

# NCERT SOLUTIONS CLASS-XI CHEMISTRY CHAPTER-6 CHEMICAL THERMODYNAMICS

Q-1: A thermodynamic state function is \_\_\_\_\_.

1. A quantity which depends upon temperature only.
2. A quantity which determines pressure-volume work.
3. A quantity which is independent of path.
4. A quantity which determines heat changes.

Ans:

(3) A quantity which is independent of path

Reason:

Functions like pressure, volume and temperature depends on the state of the system only and not on the path.

Q-2: Which of the following is a correct conditions for adiabatic condition to occur.

1.  $q = 0$
2.  $w = 0$
3.  $\Delta p = 0$
4.  $\Delta T = 0$

Ans:

1:  $q = 0$

Reason:

For an adiabatic process heat transfer is zero, i.e.  $q = 0$ .

Q-3: The value of enthalpy for all elements in standard state is \_\_\_\_\_.

- (1) Zero
- (2)  $< 0$
- (3) Different for every element
- (4) Unity

Ans:

(1)  $C(s)$ ,  $O_2(g)$ ,  $Cl_2(g)$ ,  $H_2(g)$ ,  $N_2(g)$ ,  $Br_2(l)$ ,  $I_2(s)$

(4)  $C(g) + 4Cl(g) \rightarrow CCl_4(g)$ ;  $\Delta_f H^\ominus = -135.5 \text{ kJmol}^{-1}$

$\Delta H$  for the process  $CCl_4(g) \rightarrow C(g) + 4Cl(g)$  can be measured as:

$$\begin{aligned} \Delta H &= \Delta_a H^\ominus(C) + 2\Delta_a H^\ominus(Cl_2) - \Delta_{vap} H^\ominus - \Delta_f H \\ &= (715 \text{ kJmol}^{-1}) + 2(\text{kJmol}^{-1}) - (30.5 \text{ kJmol}^{-1}) - (-135.5 \text{ kJmol}^{-1}) \end{aligned}$$

Therefore,  $H = 1304 \text{ kJmol}^{-1}$

The value of bond enthalpy for C-Cl in  $CCl_4(g)$

$$\begin{aligned} &= \frac{1304}{4} \text{ kJmol}^{-1} \\ &= 326 \text{ kJmol}^{-1} \end{aligned}$$

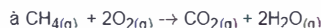
Q-4:  $\Delta U = 0$  for isolated system, then what will be  $\Delta H$ ?

Q-5: For, Methane, di-hydrogen and and graphite the enthalpy of combustion at 298K are given  $-890.3 \text{ kJmol}^{-1}$ ,  $-285.8 \text{ kJmol}^{-1}$  and  $-393.5 \text{ kJmol}^{-1}$  respectively. Find the enthalpy of formation of Methane gas?

- (a)  $-52.27 \text{ kJmol}^{-1}$
- (b)  $52 \text{ kJmol}^{-1}$
- (c)  $+74.8 \text{ kJmol}^{-1}$
- (d)  $-74.8 \text{ kJmol}^{-1}$

Ans:

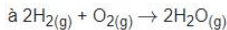
(d)  $-74.8 \text{ kJmol}^{-1}$



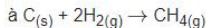
$$\Delta H = -890.3 \text{ kJ mol}^{-1}$$



$$\Delta H = -393.5 \text{ kJ mol}^{-1}$$



$$\Delta H = -285.8 \text{ kJ mol}^{-1}$$



$$\Delta_f H_{\text{CH}_4} = \Delta_c H_c + 2\Delta_f H_{\text{H}_2} - \Delta_f H_{\text{CO}_2}$$

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

$$= -74.8 \text{ kJ mol}^{-1}$$

Q-6: A reaction,  $X + Y \rightarrow U + V + q$  is having a +ve entropy change. Then the reaction \_\_\_\_.

(a) will be possible at low temperature only

(b) will be possible at high temperature only

(c) will be possible at any temperature

(d) won't be possible at any temperature

Ans:

(c) will be possible at any temperature

$\Delta G$  should be -ve, for spontaneous reaction to occur

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

As per given in question,

$\Delta H$  is -ve ( as heat is evolved)

$\Delta S$  is +ve

Therefore,  $\Delta G$  is negative

So, the reaction will be possible at any temperature.

