**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**](http://cbse.meritnation.com/study-online/solution/chemistry/iGfWXR7ntxWaX5rU6a76uw%21%21#optionContent1)

A thermodynamic state function is a quantity whose value is independent of a path.

Functions like *p*, *V*, *T* etc. depend only on the state of a system and not on the path.

Hence, alternative (ii) is correct.

**Question 6.2:**

For the process to occur under adiabatic conditions, the correct condition is:

(i) Δ*T*= 0

(ii) Δ*p*= 0

(iii) *q*= 0

(iv) *w* = 0

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/647%24CpFqjb5QohjZB8%24srw%21%21#optionContent1)

A system is said to be under adiabatic conditions if there is no exchange of heat between the system and its surroundings. Hence, under adiabatic conditions, *q*= 0.

Therefore, alternative (iii) is correct.

**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**](http://cbse.meritnation.com/study-online/solution/Chemistry/UZDvZIkn5wGw4B7hDYK1BQ%21%21#optionContent1)

The enthalpy of all elements in their standard state is zero.

Therefore, alternative (ii) is correct.

**Question 6.4:**

Δ*U*θof combustion of methane is – *X* kJ mol–1. The value of Δ*H*θ is

(i) = Δ*U*θ

(ii) > Δ*U*θ

(iii) < Δ*U*θ

(iv) = 0

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/ysU0WPl8VUV%40DQACX3Lu9g%21%21#optionContent1)

Since Δ*H*θ = Δ*U*θ + Δ*ng*R*T* and Δ*U*θ = –*X* kJ mol–1,

Δ*H*θ = (–*X*) + Δ*ng*R*T.*

⇒ Δ*H*θ < Δ*U*θ

Therefore, alternative (iii) is correct.

**Question 6.5:**

The enthalpy of combustion of methane, graphite and dihydrogen at 298 K are, –890.3 kJ mol–1 –393.5 kJ mol–1, and –285.8 kJ mol–1 respectively. Enthalpy of formation of CH4(*g*) will be

(i) –74.8 kJ mol–1             (ii) –52.27 kJ mol–1

(iii) +74.8 kJ mol–1           (iv) +52.26 kJ mol–1.

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/NH%24415ttIs9V8Pat%40EiiuQ%21%21#optionContent1)

According to the question,



Thus, the desired equation is the one that represents the formation of CH4 (*g*) i.e.,







Enthalpy of formation of CH4(*g*)= –74.8 kJ mol–1

Hence, alternative (i) is correct.

**Question 6.6:**

A reaction, A + B → 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

(iv) possible at any temperature

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/vq1dl%40vjCNdjl6rsfxKhOg%21%21#optionContent1)

For a reaction to be spontaneous, Δ*G* should be negative.

Δ*G* = Δ*H* – *T*Δ*S*

According to the question, for the given reaction,

Δ*S* = positive

Δ*H* = negative (since heat is evolved)

⇒ Δ*G* = negative

Therefore, the reaction is spontaneous at any temperature.

Hence, alternative (iv) is correct.

**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**](http://cbse.meritnation.com/study-online/solution/Chemistry/TIzYFoJ3X63QvFWjLHy1tw%21%21#optionContent1)

According to the first law of thermodynamics,

Δ*U* = *q* + *W* (i)

Where,

Δ*U* = change in internal energy for a process

*q* = heat

*W* = work

Given,

*q* = + 701 J (Since heat is absorbed)

*W* = –394 J (Since work is done by the system)

Substituting the values in expression (i), we get

Δ*U* = 701 J + (–394 J)

Δ*U* = 307 J

Hence, the change in internal energy for the given process is 307 J.

**Question 6.8:**

The reaction of cyanamide, NH2CN(*s*), with dioxygen was carried out in a bomb calorimeter, and Δ*U*was found to be –742.7 kJ mol–1 at 298 K. Calculate enthalpy change for the reaction at 298 K.



* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/6M8lEAOiSi1fGUr5JCHATg%21%21#optionContent1)

Enthalpy change for a reaction (Δ*H*) is given by the expression,

Δ*H* = Δ*U* + Δ*ng*R*T*

Where,

Δ*U* = change in internal energy

Δ*ng* = change in number of moles

For the given reaction,

Δ*ng* = ∑*ng* (products) – ∑*ng* (reactants)

= (2 – 2.5) moles

Δ*ng* = –0.5 moles

And,

Δ*U* = –742.7 kJ mol–1

*T* = 298 K

R = 8.314 × 10–3 kJ mol–1 K–1

Substituting the values in the expression of Δ*H*:

Δ*H* = (–742.7 kJ mol–1) + (–0.5 mol) (298 K) (8.314 × 10–3 kJ mol–1 K–1)

= –742.7 – 1.2

Δ*H* = –743.9 kJ mol–1

**Question 6.9:**

Calculate the number of kJ of heat necessary to raise the temperature of 60.0 g of aluminium from 35°C to 55°C. Molar heat capacity of Al is 24 J mol–1K–1.

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/4aUbE2L0%247E1eKug84kjWg%21%21#optionContent1)

From the expression of heat (*q*),

*q* = *m*. c. Δ*T*

Where,

c = molar heat capacity

*m* = mass of substance

Δ*T* = change in temperature

Substituting the values in the expression of *q*:



*q* = 1066.7 J

*q* = 1.07 kJ

**Question 6.10:**

Calculate the enthalpy change on freezing of 1.0 mol of water at 10.0°C to ice at –10.0°C. Δ*fusH*= 6.03 kJ mol–1 at 0°C.

*Cp*[H2O(l)] = 75.3 J mol–1 K–1

*Cp*[H2O(s)] = 36.8 J mol–1 K–1

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/9eUG3DKnigYvpvwL%245PopA%21%21#optionContent1)

Total enthalpy change involved in the transformation is the sum of the following changes:

(a) Energy change involved in the transformation of 1 mol of water at 10°C to 1 mol of water at 0°C.

(b) Energy change involved in the transformation of 1 mol of water at 0° to 1 mol of ice at 0°C.

(c) Energy change involved in the transformation of 1 mol of ice at 0°C to 1 mol of ice at –10°C.



= (75.3 J mol–1 K–1) (0 – 10)K + (–6.03 × 103 J mol–1) + (36.8 J mol–1 K–1) (–10 – 0)K

= –753 J mol–1 – 6030 J mol–1 – 368 J mol–1

= –7151 J mol–1

= –7.151 kJ mol–1

Hence, the enthalpy change involved in the transformation is –7.151 kJ mol–1.

**Question 6.11:**

Enthalpy of combustion of carbon to CO2 is –393.5 kJ mol–1. Calculate the heat released upon formation of 35.2 g of CO2 from carbon and dioxygen gas.

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/LcY%40z4Fu56qURzZp7WQLiw%21%21#optionContent1)

Formation of CO2 from carbon and dioxygen gas can be represented as:



(1 mole = 44 g)

Heat released on formation of 44 g CO2 = –393.5 kJ mol–1

Heat released on formation of 35.2 g CO2



= –314.8 kJ mol–1

**Question 6.12:**

Enthalpies of formation of CO(*g*),CO2(*g*),N2O(*g*)and N2O4(*g*)are –110 kJ mol–1, – 393 kJ mol–1, 81 kJ mol–1 and 9.7 kJ mol–1 respectively. Find the value of Δ*rH*for the reaction:

N2O4(*g*) + 3CO(*g*) N2O(*g*) + 3CO2(*g*)

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/C%24Ek2iy%24G043va6z1BrO%40A%21%21#optionContent1)

Δ*rH* for a reaction is defined as the difference between Δ*fH* value of products and Δ*fH* value of reactants.



For the given reaction,

N2O4(*g*) + 3CO(*g*)  N2O(*g*) + 3CO2(*g*)



Substituting the values of Δ*fH* for N2O, CO2, N2O4, and CO from the question, we get:



Hence, the value of Δ*rH*for the reaction is.

