NCERT solutions for class 11 chemistry chapter 6 Thermodynamics

Question 6.1 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 :

A state function refers to the equilibrium state of a system. A thermodynamic state function is a quantity whose value is independent of path.

Like- p, V, T depends on the state of the system, not on the path.

So, the correct option is (ii)

Question 6.2 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 :

In adiabatic condition, the heat exchange between the system and the surrounding through its boundary is not allowed.

So, the correct option is (iii) q = 0

Question 6.3 The enthalpies of all elements in their standard states are:

(i) unity

(ii) zero

(iii) < 0

(iv) different for each element

Answer :

The enthalpies of all elements in their standard states are Zero.

So, the correct option is (ii)

Question 6.4 $\Delta U^{\ominus}$ of combustion of methane is $-X\,KJmol^{-1}$. The value of $\Delta H^{\ominus}$ is

(i) $= \Delta U^{\ominus}$

(ii) $> \Delta U^{\ominus}$

(iii) $< \Delta U^{\ominus}$

(iv) = 0

Answer :

Equation of combustion of methane-

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

Since, $\Delta H^o = \Delta U^o + \Delta n_g RT$ and $\Delta U^o = -X\,KJ\,mol^{-1}$

from the equation $\Delta n_g = (n_p - n_R) = 1 - 3 = -2$

$$\Delta H^o = -X - 2RT$$

$$\Delta H^o < \Delta U^o$$

so the correct option is (iii).

Question 6.5 The enthalpy of combustion of methane, graphite and dihydrogen at 298 K are, $-890.KJmol^{-1}, -393.5KJmol^{-1}$ and $-285.8KJmol^{-1}$ respectively. Enthalpy of formation of $CH_4(g)$ will be
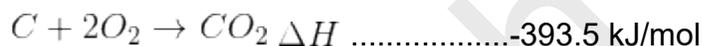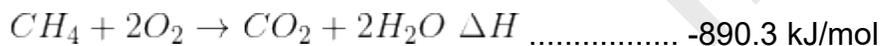
(i) –74.8 $kJmol^{-1}$

(ii) –52.27 $kJmol^{-1}$

(iii) +74.8 $kJmol^{-1}$

(iv) +52.26 $kJmol^{-1}$

Answer :

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O \ \Delta H$ .... -890.3 kJ/mol

$C + 2O_2 \rightarrow CO_2 \ \Delta H$ ....-393.5 kJ/mol

$2H_2 + O_2 \rightarrow 2H_2O \ \Delta H$ ....-258.8 kJ/mol

So, the required equation is to get the formation of $CH_4(g)$ by combining these three equations-

Thus, $C + 2H_2 \rightarrow CH_4$

$\Delta_f H_{CH_4} = \Delta_c H_c + 2\Delta_c H_{H_2}$

= [-393.5 + 2(-285.8) + 890.3]

= -74.8 kJ/mol

therefore, the enthalpy of formation of methane is -74.8 kJ/mol. So, the correct option is (i)

Question 6.6 A reaction, A + B $\rightarrow$ C + D + q is found to have a positive entropy

change. The reaction will be

(i) possible at high temperature

(ii) possible only at low temperature

(iii) not possible at any temperature

(v) possible at any temperature

Answer :

For the reaction to be feasible the ΔG should be negative

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

According to question,

ΔS = positive unit

ΔH = negative (as heat is evolved in the reaction)

Overall the ΔG is negative.