Q-7: In the process, system absorbs 801 J and work done by the system is 594 J. Find  $\Delta U$  for the given process.

Ans:

As per Thermodynamics 1<sup>st</sup> law,

$$\Delta U = q + W(i);$$

$\Delta U$  internal energy = heat

W = work done

W = -594 J (work done by system)

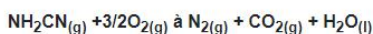
q = +801 J (+ve as heat is absorbed)

Now,

$$\Delta U = 801 + (-594)$$

$$\Delta U = 207 \text{ J}$$

Q-8: The reaction given below was done in bomb calorimeter, and at 298K we get,  $\Delta U = -753.7 \text{ kJ mol}^{-1}$ . Find  $\Delta H$  at 298K.



Ans:

$\Delta H$  is given by,

$$\Delta H = \Delta U + \Delta n_g RT \dots \dots \dots (1)$$

$\Delta n_g$  = change in number of moles

$\Delta U$  = change in internal energy

Here,

$$\Delta n_g = \sum n_g(\text{product}) - \sum n_g(\text{reactant})$$

$$= (2 - 2.5) \text{ moles}$$

$$\Delta n_g = -0.5 \text{ moles}$$

Here,

$$T = 298\text{K}$$

$$\Delta U = -753.7 \text{ kJmol}^{-1}$$

$$R = 8.314 \times 10^{-3} \text{ kJmol}^{-1}\text{K}^{-1}$$

Now, from (1)

$$\Delta H = (-753.7 \text{ kJmol}^{-1}) + (-0.5 \text{ mol})(298\text{K})(8.314 \times 10^{-3} \text{ kJmol}^{-1}\text{K}^{-1})$$

$$= -753.7 - 1.2$$

$$\Delta H = -754.9 \text{ kJmol}^{-1}$$

**Q-9: Calculate the heat (in kJ) required for 50.0 g aluminium to raise the temperature from 45°C to 65°C. For aluminium molar heat capacity is 24 Jmol<sup>-1</sup>K<sup>-1</sup>**

**Ans:**

Expression of heat(q),

$$q = mCP\Delta T \dots \dots \dots (a)$$

$\Delta T$  = Change in temperature

c = molar heat capacity

m = mass of substance

From (a)

$$q = \left(\frac{50}{27} \text{ mol}\right)(24 \text{ mol}^{-1}\text{K}^{-1})(20\text{K})$$

$$q = 888.88 \text{ J}$$

**Q-10: Calculate  $\Delta H$  for transformation of 1 mole of water into ice from 10°C to (-10)°C.**

$$\Delta_{fus}H = 6.03 \text{ kJmol}^{-1} \text{ at } 10^\circ\text{C}.$$

$$C_p[H_2O_{(l)}] = 75.3 \text{ J mol}^{-1}\text{K}^{-1}$$

$$C_p[H_2O_{(s)}] = 36.8 \text{ J mol}^{-1}\text{K}^{-1}$$

**Ans:**

$\Delta H_{total}$  = sum of the changes given below:

(a) Energy change that occurs during transformation of 1 mole of water from 10°C to 0°C.

(b) Energy change that occurs during transformation of 1 mole of water at 0°C to 1 mole of ice at 0°C.

(c) Energy change that occurs during transformation of 1 mole of ice from 0°C to (-10)°C.

$$\Delta H_{total} = C_p[H_2O_{(l)}]\Delta T + \Delta H_{freezing}C_p[H_2O_{(s)}]\Delta T$$

$$= (75.3 \text{ Jmol}^{-1}\text{K}^{-1})(0 - 10)\text{K} + (-6.03 \times 1000 \text{ Jmol}^{-1})(-10)\text{K}$$

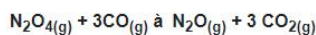
$$= -753 \text{ Jmol}^{-1} - 6030 \text{ Jmol}^{-1} - 368 \text{ Jmol}^{-1}$$

$$= -7151 \text{ Jmol}^{-1}$$

$$= -7.151 \text{ kJmol}^{-1}$$

Thus, the required change in enthalpy for given transformation is -7.151 kJmol<sup>-1</sup>.

