



# NCERT Solutions for 11th Class Chemistry: Chapter 6- Thermodynamics



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## NCERT Solutions for 11th Class Chemistry: Chapter 6-Thermodynamics

Class 11: Chemistry Chapter 6 solutions. Complete Class 11 Chemistry Chapter 6 Notes.

### NCERT Solutions for 11th Class Chemistry: Chapter 6-Thermodynamics

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**Question 1. Choose the correct answer:**

**A thermodynamic state function is a quantity**

- (i) used to determine heat changes**
- (ii) whose value is independent of path**
- (iii) used to determine pressure volume work**
- (iv) whose value depends on temperature only.**

**Answer:** (ii) whose value is independent of path

**Question 2. For the process to occur under adiabatic conditions, the correct condition is:**

- (i)  $\Delta T = 0$  (ii)  $\Delta p = 0$**
- (iii)  $q = 0$  (iv)  $w = 0$**

**Answer.** (iii)  $q = 0$

**Question 3. The enthalpies of all elements in their standard states are : ‘**

- (i) unity (ii) zero**
- (iii)  $< 0$  (iv) different for each element**

**Answer:** (ii) zero

**Question 4.**

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$\Delta U^\ominus$  of combustion of methane is  $-X \text{ kJ mol}^{-1}$ . The value of  $\Delta H^\ominus$  is

- (i)  $= \Delta U^\ominus$  (ii)  $> \Delta U^\ominus$   
 (iii)  $< \Delta U^\ominus$  (iv) 0

**Answer:**

The balanced chemical equation for the combustion reaction is :



$$\Delta_{ng} = 1 - 3 = -2$$

$$\Delta H^\ominus = \Delta U^\ominus + \Delta_{ng} RT = \Delta U^\ominus - 2RT$$

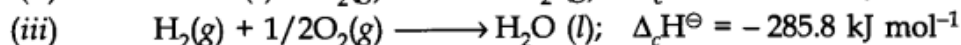
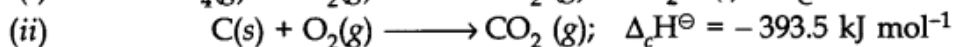
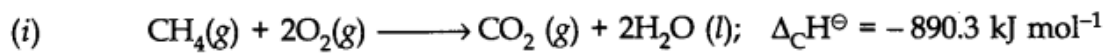
$\therefore \Delta H^\ominus < \Delta U^\ominus$  or (iii) is the correct answer.

**Question 5.** The enthalpy of combustion of methane, graphite and dihydrogen at 298 K are  $-890.3 \text{ kJ mol}^{-1}$ ,  $-393.5 \text{ kJ mol}^{-1}$  and  $-285.8 \text{ kJ mol}^{-1}$  respectively. Enthalpy of formation of  $\text{CH}_4(\text{g})$  will be

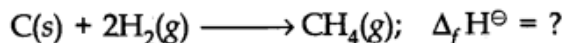
(i)  $-74.8 \text{ kJ mol}^{-1}$  (ii)  $-52.27 \text{ kJ mol}^{-1}$

(iii)  $+74.8 \text{ kJ mol}^{-1}$  (iv)  $+52.26 \text{ kJ mol}^{-1}$

**Answer:** As per the available data :



The equation we aim at:



Eqn. (ii) + 2 × Eqn. (iii) - Eqn. (i) and the correct  $\Delta_f H^\ominus$  value is:

$$= (-393.5) + 2 \times (-285.8) - (-890.3) = -74.8 \text{ kJ mol}^{-1}$$

$\therefore$  (i) is the correct answer.

**Question 6.** A reaction,  $\text{A} + \text{B} \rightarrow \text{C} + \text{D} + q$  is found to have a positive entropy change. The reaction will be

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**(i) possible at high temperature (ii) possible only at low temperature**

**(iii) not possible at any temperature (iv) possible at any temperature**

**Answer:** (iv) possible at any temperature

**Question 7. In a process, 701 J of heat is absorbed by a system and 394 J of work is done by the system. What is the change in internal energy for the process?**

**Answer:** Heat absorbed by the system,  $q = 701 \text{ J}$  Work done by the system =  $-394 \text{ J}$  Change in internal energy ( $\Delta U$ ) =  $q + w = 701 - 394 = 307 \text{ J}$ .