Therefore the reaction is possible at any temperature. So, the correct option is (iv)

Question 6.7 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 :

The first law of Thermodynamics state that,

$$\Delta U = q + W$$

where, ΔU = change in internal energy for the process

q = heat and W = work

Given that,

q = +701 J (heat is absorbed)

W = - 394 (work is done by the system)

Substituting the value in equation of first law we get,

$$\Delta U = 701 + (-394) = 307 J$$

So, the change in internal energy for the process is 307 J

Question 6.8 The reaction of cyanamide, NH_2CN (s), with dioxygen was carried out in a bomb calorimeter, and ΔU was found to be $-742.7 KJ mol^{-1}$ at 298 K. Calculate enthalpy change for the reaction at 298 K.

$$NH_2CN(g) + \frac{3}{2}O_2(g) \rightarrow N_2 + CO_2(g)H_2O(l)$$
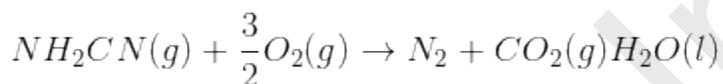
Answer :

Given information,

$$\Delta U = -742.7 kJ mol^{-1}$$

T = 298 K

R = 8.314 $\times 10^{-3}$

$\Delta n_g = n_g$ (products) $-n_g$ (reactants)

= (2 - 1.5)

= 0.5 moles

The enthalpy change for the reaction is expressed as;

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

where, ΔU = change in internal energy and

Δn_g = change in no. of moles

By putting the values we get,

$$\Delta H = (-742.7) + 0.5(298)(8.314 \times 10^{-3})$$

$$= -741.5 kJmol^{-1}$$

Question 6.9 Calculate the number of kJ of heat necessary to raise the temperature of 60.0 g of aluminium from $35^\circ C$ to $55^\circ C$. Molar heat capacity of Al is $24 Jmol^{-1}K^{-1}$

Answer :

We know that

$$q = mc\Delta T$$

m = mass of the substance

c = heat capacity

ΔT = change in temperature

By putting all these values we get,

$$q = (\frac{60}{27}mol)(24J/mol/K)(20K)$$

= 1066.7 J

= 1.07 kJ

Question 6.10 Calculate the enthalpy change on freezing of 1.0 mol of water at $10.0^\circ C$ to ice at $-10.0^\circ C$.

$\Delta_{fus} H = 6.03 KJmol^{-1}$ at $0^\circ C$.

$$Cp[H_2O(l)] = 75.3 Jmol^{-1}K^{-1}$$

$$Cp[H_2O(s)] = 36.8 Jmol^{-1}K^{-1}$$

Answer :

Total enthalpy change is equal to the summation of all the energy required at three different stages-

(i) from 10^oC to 0^oC of 1 mol of water (water $\to$ water)

(ii) from 0^oC to 0^oC of 1 mol of ice (water $\to$ ice)

(iii) from 0^oC to -10^oC of 1 mole of ice (ice $\to$ ice)

So, the total enthalpy change

$$\Delta H = C_p[H_2O(l)](\Delta T) + \Delta H_{(freezing)} + C_p[H_2O(s)]\Delta T$$

$$= (75.3)(-10) + (-6.03 \times 10^3) + (36.8)(-10)$$
$$= -753 - 6030 - 368$$
$$= -7151 Jmol^{-1}$$
$$= 7.151 kJ/mol$$

Hence the total enthalpy change in the transformation process is 7.151 kJ/mol

Question 6.11 Enthalpy of combustion of carbon to CO_2 is $-393.5 KJmol^{-1}$ Calculate the heat released upon formation of 35.2 g of CO_2 from carbon and dioxygen gas.

Answer :

Formation of carbon dioxide from carbon and dioxygen reaction is-

$C + O_2 \to CO_2$ .... ΔH_f = -393.5 kJ/mol

1 mole of CO_2 = 44 g

So, the heat released in the formation of 44g of CO_2 = -393.5 kJ/mol

therefore, In 35.5 g of CO_2 the amount of **heat released** =
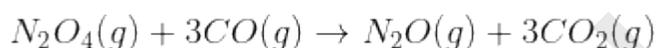$$\frac{-393.5}{44} \times 35.2 = -314.8 kJ/mol$$