**Q-11: Enthalpies of formation for CO<sub>2</sub>(g), CO(g), N<sub>2</sub>O<sub>4</sub>(g), N<sub>2</sub>O(g) are -393 kJmol<sup>-1</sup>, -110 kJmol<sup>-1</sup>, 9.7 kJmol<sup>-1</sup> and 81 kJmol<sup>-1</sup> respectively. Then,  $\Delta_r H =$  \_\_\_\_\_.**

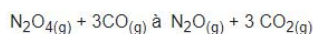


**Ans:**

" $\Delta_r H$  for any reaction is defined as the difference between  $\Delta_f H$  value of products and  $\Delta_f H$  value of reactants."

$$\Delta_r H = \sum \Delta_f H(\text{products}) - \sum \Delta_f H(\text{reactants})$$

Now, for



$$\Delta_r H = [(\Delta_f H(N_2O) + (3\Delta_f H(CO_2)) - (\Delta_f H(N_2O_4) + 3\Delta_f H(CO)))]$$

Now, substituting the given values in the above equation, we get:

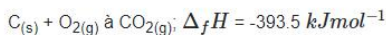
$$\Delta_r H = \{[81 \text{ kJmol}^{-1} + 3(-393 \text{ kJmol}^{-1})] - [9.7 \text{ kJmol}^{-1} + 3(-110 \text{ kJmol}^{-1})]\}$$

$$\Delta_r H = -777.7 \text{ kJmol}^{-1}$$

**Q-12 Enthalpy of combustion of C to CO<sub>2</sub> is -393.5 kJmol<sup>-1</sup>. Determine the heat released on the formation of 37.2g of CO<sub>2</sub> from dioxygen and carbon.**

**Ans:**

Formation of carbon dioxide from di-oxygen and carbon gas is given as:



1 mole CO<sub>2</sub> = 44g

Heat released during formation of 44g CO<sub>2</sub> = -393.5 kJmol<sup>-1</sup>

Therefore, heat released during formation of 37.2g of CO<sub>2</sub> can be calculated as

$$= \frac{-393.5 \text{ kJmol}^{-1}}{44 \text{ g}} \times 37.2 \text{ g}$$

$$= -332.69 \text{ kJmol}^{-1}$$

**Q-13: N<sub>2(g)</sub> + 3H<sub>2(g)</sub> → 2NH<sub>3(g)</sub>; Δ<sub>r</sub>H<sup>Θ</sup> = -92.4 kJmol<sup>-1</sup>**

**Standard Enthalpy for formation of ammonia gas is \_\_\_\_\_.**

**Ans:**

"Standard enthalpy of formation of a **compound** is the enthalpy that takes place during the formation of 1 mole of a substance in its standard form, from its constituent elements in their standard state."

Dividing the chemical equation given in the question by 2, we get



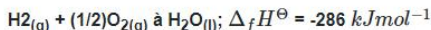
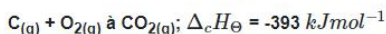
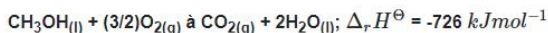
Therefore, Standard Enthalpy for formation of ammonia gas

$$= (0.5) \Delta_r H^\ominus$$

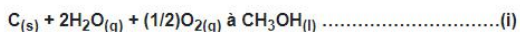
$$= (0.5)(-92.4 \text{ kJmol}^{-1})$$

$$= -46.2 \text{ kJmol}^{-1}$$

**Q-14: Determine Standard Enthalpy of formation for CH<sub>3</sub>OH(l) from the data given below:**



**Ans:**



CH<sub>3</sub>OH(l) can be obtained as follows,

$$\Delta_f H_\ominus [CH_3OH(l)] = \Delta_c H_\ominus$$

$$2\Delta_f H_\ominus - \Delta_r H_\ominus$$

$$= (-393 \text{ kJmol}^{-1}) + 2(-286 \text{ kJmol}^{-1}) - (-726 \text{ kJmol}^{-1})$$