**Question 6.13:**

Given

; Δ*rH*θ = –92.4 kJ mol–1

What is the standard enthalpy of formation of NH3 gas?

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/qCUXBZsAsTrHrEJJMlcapA%21%21#optionContent1)

Standard enthalpy of formation of a compound is the change in 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.

Re-writing the given equation for 1 mole of NH3(*g*),



Standard enthalpy of formation of NH3(*g*)

= ½ Δ*rH*θ

= ½ (–92.4 kJ mol–1)

= –46.2 kJ mol–1

**Question 6.14:**

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

CH3OH(*l*) + O2(*g*)  CO2(*g*) + 2H2O(*l*) ; Δ*rH*θ = –726 kJ mol–1

C(*g*) + O2(*g*)  CO2(*g*) ; Δ*cH*θ = –393 kJ mol–1

H2(*g*) +O2(*g*)  H2O(*l*) ; Δ*fH*θ = –286 kJ mol–1.

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/QEeEVXRyYWq5ABrP4HnXwQ%21%21#optionContent1)

The reaction that takes place during the formation of CH3OH(*l*) can be written as:

C(*s*) + 2H2O(*g*) + O2(*g*)  CH3OH(*l*) (**1**)

The reaction (**1**) can be obtained from the given reactions by following the algebraic calculations as:

Equation (ii) + 2 × equation (iii) – equation (i)

Δ*fH*θ [CH3OH(*l*)] = Δc*H*θ+ 2Δ*fH*θ [H2O(*l*)] – Δ*rH*θ

= (–393 kJ mol–1) + 2(–286 kJ mol–1) – (–726 kJ mol–1)

= (–393 – 572 + 726) kJ mol–1

Δ*fH*θ [CH3OH(*l*)] = –239 kJ mol–1

**Question 6.15:**

Calculate the enthalpy change for the process

CCl4(*g*) → C(*g*) + 4Cl(*g*)

and calculate bond enthalpy of C–Cl in CCl4(*g*).

Δ*vapH*θ (CCl4) = 30.5 kJ mol–1.

Δ*fH*θ (CCl4) = –135.5 kJ mol–1.

Δ*aH*θ (C) = 715.0 kJ mol–1, where Δ*aH*θ is enthalpy of atomisation

Δ*aH*θ (Cl2) = 242 kJ mol–1

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/4zz%40Pd3Y%40aRAWzCDtWOMsA%21%21#optionContent1)

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

Δ*vapH*θ = 30.5 kJ mol–1

Δ*aH*θ = 715.0 kJ mol–1

 Δ*aH*θ = 242 kJ mol–1

 Δ*fH* = –135.5 kJ mol–1

Enthalpy change for the given process can be calculated using the following algebraic calculations as:

Equation (ii) + 2 × Equation (iii) – Equation (i) – Equation (iv)

Δ*H* = Δ*aH*θ(C) + 2Δ*aH*θ (Cl2) – Δ*vapH*θ – Δ*fH*

= (715.0 kJ mol–1) + 2(242 kJ mol–1) – (30.5 kJ mol–1) – (–135.5 kJ mol–1)

Δ*H* = 1304 kJ mol–1

Bond enthalpy of C–Cl bond in CCl4 (*g*)



= 326 kJ mol–1

**Question 6.16:**

For an isolated system, Δ*U*= 0, what will be Δ*S?*

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/dGeK6FJ%40yQh8WCXOxYKDiA%21%21#optionContent1)

ΔS will be positive i.e., greater than zero

Since Δ*U* = 0, Δ*S* will be positive and the reaction will be spontaneous.

**Question 6.17:**

For the reaction at 298 K,

2A + B → C

Δ*H*= 400 kJ mol–1 and Δ*S*= 0.2 kJ K–1 mol–1

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

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/8dgcMnYYiDeoZJDRMYc76w%21%21#optionContent1)

From the expression,

Δ*G* = Δ*H* – *T*Δ*S*

Assuming the reaction at equilibrium, Δ*T* for the reaction would be:



(Δ*G* = 0 at equilibrium)



*T* = 2000 K

For the reaction to be spontaneous, Δ*G* must be negative. Hence, for the given reaction to be spontaneous, *T* should be greater than 2000 K.