Question 6.12 Enthalpies of formation of $CO(g), CO_2(g), N_2O(g)$ and $N_2O_4(g)$ are –110, – 393, 81 and 9.7 $K Jmol^{-1}$ respectively. Find the value of $\Delta_r H$ for the reaction:

$$N_2O_4(g) + 3CO(g) \rightarrow N_2O(g) + 3CO_2(g)$$
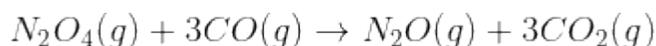
Answer :

Given,

Enthalpies of formation of $CO(g), CO_2(g), N_2O(g)$ and $N_2O_4(g)$ are –110, – 393, 81 and 9.7 $K Jmol^{-1}$

$$N_2O_4(g) + 3CO(g) \rightarrow N_2O(g) + 3CO_2(g)$$

We know that the $\Delta_r H = \sum \Delta_f H$ (product) $- \sum \Delta_f H$ (reactants)

For the above reaction,

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

Substituting the given values we get,
$$= [81 + 3(-393)] - [9.7 + 3(-110)]$$
$$= -777.7 kJ/mol$$

Thus the value of $\Delta_r H$ of the reaction is -777.7 kJ/mol

Question 6.13 Given $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g); \Delta r H = -92.4 kJmol^{-1}$ .What is the standard enthalpy of formation of NH_3 gas?

Answer :

the standard enthalpy of formation of any compound is the required change in enthalpy of formation of 1 mole of a substance in its standard form from its constituent elements.

Thus we can re-write the reaction as;

$$\frac{1}{2}N_2 + \frac{3}{2}H_2 \rightarrow NH_3$$
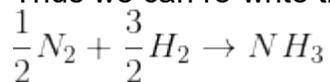
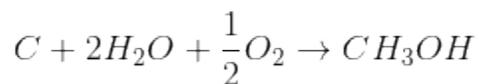Therefore, standard enthalpy of formation of ammonia is = $\frac{1}{2}\Delta_r H^\ominus$

= 1/2 (-92.4)

= -46.2 kJ/mol

Question 6.14 Calculate the standard enthalpy of formation of CH3OH(l) from the following data:

Answer :

for the formation of CH_3OH the reaction is ,

$$C + 2H_2O + \frac{1}{2}O_2 \rightarrow CH_3OH$$

This can be obtained by the following expressions-

required eq = eq (i) + 2 (eq iii) - eq (i)

$$\Delta_f H[CH_3OH] = \Delta_c H^\ominus + 2\Delta_f H[H_2O] - \Delta_r H^\ominus$$
$$= (-393) + 2(-286) - (-726)$$
$$= -239 kJmol^{-1}$$

Question 6.15 Calculate the enthalpy change for the process $CCl_4(g) \rightarrow C(g) + 4Cl(g)$ and calculate bond enthalpy of $C - Cl$ in $CCl_4(g)$

Answer :

We have the following chemical reactions equations-

(1) $CCl_4(l) \rightarrow CCl_4(g)$ .... $\Delta_{vap}H^o = 30.5 kJ/mol$

(2) $C(s) \rightarrow C(g)$ .... $\Delta_a H^o = 715.0 kJ/mol$

(3) $Cl2(g) \rightarrow 2Cl(g)$ .... $\Delta_a H^o = 242 kJ/mol$

(4) $C(g) + 4Cl \rightarrow CCl_4(g)$ .... $\Delta_f H = -135.5 kJ/mol$

The enthalpy change for the process $CCl_4(g) \rightarrow C(g) + 4Cl(g)$ by the above

reaction is calculated as;

$$\Delta H = \Delta_a H^o(C) + 2(\Delta_a H^o(Cl_2)) - \Delta_{vap} H^o - \Delta_f H$$
$$= [(715) + 2(242) - (30.5) - (-135.5)] kJ/mol$$
$$= 1304 kJmol^{-1}$$

And the bond enthalpy of C-Cl bond in CCl_4 (g) = 1304/4 = 326 kJ/mol

Question 6.16 For an isolated system, $\Delta U = 0$, what will be ΔS ?