$$= (-393 - 572 + 726) \text{ kJmol}^{-1}$$

$$= -239 \text{ kJmol}^{-1}$$

$$\text{Thus, } \Delta_f H_\ominus [CH_3OH(l)] = -239 \text{ kJmol}^{-1}$$

**Q-15: Calculate ΔH for the following process**

**CCl<sub>4(g)</sub> → C<sub>(s)</sub> + 4Cl<sub>2(g)</sub> and determine the value of bond enthalpy for C-Cl in CCl<sub>4(g)</sub>.**

$$\Delta_{\text{vap}}H^\ominus(\text{CCl}_4) = 30.5 \text{ kJmol}^{-1}$$

$$\Delta_fH^\ominus(\text{CCl}_4) = -135.5 \text{ kJmol}^{-1}$$

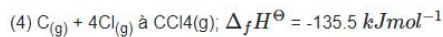
$$\Delta_aH^\ominus(\text{C}) = 715 \text{ kJmol}^{-1}$$

$\Delta_aH^\ominus$  is a enthalpy of atomisation

$$\Delta_aH^\ominus(\text{Cl}_2) = 242 \text{ kJmol}^{-1}$$

**Ans:**

"The chemical equations implying to the given values of enthalpies" are:



$\Delta H$  for the process  $\text{CCl}_4(l) \rightarrow \text{C}(g) + 4\text{Cl}(g)$  can be measured as:

$$\Delta H = \Delta_aH^\ominus(\text{C}) + 2\Delta_aH^\ominus(\text{Cl}_2) - \Delta_{\text{vap}}H^\ominus - \Delta_fH$$

$$= (715 \text{ kJmol}^{-1}) + 2(\text{kJmol}^{-1}) - (30.5 \text{ kJmol}^{-1}) - (-135.5 \text{ kJmol}^{-1})$$

$$\text{Therefore, } H = 1304 \text{ kJmol}^{-1}$$

**The value of bond enthalpy for C-Cl in  $\text{CCl}_4(g)$**

$$= \frac{1304}{4} \text{ kJmol}^{-1}$$

$$= 326 \text{ kJmol}^{-1}$$

**Q-16:**  $\Delta U = 0$  for isolated system, then what will be  $\Delta U$ ?

**Ans:**

$\Delta U$  is positive ;  $\Delta U > 0$ .

As,  $\Delta U = 0$  then  $\Delta S$  will be +ve, as a result reaction will be spontaneous.

**Q-17:**

**Following reaction takes place at 298K,**



$$\Delta H = 400 \text{ kJmol}^{-1}$$

$$\Delta H = 0.2 \text{ kJmol}^{-1} \text{K}^{-1}$$

**Find the temperature at which the reaction become spontaneous considering  $\Delta S$  and  $\Delta H$  to be constant over the entire temperature range?**

**Ans:**

Now,

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

Let, the given reaction is at equilibrium, then  $\Delta T$  will be:

$$T = \frac{(\Delta H - \Delta G)}{\Delta S}$$

$$\frac{\Delta H}{\Delta S}; (\Delta G = 0 \text{ at equilibrium})$$

$$= 400 \text{ kJmol}^{-1} / 0.2 \text{ kJmol}^{-1} \text{K}^{-1}$$

$$\text{Therefore, } T = 2000 \text{K}$$

Thus, for the spontaneous,  $\Delta G$  must be -ve and  $T > 2000 \text{K}$ .

**Q-18:**  $2\text{Cl}(g) \rightarrow \text{Cl}_2(g)$

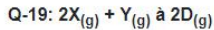
**In above reaction what can be the sign for  $\Delta S$  and  $\Delta H$ ?**

**Ans:**

$\Delta S$  and  $\Delta H$  are having negative sign.

The reaction given in the question represents the formation of Cl molecule from Cl atoms. As the formation of bond takes place in the given reaction. So, energy is released. So,  $\Delta H$  is negative.