**Question 6.18:**

For the reaction,

2Cl(*g)* → Cl2(*g*), what are the signs of Δ*H*and Δ*S*?

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/t2CEdYiRwK4mWfGpO5onAg%21%21#optionContent1)

Δ*H* and Δ*S*are negative

The given reaction represents the formation of chlorine molecule from chlorine atoms. Here, bond formation is taking place. Therefore, energy is being released. Hence, Δ*H* is negative.

Also, two moles of atoms have more randomness than one mole of a molecule. Since spontaneity is decreased, Δ*S* is negative for the given reaction.

**Question 6.19:**

For the reaction

2A(*g*) + B(*g*) → 2D(*g*)

Δ*U*θ = –10.5 kJ and Δ*S*θ= –44.1 JK–1.

Calculate Δ*G*θ for the reaction, and predict whether the reaction may occur spontaneously.

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/v9Sr8VoIohIGTX4lHSuj9g%21%21#optionContent1)

For the given reaction,

2 A(*g*) + B(*g*) → 2D(*g*)

Δ*ng* = 2 – (3)

= –1 mole

Substituting the value of Δ*U*θ in the expression of Δ*H*:

Δ*H*θ = Δ*U*θ + Δ*ng*R*T*

= (–10.5 kJ) – (–1) (8.314 × 10–3 kJ K–1 mol–1) (298 K)

= –10.5 kJ – 2.48 kJ

Δ*H*θ = –12.98 kJ

Substituting the values of Δ*H*θ and Δ*S*θ in the expression of Δ*G*θ:

Δ*G*θ = Δ*H*θ – *T*Δ*S*θ

= –12.98 kJ – (298 K) (–44.1 J K–1)

= –12.98 kJ + 13.14 kJ

Δ*G*θ = + 0.16 kJ

Since Δ*G*θ for the reaction is positive, the reaction will not occur spontaneously.

**Question 6.20:**

The equilibrium constant for a reaction is 10. What will be the value of Δ*G*θ? R = 8.314 JK–1 mol–1, *T* = 300 K.

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/oHDDwvpkKQddvtVRfLs71A%21%21#optionContent1)

From the expression,

Δ*G*θ = –2.303 R*T* log*Keq*

Δ*G*θ for the reaction,

= (2.303) (8.314 JK–1 mol–1) (300 K) log10

= –5744.14 Jmol–1

= –5.744 kJ mol–1

**Question 6.21:**

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

N2(*g*)+ O2(*g*)→ NO(*g*) ; Δ*rH*θ = 90 kJ mol–1

NO(*g*) +O2(*g*)→ NO2(*g*): Δ*rH*θ= –74 kJ mol–1

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/vaActFlC3fHGlgTFo3uZ0Q%21%21#optionContent1)

The positive value of Δ*rH* indicates that heat is absorbed during the formation of NO(*g*). This means that NO(*g*) has higher energy than the reactants (N2 and O2). Hence, NO(*g*) is unstable.

The negative value of Δ*rH* indicates that heat is evolved during the formation of NO2(*g*) from NO(*g*) and O2(*g*). The product, NO2(*g*) is stabilized with minimum energy.

Hence, unstable NO(*g*) changes to stable NO2(*g*)

**Question 6.22:**

Calculate the entropy change in surroundings when 1.00 mol of H2O(*l*) is formed under standard conditions. Δ*fH*θ = –286 kJ mol–1.

* [**Answer**](http://cbse.meritnation.com/study-online/solution/Chemistry/qkk81tXsLyhPCTwtx3u4fA%21%21#optionContent1)

It is given that 286 kJ mol–1 of heat is evolved on the formation of 1 mol of H2O(*l*). Thus, an equal amount of heat will be absorbed by the surroundings.

*qsurr* = +286 kJ mol–1

Entropy change (Δ*Ssurr*) for the surroundings = 



Δ*Ssurr* = 959.73 J mol–1 K–1