

## Thermochemistry Chemistry Questions with Solutions

**Q-1:** Given that standard molar enthalpies of formation of NO(g) and NO<sub>2</sub>(g) are respectively 90.3 kJ/mol and 33.2 kJ/mol. Calculate the enthalpy change for the reaction 2NO(g) + O<sub>2</sub>(g) → 2NO<sub>2</sub>(g).

**Answer:**  $\Delta H = \sum \Delta H_{\text{products}} - \sum \Delta H_{\text{reactants}}$

$$\Delta H = 2(33.2) - [2(90.3) + 0]$$

$$\Delta H = -114.2 \text{ kJ}$$

**Note:** Enthalpy of formation of free elements like O<sub>2</sub> is always 0.

**Q-2:** The enthalpy of formation of AgCl is obtained from the enthalpy change from which one of the following process?

- a)  $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$
- b)  $\text{Ag}(\text{s}) + \frac{1}{2} \text{Cl}_2(\text{g}) \rightarrow \text{AgCl}(\text{s})$
- c)  $\text{AgCl}(\text{s}) \rightarrow \text{Ag}(\text{s}) + \frac{1}{2} \text{Cl}_2(\text{g})$
- d)  $\text{Ag}(\text{s}) + \text{AuCl} \rightarrow \text{Au}(\text{s}) + \text{AgCl}(\text{s})$

**Answer: b)**  $\text{Ag}(\text{s}) + \frac{1}{2} \text{Cl}_2(\text{g}) \rightarrow \text{AgCl}(\text{s})$

Explanation: The enthalpy change accompanying the formation of one mole of a pure substance in its standard state from the constituents elements in their standard state is called enthalpy of formation. For Ag, the standard state is solid and for chlorine, it is gas. Thus, the process that gives the enthalpy of formation of AgCl is:  $\text{Ag}(\text{s}) + \frac{1}{2} \text{Cl}_2(\text{g}) \rightarrow \text{AgCl}(\text{s})$

**Q-3:**  $\Delta H$  and  $\Delta E$  for the reaction,  $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + \text{H}_2\text{O}(\text{l})$  at constant temperature are related as

- a)  $\Delta H = \Delta E$
- b)  $\Delta H = \Delta E + RT$
- c)  $\Delta H = \Delta E + 3RT$
- d)  $\Delta H = \Delta E - 3RT$

**Answer: d)**  $\Delta H = \Delta E - 3RT$

Explanation: For any chemical reaction,

$$\Delta H = \Delta E + \Delta_{\text{ng}}RT$$

Where,  $\Delta_{\text{ng}}$  = total number of moles of gaseous product - total number of moles of gaseous reactants.

For the given reaction,

$$\Delta_{ng} = 0 - 3 = -3$$

$$\Delta H = \Delta E + (-3)RT$$

$$\Delta H = \Delta E - 3RT$$

**Q-4:** For the reaction,  $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ , compute the entropy change (in J/K/mol) for the process and comment on the sign of the property.

Data	Species	$NH_3(g)$	$N_2(g)$	$H_2(g)$
	$S^\circ(J/K/mol)$	192.3	191.5	130.6

**Answer:**  $\Delta S^\circ = \sum \Delta S^\circ_{\text{products}} - \sum \Delta S^\circ_{\text{reactants}}$

$$\Delta S^\circ = 2 \times 192.3 - [191.5 + 3(130.6)] = -198.7 J/K/mol$$

In the reaction, two gaseous species are there in the reactant side and the number of moles of gaseous species on the product side is 1. So there is a decrease in the number of gaseous species which is represented by the negative sign of entropy.

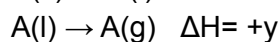
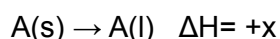
**Q-5:** The rate of evaporation of a liquid is always faster at a higher temperature because

- The enthalpy of vaporisation is always endothermic
- The enthalpy of vaporisation is always exothermic
- The enthalpy of vaporisation is zero
- The internal pressure of the liquid is less than that of the gas

**Answer:** a) The enthalpy of vaporisation is always endothermic

Explanation: At the time of evaporation, energy is required to overcome the intermolecular forces between the molecules of liquid, that is, the enthalpy of vaporisation is always endothermic. Since at high temperatures more energy is available, the rate of evaporation is faster at elevated temperatures.