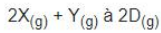
Also, 2 moles of Chlorine atoms is having more randomness than 1 mole of chlorine molecule. So, the spontaneity is decreased. Thus,  $\Delta S$  is negative.



$$\Delta U^\ominus = -10.5 \text{ kJ and } \Delta S^\ominus = -44.1 \text{ JK}^{-1}$$

**Determine  $\Delta G^\ominus$  for the given reaction, and predict that whether given reaction can occur spontaneously or not.**

**Ans:**



$$\Delta n_g = 2 - 3$$

$$= -1 \text{ mole}$$

Putting value of  $\Delta U^\ominus$  in expression of  $\Delta H$ :

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

$$= (-10.5 \text{ kJ}) - (-1)(8.314 \times 10^{-3} \text{ kJK}^{-1} \text{ mol}^{-1})(298 \text{ K})$$

$$= -10.5 \text{ kJ} - 2.48 \text{ kJ}$$

$$\Delta H^\ominus = -12.98 \text{ kJ}$$

Putting value of  $\Delta S^\ominus$  and  $\Delta H^\ominus$  in expression of  $\Delta G^\ominus$ :

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

$$= -12.98 \text{ kJ} - (298 \text{ K})(-44.1 \text{ JK}^{-1})$$

$$= -12.98 \text{ kJ} + 13.14 \text{ kJ}$$

$$\Delta G^\ominus = 0.16 \text{ kJ}$$

**As,  $\Delta G^\ominus$  is positive, the reaction won't occur spontaneously.**

**Q-20:** Find the value of  $\Delta G^\ominus$  for the reaction, if equilibrium is given 10. given that  $T = 300 \text{ K}$  and  $R = 8.314 \times 10^{-3} \text{ kJK}^{-1} \text{ mol}^{-1}$ .

**Ans:**

Now,

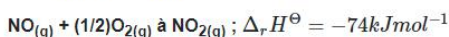
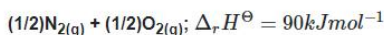
$$\Delta G^\ominus = -2.303RT \log eq$$

$$= (2.303)(8.314 \times 10^{-3} \text{ kJK}^{-1} \text{ mol}^{-1})(300 \text{ K}) \log 10$$

$$= -5744.14 \text{ Jmol}^{-1}$$

$$= -5.744 \text{ kJmol}^{-1}$$

**Q-21:** What can be said about the thermodynamic stability of  $\text{NO}_{(g)}$ , given



**Ans:**

The +ve value of  $\Delta_r H$  represents that during  $\text{NO}_{(g)}$  formation from  $\text{O}_2$  and  $\text{N}_2$ , heat is absorbed. The obtained product,  $\text{NO}_{(g)}$  is having more energy than reactants. Thus,  $\text{NO}_{(g)}$  is unstable.

The -ve value of  $\Delta_r H$  represents that during  $\text{NO}_{2(g)}$  formation from  $\text{O}_{2(g)}$  and  $\text{NO}_{(g)}$ , heat is evolved. The obtained product,  $\text{NO}_{2(g)}$  gets stabilized with minimum energy.

**Thus, unstable  $\text{NO}_{(g)}$  converts into unstable  $\text{NO}_{2(g)}$ .**

**Q-22:** Determine  $\Delta S$  in surrounding given that mole of  $\text{H}_2\text{O}(l)$  is formed I standard condition.

$$\Delta_r H^\ominus = -286 \text{ kJmol}^{-1}$$

**Ans:**

$\Delta_r H^\ominus = -286 \text{ kJ mol}^{-1}$  is given so that amount of heat is evolved during the formation of 1 mole of  $\text{H}_2\text{O}_{(l)}$ .

Thus, the same heat will be absorbed by surrounding.  $Q_{\text{surr}} = +286 \text{ kJ mol}^{-1}$ .

Now,  $\Delta S_{\text{surr}} = Q_{\text{surr}}/T$

$$= \frac{286 \text{ kJ mol}^{-1}}{298 \text{ K}}$$

Therefore,  $\Delta S_{\text{surr}} = 959.73 \text{ J mol}^{-1} \text{ K}^{-1}$

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