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**Question 8. The reaction of cyanamide,  $\text{NH}_2\text{CN}(\text{s})$  with dioxygen was carried out in a bomb calorimeter and  $\Delta U$  was found to be  $-742.7 \text{ kJ mol}^{-1}$  at  $298 \text{ K}$ . Calculate the enthalpy change for the reaction at  $298 \text{ K}$ .**  
 $\text{NH}_2\text{CN}(\text{s}) + \frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$

**Answer:**  $\Delta U = -742.7 \text{ kJ mol}^{-1}$ ;  $\Delta n_g = 2 - \frac{3}{2} = +\frac{1}{2} \text{ mol}$ .

$R = 8.314 \times 10^{-3} \text{ kJ mol}^{-1} \text{ K}^{-1}$ ;  $T = 298 \text{ K}$

According to the relation,  $\Delta H = \Delta U + \Delta n_g RT$

$\Delta H = (-742.7 \text{ kJ}) + (\frac{1}{2} \text{ mol}) \times (8.314 \times 10^{-3} \text{ kJ mol}^{-1} \text{ K}^{-1}) \times (298 \text{ K})$   
 $= -742.7 \text{ kJ} + 1.239 \text{ kJ} = -741.5 \text{ kJ}$ .

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**Question 9.** Calculate the number of kJ of heat necessary to raise the temperature of 60 g of aluminium from 35°C to 55°C. Molar heat capacity of Al is 24 J mol<sup>-1</sup> K<sup>-1</sup>.

**Answer:** No. of moles of Al (m) = (60g)/(27 g mol<sup>-1</sup>) = 2.22 mol

Molar heat capacity (C) = 24 J mol<sup>-1</sup> K<sup>-1</sup>.

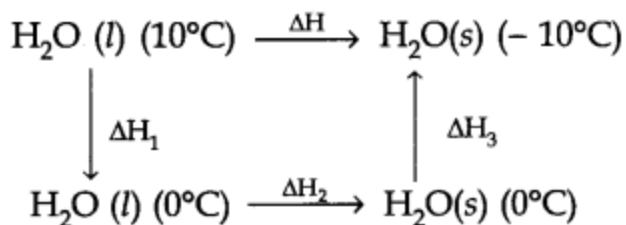
Rise in temperature ( $\Delta T$ ) = 55 – 35 = 20°C = 20 K

Heat evolved (q) = C x m x T = (24 J mol<sup>-1</sup> K<sup>-1</sup>) x (2.22 mol) x (20 K)  
= 1065.6 J = 1.067 kJ

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**Question 10.** Calculate the enthalpy change on freezing of 1.0 mol of water at 10.0°C to ice at – 10.0°C.  $\Delta H_f = 6.03$  KJ mol<sup>-1</sup> at 0°C.  $C_p$  [H<sub>2</sub>O(l)] = 75.3 J mol<sup>-1</sup> K<sup>-1</sup>;  $C_p$  [H<sub>2</sub>O(s)] = 36.8 J mol<sup>-1</sup> K<sup>-1</sup>.

