

## THERMODYNAMICS' PROPERTIES AND PROCESS - LEVEL I

**1** A necessary condition for an adiabatic change is-

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

**2** Which, among the following, is not a state function?

1. Internal energy
2. Free energy
3. Work
4. Enthalpy

**3** Which one of the following statements is false?

- |    |  |
|----|--|
| 1. | Temperature is a state function.   |
| 2. | Work is a state function.  |
| 3. | Change in the state is completely defined when the initial and final states are specified. |
| 4. | Work appears at the boundary of the system.  |

**4** The enthalpy of formation of all elements in their standard state is-

1. Unity
2. Zero
3. Less than zero
4. Different for each element

**5** Thermodynamics is not concerned about-

1. Energy changes involved in a chemical reaction
2. The extent to which a chemical reaction proceeds
3. The rate at which a reaction proceeds
4. The feasibility of a chemical reaction

**6** In an adiabatic process, no transfer of heat takes place between the system and its surroundings. The correct option for free expansion of an ideal gas under adiabatic condition from the following is -

1.  $q = 0$ ,  $\Delta T \neq 0$ ,  $W = 0$
2.  $q \neq 0$ ,  $\Delta T = 0$ ,  $W = 0$
3.  $q = 0$ ,  $\Delta T = 0$ ,  $W = 0$
4.  $q = 0$ ,  $\Delta T = 0$ ,  $W \neq 0$

**7** The amount of heat needed to raise the temperature of 60.0 g of aluminium from 35°C to 55°C would be - (Molar heat capacity of Al is  $24 \text{ J mol}^{-1} \text{ K}^{-1}$ )

1. 1.07 J
2. 1.07 kJ
3. 106.7 kJ
4. 100.7 kJ

**8** The correct statement among the following is-

- |    |   |
|----|---|
| 1. | The presence of reacting species in a covered beaker is an example of an open system.                               |
| 2. | There is an exchange of energy as well as matter between the system and the surroundings in a closed system.        |
| 3. | The presence of reactants in a closed vessel made up of copper is an example of a closed system.                    |
| 4. | The presence of reactants in a thermos flask or any other closed insulated vessel is an example of a closed system. |

## THERMODYNAMICS' PROPERTIES AND PROCESS - LEVEL II

**9** A Beckmann thermometer is used to measure -

1. High temperature
2. Low temperature
3. Normal temperature
4. All temperatures

**10** Work done is maximum in -

1. Isobaric work
2. Isothermal work
3. Isochoric work
4. Adiabatic work

**11** When a gas undergoes adiabatic expansion, it gets cooled due to -

1. Loss of energy
2. Fall in pressure
3. Decrease in velocity
4. Increase in energy with work done

## CLASSIFICATION OF SYSTEM, EXTENSIVE & INTENSIVE PROPERTIES - LEVEL I

**12** Which of the following is not a state function?

1.	$\Delta S$	2.	$\Delta G$
3.	$\Delta H$	4.	$\Delta Q$

**13** If the volume of a gas is reduced to half from its original volume, then the specific heat will-

1.	Reduce to half	2.	Be Doubled
3.	Remain constant	4.	Increase four times

**14** Which among the following state functions is an extensive property of the system?

1.	Temperature	2.	Volume
3.	Refractive index	4.	Viscosity

**15** Plants and living beings are examples of -

1.	Isolated system	2.	Adiabatic system
3.	Open system	4.	Closed system

**16**

<b>Assertion (A):</b>	Specific heat is an intensive property.
<b>Reason (R):</b>	Heat capacity is an extensive property.

1.	Both (A) and (R) are true and (R) is the correct explanation of (A).
2.	Both (A) and (R) are true but (R) is not the correct explanation of (A).
3.	(A) is true but (R) is false.
4.	Both (A) and (R) are false.

## CLASSIFICATION OF SYSTEM, EXTENSIVE & INTENSIVE PROPERTIES - LEVEL II

**17** Which of the following pairs correctly represents intensive property?

1. Entropy, Gibb's energy
2. Enthalpy, Heat capacity
3. Electrode potential, Vapour pressure
4. Resistance, Conductance

**18** For an isolated system with  $\Delta U = 0$ , the  $\Delta S$  value will be-

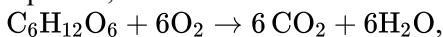
1. Positive
2. Negative
3. Zero
4. Not possible to define

## ENTHALPY & INTERNAL ENERGY - LEVEL I

**19** Assume each reaction is carried out in an open container. For which of the following reactions will  $\Delta H$  be equal to  $\Delta U$ ?

