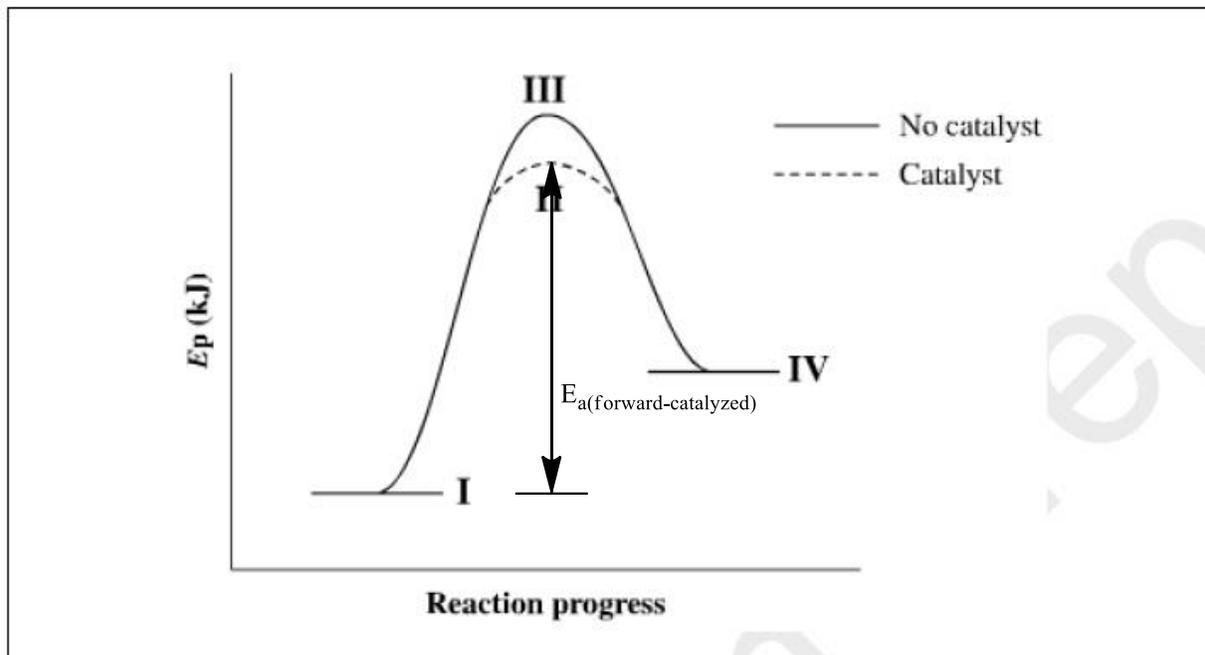
Question 70

Options A and B are incorrect. Both those statements imply that the activation energy for the reaction is increased due to the presence of a catalyst.

Option C is incorrect. Increasing the time of a reaction would mean that the rate of the reaction is decreased due to the presence of a catalyst.

Option D is correct! Decreasing the time of a reaction means that the reaction rate is increased due to the presence of a catalyst.

Question 71



The activation energy for the forward, catalyzed reaction [$E_{a(\text{forward-catalyzed})}$] goes from level I to level II.

$$E_{a(\text{forward-catalyzed})} = II - I$$

Option A is correct!

Question 72

Option A is correct! The **enthalpy change** for a catalyzed reaction is **THE SAME** as the enthalpy change for the uncatalyzed reaction.