**Answer:** The change may be represented as:



According to Hess's Law;

$$\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$$

$$\Delta H_1 = 75.3 \text{ J mol}^{-1} \text{ K}^{-1} (10 \text{ K}) = 753 \text{ J mol}^{-1}$$

$$\Delta H_2 \text{ (solidification)} = -6.03 \text{ kJ mol}^{-1} = -6030 \text{ J mol}^{-1}$$

(sign changed)

$$\Delta H_3 = 36.8 \text{ J mol}^{-1} \text{ K}^{-1} (-10 \text{ K}) = -368 \text{ J mol}^{-1}$$

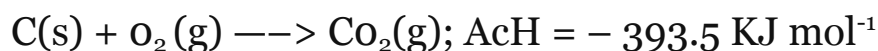
$$\Delta H = (753 - 6030 - 368) \text{ J mol}^{-1} = -5645 \text{ J mol}^{-1}$$

∴

$$= -5.645 \text{ kJ mol}^{-1}.$$

**Question 11. Enthalpy of combustion of carbon to carbon dioxide is  $-393.5 \text{ J mol}^{-1}$ . Calculate the heat released upon formation of  $35.2 \text{ g}$  of  $\text{CO}_2$  from carbon and oxygen gas.**

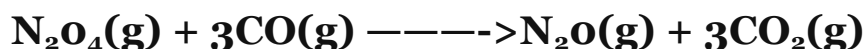
**Answer:** The combustion equation is:



Heat released in the formation of  $44 \text{ g}$  of  $\text{CO}_2 = 393.5 \text{ kJ}$

Heat released in the formation of  $35.2 \text{ g}$  of  $\text{CO}_2 = (393.5 \text{ kJ}) \times (35.2 \text{ g}) / (44 \text{ g}) = 314.8 \text{ kJ}$

**Question 12. Calculate the enthalpy of the reaction:**



Given that;  $\Delta_f H^\circ \text{CO}(g) = -110 \text{ kJ mol}^{-1}$ ;  $\Delta_f H^\circ \text{CO}_2(g) = -393 \text{ kJ mol}^{-1}$

$\Delta_f H^\circ \text{N}_2\text{O}(g) = 81 \text{ kJ mol}^{-1}$ ;  $\Delta_f H^\circ \text{N}_2\text{O}_4(g) = 9.7 \text{ kJ mol}^{-1}$

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**Answer:** Enthalpy of reaction ( $\Delta_r H$ ) =  $[81 + 3(-393)] - [9.7 + 3(-110)]$

$$= [81 - 1179] - [9.7 - 330] = -778 \text{ kJ mol}^{-1}$$

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**Question 13. Given :**  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \longrightarrow 2\text{NH}_3(\text{g}); \Delta_r H^\ominus = -92.4 \text{ kJ mol}^{-1}$  What is the standard enthalpy of formation of  $\text{NH}_3$  gas?

**Answer:**  $\Delta_f H^\ominus \text{NH}_3(\text{g}) = -(92.4)/2 = -46.2 \text{ kJ mol}^{-1}$

**Question 14. Calculate the standard enthalpy of formation of  $\text{CH}_3\text{OH}$ . from the following data:**

(i)  $\text{CH}_3\text{OH}(\text{l}) + 3/2 \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}); \Delta_r H^\ominus = -726 \text{ kJ mol}^{-1}$

(ii)  $\text{C}(\text{s}) + \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g}); \Delta_c H^\ominus = -393 \text{ kJ mol}^{-1}$

(iii)  $\text{H}_2(\text{g}) + 1/2 \text{O}_2(\text{g}) \longrightarrow \text{H}_2\text{O}(\text{l}); \Delta_f H^\ominus = -286 \text{ kJ mol}^{-1}$

**Answer:** The equation we aim at;

$\text{C}(\text{s}) + 2\text{H}_2(\text{g}) + 1/2 \text{O}_2(\text{g}) \longrightarrow \text{CH}_3\text{OH}(\text{l}); \Delta_f H^\ominus = \pm? \dots$  (iv)

Multiply eqn. (iii) by 2 and add to eqn. (ii)

$\text{C}(\text{s}) + 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

$\Delta H = -(393 + 522) = -915 \text{ kJ mol}^{-1}$  Subtract eqn. (iv) from eqn. (i)

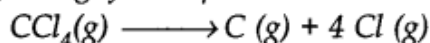
$\text{CH}_3\text{OH}(\text{l}) + 3/2 \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}); \Delta H = -726 \text{ kJ mol}^{-1}$