**Q-6:** Given that



The heat of sublimation of A will be

- a)  $x-y$
- b)  $x+y$
- c)  $x$  or  $y$
- d)  $-x+y$

**Answer: b)  $x+y$**

Explanation: It is a physical process by which a substance is directly converted from solid to gaseous.

The reaction for heat of sublimation is  $A(s) \rightarrow A(g)$  which is obtained by adding  $A(s) \rightarrow A(l)$  and  $A(l) \rightarrow A(g)$ . Thus, the heat of sublimation of A will be  $x+y$ .

**Q-7:**  $\Delta H = -25 \text{ kcal}$  for the reaction  $\text{CH}_4(g) + \text{Cl}_2(g) \rightarrow \text{CH}_3\text{Cl}(g) + \text{HCl}(g)$ .

Bond	Bond Energy(kCal)
C-Cl	84
H-Cl	103
C-H	$x$
Cl-Cl	$y$
$x:y$	9:5

What is the bond enthalpy of Cl-Cl bond? Use the given data to calculate the answer.

**Answer:**  $\Delta H = \sum \Delta B.E_{\text{reactants}} - \sum \Delta B.E_{\text{products}}$   
 $\Delta H = 4(\text{C-H}) + \text{Cl-Cl} - [3(\text{C-H}) + 1(\text{C-Cl}) + 1(\text{H-Cl})]$   
 $-25 = 4x + y - [3x + 84 + 103]$

This implies,  $x + y - 187 = -25$   
 Or  $x + y = 162$

**Given:**  $x:y = 9:5$ . Therefore,  $x = 9y/5$

Substituting  $x = 9y/5$  in  $x + y = 162$ , we get,  $y = 57.85 \text{ kCal}$

**Q-8:** For the reaction,  
 $\text{N}_2(g) + 2\text{O}_2(g) \rightarrow 2\text{NO}_2(g) \quad \Delta H = -66 \text{ kJ}$

Calculate the value of  $\Delta_f H$  of  $\text{NO}_2$ .

**Answer:** Enthalpy of formation( $\Delta_f H$ ) is the enthalpy change accompanying the formation of one mole of a pure substance in its standard state from the constituent elements in their standard state.

The given  $\Delta H$  of  $\text{NO}_2$  is for its two mole formation. Since we need to calculate the  $\Delta_f H$  of  $\text{NO}_2$  for one mole by definition, therefore it will be equal to

$$\Delta_f H = \Delta H/2 = -66/2 = -33\text{kJ}$$

**Q-9:** Which of the following is not an endothermic reaction?

- a) Decomposition of water
- b) Conversion of graphite to diamond
- c) Combustion of methane
- d) Dehydrogenation of ethane to ethylene

**Answer: c)** Combustion of methane

Explanation: Heat is produced in any combustion reaction. The reactions that involve the release of heat are said to be exothermic.

**Q-10:** The enthalpy changes of the following reactions at  $27^\circ\text{C}$  are

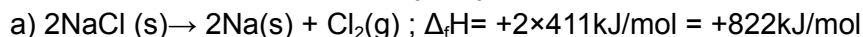
- 1)  $\text{Na(s)} + \frac{1}{2} \text{Cl}_2(\text{g}) \rightarrow \text{NaCl(s)} ; \Delta_f H = -411\text{kJ/mol}$
- 2)  $\text{H}_2(\text{g}) + \text{S(s)} + 2\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{SO}_4(\text{l}) ; \Delta_f H = -811\text{kJ/mol}$
- 3)  $2\text{Na(s)} + \text{S(s)} + 2\text{O}_2(\text{g}) \rightarrow \text{Na}_2\text{SO}_4(\text{s}) ; \Delta_f H = -1382\text{kJ/mol}$
- 4)  $\frac{1}{2} \text{H}_2(\text{g}) + \frac{1}{2} \text{Cl}_2(\text{g}) \rightarrow \text{HCl(g)} ; \Delta_f H = -92\text{kJ/mol}$

From the above data, the heat change of reaction at constant volume (in kJ/mol) at  $27^\circ\text{C}$  for the process

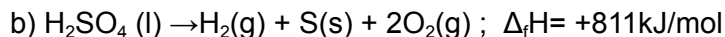


**Answer:**

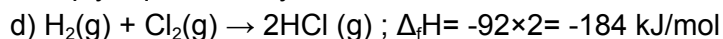
Reverse equation 1 and multiply it by 2



Reverse equation 2



Multiply equation 4 by 2



On adding equation, a, b, c, and d, we get



$$\Delta_f H = 822 + 811 + (-1382 - 184) = 67 \text{ kJ/mol}$$

We know that,  $\Delta_f H = \Delta U + \Delta n_g RT$

Where,  $\Delta n_g$  = total number of moles of gaseous product- total number of moles of gaseous reactants.