Answer :

Since $\Delta U = 0$

So, the $\Delta S > 0$ (positive) therefore the reaction will be feasible.

Question 6.17 For the reaction at 298 K,

2A + B → C

$\Delta H = 400 kJmol^{-1}$ and $\Delta S = 0.2 kJK^{-1}mol^{-1}$

At what temperature will the reaction become spontaneous considering ΔH and ΔS to

be constant over the temperature range.

Answer :

From the equation,

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

Suppose the reaction is at equilibrium, So the change in temperature is given as;

$$T = \frac{\Delta H - \Delta G}{\Delta S}$$
$$= \frac{400}{0.2} = 2000K$$
$\qquad$ (ΔG at equilibrium is zero)

To reaction should be spontaneous, ΔG should be neagtive. So, that for the given reaction T should be greater than 2000 K
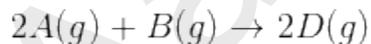
Question 6.18 For the reaction, $2Cl(g) \rightarrow Cl_2(g)$, what are the signs of $\Delta H \ and \ \Delta S$?

Answer :

The given reaction represents the formation of a chlorine molecule from its atom. Bond formation taking place, therefore the energy is released during this. So, ΔH is negative.

Two moles of an atom have more randomness than the one mole of chlorine. So, spontaneity is decreased. Hence ΔS is negative.

Question 6.19 For the reaction
$$2A(g) + B(g) \rightarrow 2D(g)$$
$$\Delta U^{\ominus} = -10.5KJ \ and \ \Delta S^{\ominus} = -44.1JK^{-1}$$
Calculate $\Delta G^{\ominus}$ for the reaction, and predict whether the reaction may occur spontaneously.

Answer :

Given reaction is $2A(g) + B(g) \rightarrow 2D(g)$

We know that, $\Delta G^o = \Delta H^o - T\Delta S^o$

and $\Delta H^o = \Delta U^o + \Delta n_g RT$

here $\Delta n_g = 2 - (3) = -1$

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

substituting the given values in equations-

$\Delta G = (-10.5 + 8.314 \times 10^{-3} \times 298) - 298(-44.1)$
$\Delta G = (-12.98 + 13.14)kJ$
$\Delta G = +0.16kJ$

Hence the reaction will not occur spontaneously because the $\Delta G^\ominus$ value is positive for this reaction.

Question 6.20 The equilibrium constant for a reaction is 10. What will be the value of

$\Delta G^\ominus ? R = 8.314 JK^{-1}mol^{-1}, T = 300K$

Answer :

Given values are,

$R = 8.314 JK^{-1}mol^{-1}, T = 300K$

and equilibrium constant = 10

It is known that,

$\Delta G^o = -2.303\ RT \log K_{eq}$
$= -2.303(8.314)(300) \log 10$
$= -5744.14 J mol^{-1}$
$= -5.744 kJ mol^{-1}$

Hence the value of ΔG^o is -5.744 kJ/mol

Question 6.21 Comment on the thermodynamic stability of NO(g), given

Answer :

The formation of NO is unstable because $\Delta_r H$ is positive it means heat is absorbed during the reaction. So, NO (g) has higher energy than its reactants N_2 and O_2. On the other hand, NO_2 is stable because $\Delta_r H$ is negative means heat is released during the formation of NO_2. It is stablised with minimum energy. Hence unstable NO changes to stable NO_2.

Question 6.22 Calculate the entropy change in surroundings when 1.00 mol of $H_2O(l)$ is formed under standard conditions. $\Delta_f H^\ominus = -286 KJ mol^{-1}$

Answer :

-286 kJ/mol heat is evolved when 1 mol of the water molecule is formed. It means the same amount of energy is absorbed by the surrounding also.

$q_{surr} = +286$ kJ/mol

We know that,

Entropy changes for the surrounding is = q(surr)/ temp.

= 286000/ 298

= 959.73 J/mol/K