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Subtract:  $C(s) + 2H_2(g) + 1/2O_2(g) \longrightarrow CH_3OH(l)$ ;  $\Delta_f H^\ominus = -239 \text{ kJ mol}^{-1}$

### Question 15.

Calculate the enthalpy change for the process



and calculate bond enthalpy of C–Cl in  $CCl_4(g)$

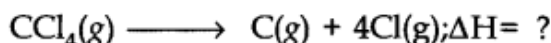
Given:  $\Delta_{vap} H^\ominus (CCl_4) = 30.5 \text{ kJ mol}^{-1}$ ;  $\Delta_f H^\ominus (CCl_4) = -135.5 \text{ kJ mol}^{-1}$   
 $\Delta_a H^\ominus (C) = 715.0 \text{ kJ mol}^{-1}$  where  $\Delta_a H^\ominus$  is enthalpy of atomisation  
 $\Delta_a H^\ominus (Cl_2) = 242 \text{ kJ mol}^{-1}$ .

### Answer:

The available data is:

- (i)  $CCl_4(l) \longrightarrow CCl_4(g)$ ,  $\Delta_{vap} H^\ominus = 30.5 \text{ kJ mol}^{-1}$
- (ii)  $C(s) + 2Cl_2(g) \longrightarrow CCl_4(l)$ ,  $\Delta_f H^\ominus = -135.5 \text{ kJ mol}^{-1}$
- (iii)  $C(s) \longrightarrow C(g)$ ,  $\Delta_a H^\ominus = 715.0 \text{ kJ mol}^{-1}$
- (iv)  $Cl_2(g) \longrightarrow 2 Cl(g)$ ,  $\Delta_a H^\ominus = 242 \text{ kJ mol}^{-1}$

The equation we aim at is:



Eqn. (iii) + 2 × Eqn. (iv) – Eqn. (i) – Eqn. (ii) gives the required equation with

$$\begin{aligned} \Delta H &= 715.0 + 2(242) - 30.5 - (-135.5) \text{ kJ mol}^{-1} \\ &= 1304 \text{ kJ mol}^{-1} \end{aligned}$$

Bond enthalpy of C–Cl in  $CCl_4$  (average value) =  $\frac{1304}{4} = 326 \text{ kJ mol}^{-1}$ .

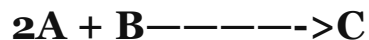
### Question 16. For an isolated system $\Delta U = 0$ ; what will be $\Delta S$ ?

**Answer:** Change in internal energy ( $\Delta U$ ) for an isolated system is zero for it does not exchange any energy with the surroundings. But entropy tends to increase in case of spontaneous reaction. Therefore,  $\Delta S > 0$  or positive.

### Question 17. For a reaction at 298 K

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$$\Delta H = 400 \text{ kJ mol}^{-1} \text{ and } \Delta S = 0.2 \text{ kJ K}^{-1} \text{ mol}^{-1}.$$

**At what temperature will the reaction become spontaneous considering  $\Delta H$  and  $\Delta S$  to be constant over the temperature range?**

**Answer:** As per the Gibbs Helmholtz equation:

$$\Delta G = \Delta H - T\Delta S \text{ For } \Delta G=0 ; \Delta H=T\Delta S \text{ or } T=\Delta H/\Delta S$$

$$T = (400 \text{ KJ mol}^{-1}) / (0.2 \text{ KJ K}^{-1} \text{ mol}^{-1}) = 2000 \text{ K}$$

Thus, reaction will be in a state of equilibrium at 2000 K and will be spontaneous above this temperature.

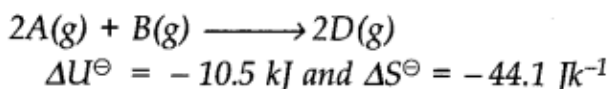
**Question 18. For the reaction;  $2\text{Cl}(\text{g}) \longrightarrow \text{Cl}_2(\text{g})$ ; what will be the signs of  $\Delta H$  and  $\Delta S$ ?**

**Answer:**  $\Delta H$  : negative (- ve) because energy is released in bond formation

$\Delta S$  : negative (- ve) because entropy decreases when atoms combine to form molecules.