For the reaction given,  $\Delta n_g = 2$

$$R = 8.3 \text{ J/K mol} = 0.0083 \text{ kJ/mol}$$

$$T = 27^\circ\text{C} = 300 \text{ K}$$

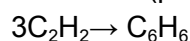
$\Delta U$  is the heat change of reaction at constant volume

On substituting the values, we get

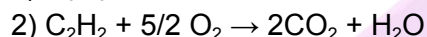
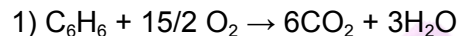
$$67 = \Delta U + 2 \times 0.0083 \times 300$$

$$\Delta U = 62.02 \text{ kJ/mol}$$

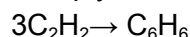
**Q-11:** Heat of combustion for benzene and acetylene are -3900 and -642 joule. Then calculate the heat of reaction(per mole) for the following reaction.



**Answer:**



Multiply reaction 2 with 3 and reverse reaction 1 and add them, we get



$$\Delta_f H = 3(\Delta H_c)_{\text{acetylene}} + (\Delta H_c)_{\text{benzene}} = 3(-642) + 3900 = 1974 \text{ J for 3 mole}$$

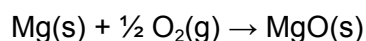
$$\text{Therefore, } \Delta_f H \text{ for 1 mole} = 1974/3 = 658 \text{ J/mol}$$

**Q-12:** For combustion of one mole of magnesium in an open container at 300 K and 1 bar pressure,  $\Delta_c H = -601.70 \text{ kJ/mol}$ , the magnitude of change in internal energy for the reaction is \_\_\_\_\_ kJ.

(Nearest integer)

Given:  $8.3 \text{ J/Kmol}$

**Answer:** The combustion reaction for one mole of magnesium is



$$\Delta_c H = \Delta E + \Delta n_g RT$$

Where,  $\Delta n_g$  = total number of moles of gaseous product- total number of moles of gaseous reactants.

$\Delta E$  is the internal energy

For the given reaction,  $\Delta n_g = -\frac{1}{2}$

$$\Delta E = -601.70 \text{ kJ/mol} - \frac{1}{2} (8.3 \times 10^{-3} \times 300) = 599.455 \text{ kJ} = 600 \text{ kJ}$$

**Q-13:** NaOH(s) has a heat of solution of -42.6 kcal/mol NaOH. When NaOH is dissolved in water, the temperature of the solution

- a) increases
- b) decreases
- c) remains constant
- d) can't be predicted

**Answer:** a) increases

**Q-14:** What is Hess's Law?

**Answer:** Hess's law is an important result of the first law of thermodynamics. It states that the enthalpy change in a chemical or physical process is the same whether it is performed in one step or several steps.

**Q-15:** If total enthalpies of reactants and products are  $H_R$  and  $H_P$  respectively, then for an endothermic reaction

- a)  $H_R < H_P$
- b)  $H_R > H_P$
- c)  $H_R = H_P$
- d) Data insufficient

**Answer:** a)  $H_R < H_P$

Explanation: For an endothermic reaction, enthalpy of reaction is positive. Enthalpy of reaction is calculated by using below equation:

$$\Delta H = \sum \Delta H_P - \sum \Delta H_R$$

For  $\Delta H$  to be positive,  $H_R$  must be less than  $H_P$ .

## Practise Questions on Thermochemistry

**Q-1:** Which of the following reactions does not represent  $\Delta_f H$ ?

- a)  $C(s) + O_2 \rightarrow CO_2$
- b)  $Br_2(l) + H_2(g) \rightarrow 2HBr$
- c)  $CO + O_2 \rightarrow CO_2$
- d)  $C(s) \rightarrow C(g)$

**Answer:** b) and c)

**Q-2:** Which of the following pairs are correctly matched?

i) Arrhenius Equation	Variation of enthalpy of a reaction with temperature
ii) Kirchhoff equation	Variation of rate constant with temperature
iii) Second law of thermodynamics	Entropy of an isolated system tends to increase and reach a maximum value
iv) Hess's law of constant heat summation	Enthalpy change in a reaction is always constant and independent of the manner in which the reaction occurs.