Question 73



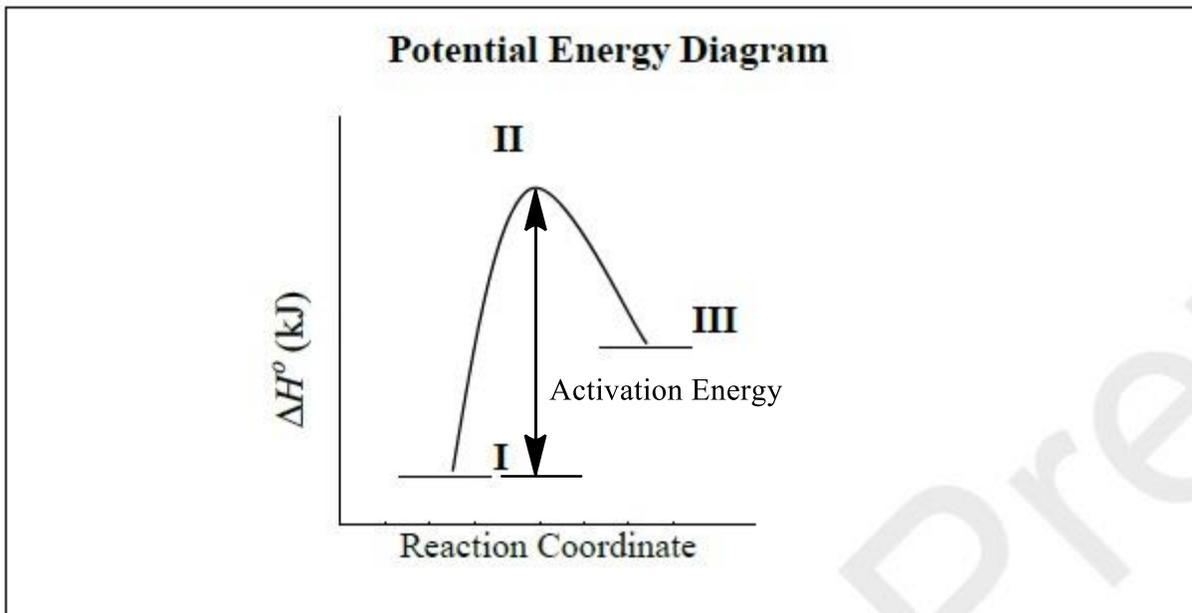
The energy term is on the products side of the balanced chemical equation meaning that this reaction is exothermic. An energy pathway diagram should show an overall decrease in E_p from reactants to products.

Option B is incorrect. The energy pathway diagram shows an overall increase in E_p from reactants to products.

Options A and D are incorrect. There are differences in the overall enthalpy changes between the catalyzed and uncatalyzed processes.

Option C is correct! There is an overall decrease in E_p from reactants to products. There is no difference in the overall enthalpy change between the catalyzed and uncatalyzed processes.

Question 74

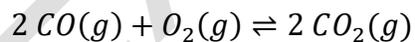


i = "activation energy"

ii = "I and II"

Option C is correct!

Question 75



Catalysts do not change the ΔH of a reaction so the “**energy transferred**” during the reaction **does not change (i)**.

Catalysts do not change the position of an equilibrium, but they do increase the rate at which equilibrium is reached! They do so by increasing the rates of BOTH the forward and reverse reactions. The **rate of the forward reaction is increased (ii)**.

Option C is correct!

Question 76

Photosynthesis:



The energy is on the reactants side. This reaction is endothermic! **2**

Cellular Respiration:



The energy is on the products side. This reaction is exothermic! **1**

Formation of Glucose:



From the data booklet:

$$\Delta_f H^\circ_{\text{C}_6\text{H}_{12}\text{O}_6(s)} = -1273.3 \text{ kJ/mol}$$

Because of the negative ΔH , this reaction is exothermic! **1**

The combustion of propane is exothermic! **1**

The correct response is **2111**.

Question 77

One should recognize that Dr. Richard Trotter's reaction is simply the complete combustion of methane. This reaction is exothermic!

Option C is incorrect. It is the energy pathway diagram of an endothermic reaction.

Option D is incorrect. There is a difference in the overall enthalpy change between the catalyzed and uncatalyzed processes. It is also the energy pathway diagram of an endothermic reaction.

Option B is incorrect. It shows the catalyzed process as having a higher activation energy.

Option A is correct! It is the energy pathway diagram of an exothermic reaction. There is no difference in the overall enthalpy change between the catalyzed and uncatalyzed processes.

Question 78

Only the **potential (i)** energy changes in a chemical reaction, such as the combustion of propane, are considered.

The changes in potential energy of a reaction result from a **rearrangement of bonds (ii)** as reactants transform into products.

Option A is correct!

Note: Only changes in the kinetic energy of the surroundings of a chemical reaction are considered in a calorimetry experiment. The kinetic energy changes result from an increase or decrease in molecular motion of the surrounding substance.

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