**Question 19.**

*For the reaction*



*Calculate  $\Delta G^\ominus$  for the reaction, and predict whether the reaction may occur spontaneously.*

**Answer:**

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$$\Delta H^\ominus = \Delta U^\ominus + \Delta n_g RT$$

$$\Delta U^\ominus = -10.5 \text{ kJ}; \quad \Delta n_g = 2 - 3 = -1 \text{ mol}$$

$$R = 8.314 \times 10^{-3} \text{ kJ K}^{-1} \text{ mol}^{-1}; \quad T = 298 \text{ K}$$

$$\begin{aligned} \therefore \Delta H^\ominus &= (-10.5 \text{ kJ}) + [(-1 \text{ mol}) \times (8.314 \times 10^{-3} \text{ kJ K}^{-1} \text{ mol}^{-1}) \times (298 \text{ K})] \\ &= -10.5 \text{ kJ} - 2.478 \text{ kJ} = -12.978 \text{ kJ} \end{aligned}$$

According to Gibbs Helmholtz equation:

$$\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$$

$$\Delta G^\ominus = (-12.978 \text{ kJ}) - (298 \text{ K}) \times (-0.0441 \text{ kJ K}^{-1})$$

$$= -12.978 + 13.112 = -12.978 + 13.142 = \mathbf{0.164 \text{ kJ}}$$

Since  $\Delta G^\ominus$  is positive, the reaction is non-spontaneous in nature.

### Question 20.

The equilibrium constant for the reaction is 10. Calculate the value of  $\Delta G^\ominus$ ; Given

$$R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}; \quad T = 300 \text{ K}.$$

**Answer:**

$$\Delta G^\ominus = -RT \ln K = -2.303 RT \log K.$$

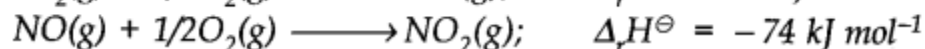
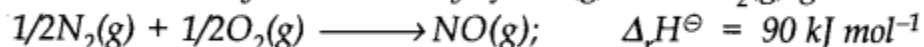
$$R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}; \quad T = 300 \text{ K}; \quad K = 10$$

$$\Delta G^\ominus = -2.303 \times 8.314 \text{ JK}^{-1} \text{ mol}^{-1} \times (300 \text{ K}) \times \log 10$$

$$= -5527 \text{ J mol}^{-1} = \mathbf{-5.527 \text{ kJ mol}^{-1}}.$$

### Question 21.

Comment on the thermodynamic stability of  $\text{NO}(\text{g})$  and  $\text{NO}_2(\text{g})$  given: :



**Answer:**

For  $\text{NO}(\text{g})$ ;  $\Delta_r H^\ominus = +\text{ve}$ : Unstable in nature

For  $\text{NO}_2(\text{g})$ ;  $\Delta_r H^\ominus = -\text{ve}$ : Stable in nature

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## Question 22.

Calculate the entropy change in surroundings when 1.0 mol of  $\text{H}_2\text{O}(l)$  is formed under standard conditions. Given  $\Delta H^\ominus = -286 \text{ kJ mol}^{-1}$ .

**Answer:**

$$q_{\text{rev}} = (-\Delta_f H^\ominus) = -286 \text{ kJ mol}^{-1} = 286000 \text{ J mol}^{-1}$$

$$\Delta S_{(\text{Surroundings})} = \frac{q_{\text{rev}}}{T} = \frac{(286000 \text{ J mol}^{-1})}{298 \text{ K}} = 959 \text{ J K}^{-1} \text{ mol}^{-1}.$$

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- Chapter 3-Classification of Elements and Periodicity in Properties
- Chapter 4-Chemical Bonding and Molecular Structure
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