**Answer:** iii and iv

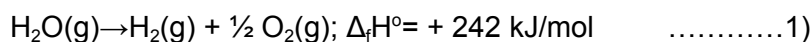
Explanation: Arrhenius equation is the variation of rate constant with temperature. On the other hand, the Kirchhoff equation explained the temperature dependence of enthalpy of reactions.

**Q-3:** From the following data at 25°C

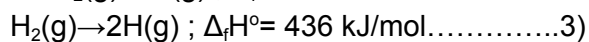
Reaction	$\Delta_f H^\circ$ (kJ/mol)
$\frac{1}{2} \text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{OH}(\text{g})$	42
$\text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g})$	-242
$\text{H}_2(\text{g}) \rightarrow 2\text{H}(\text{g})$	436
$\text{O}_2(\text{g}) \rightarrow 2\text{O}(\text{g})$	495

Calculate the  $\Delta_f H^\circ$  for the reaction  $\text{H}_2\text{O}(\text{g}) \rightarrow 2\text{H}(\text{g}) + \text{O}(\text{g})$

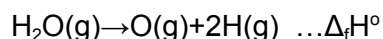
**Answer:** On reversing the equation  $\text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g})$ , we get



Multiply  $\frac{1}{2}$  to the equation  $\text{O}_2(\text{g}) \rightarrow 2\text{O}(\text{g})$ , we get



On adding, 1), 2) and 3), we get



$$\Delta_f H^\circ = (242 + 495/2 + 436) \text{ kJ/mol} = 925.5 \text{ kJ/mol}$$

**Q-4:** The heat of combustion of carbon is 394 kJ/mol. The heat evolved in combustion of  $6.022 \times 10^{22}$  atoms of carbon is:

**Answer:** The reaction for the combustion of carbon is  
 $\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} \quad \Delta H = 394 \text{ kJ/mol}$

From the equation, we can say that  
 Heat of combustion for 1 mol of carbon = 394 kJ/mol  
 1 mol of carbon contains  $6.022 \times 10^{23}$  atoms of carbon

It means,

Heat of combustion for  $6.022 \times 10^{23}$  atoms of carbon = 394 kJ/mol

$$\begin{aligned} \text{Heat of combustion for } 6.022 \times 10^{22} \text{ atoms of carbon} &= (394 \times 6.022 \times 10^{22}) / 6.022 \times 10^{23} \\ &= 39.4 \text{ kJ/mol} \end{aligned}$$

**Q-5:** Calculate standard enthalpies of formation of  $\text{CS}_2\text{(l)}$ . Given the standard enthalpy of combustion of carbon (s), sulphur(s) and  $\text{CS}_2\text{(l)}$  are: -393.3, -293.72 and -1108.76 kJ/mol respectively.

**Answer:**

The combustion reaction for Carbon is  $\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} \quad \Delta H = -393.3 \text{ kJ/mol}$

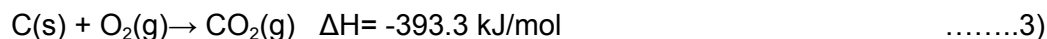
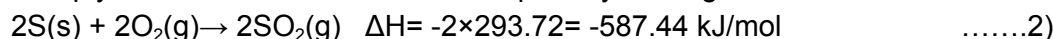
The combustion reaction for sulphur is  $\text{S(s)} + \text{O}_2\text{(g)} \rightarrow \text{SO}_2\text{(g)} \quad \Delta H = -293.72 \text{ kJ/mol}$

The combustion reaction for  $\text{CS}_2\text{(l)}$  is  $\text{CS}_2\text{(l)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{SO}_2\text{(g)} + \text{CO}_2\text{(g)} \quad \Delta H = -1108.76 \text{ kJ/mol}$

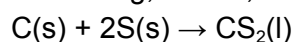
Reverse the reaction for the combustion reaction for  $\text{CS}_2\text{(l)}$ , we get



Multiply the combustion reaction for sulphur by 2, we get



On adding, 1+2+3, we get,



Above equation represents the formation reaction of  $\text{CS}_2\text{(l)}$  from its constituent elements in their standard states. Thus,  $\Delta_f H(\text{CS}_2\text{(l)}) = +1108.76 - 587.44 - 393.3 = 128.02 \text{ kJ/mol}$