1.  $\text{PCl}_3(\text{g}) \rightarrow \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$
2.  $2\text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g})$
3.  $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightarrow 2\text{HBr}(\text{g})$
4.  $\text{C}(\text{s}) + 2\text{H}_2\text{O}(\text{g}) \rightarrow 2\text{H}_2(\text{g}) + \text{CO}_2(\text{g})$

**20** Combustion of glucose takes place according to the equation,

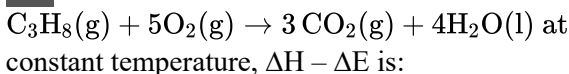


$$\Delta H = -72 \text{ K cal}$$

Energy required for the production of 1.6 g of glucose is - (Molecular mass of glucose = 180 g)

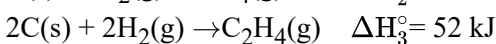
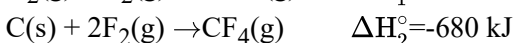
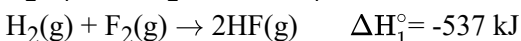
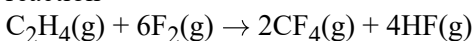
1. 0.064 kcal
2. 0.64 kcal
3. 6.4 kcal
4. 64 kcal

**21** For the reaction:



1.	+ RT	2.	- 3RT
3.	+ 3RT	4.	- RT

**22** For the following given equations and  $\Delta H^\circ$  values, determine the enthalpy of reaction at 298 K for the reaction-



1. -1165
2. -2486
3. +1165
4. +2486

23

**Assertion (A):**  $\Delta H$  is positive for endothermic reactions.

**Reason (R):** If the total enthalpy of reactants and products are  $H_R$  &  $H_P$  respectively, then for an endothermic reaction  $H_R < H_P$ .

1. Both (A) and (R) are true and (R) is the correct explanation of (A).
2. Both (A) and (R) are true but (R) is not the correct explanation of (A).
3. (A) is true but (R) is false.
4. Both (A) and (R) are false.

24 The bond energies of  $C \equiv C$ ,  $C-H$ ,  $H-H$ , and  $C=C$  are 198, 98, 103 and 145 kcal respectively.

The enthalpy change of the reaction  $HC \equiv CH + H_2 \rightarrow C_2H_4$  would be-

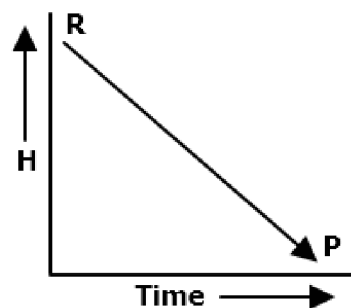
1. 48 kcal
2. 96 kcal
3. -40 kcal
4. -152 kcal

25 During complete combustion of one mole of butane, 2658 kJ of heat is released. The thermochemical reaction for this change is -

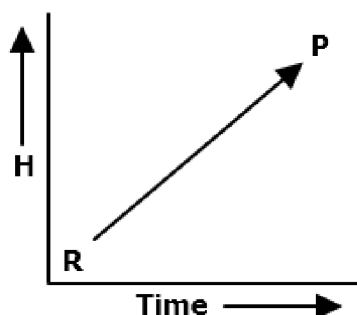
1.  $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$ ;  
 $\Delta_c H = -2658.0 \text{ kJ mol}^{-1}$
2.  $C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$ ;  
 $\Delta_c H = -1329.0 \text{ kJ mol}^{-1}$
3.  $C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$ ;  
 $\Delta_c H = -2658.0 \text{ kJ mol}^{-1}$
4.  $C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$ ;  
 $\Delta_c H = +2658.0 \text{ kJ mol}^{-1}$

26 Which plot represents an exothermic reaction?

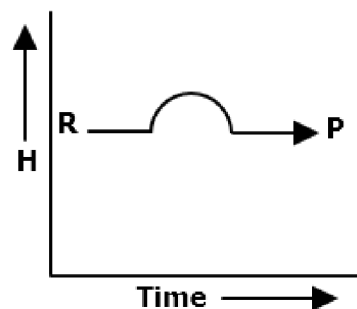
1.



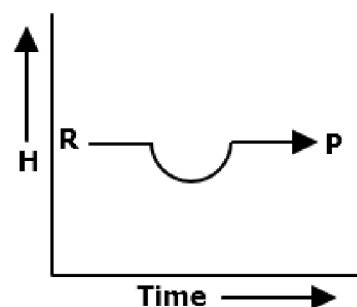
2.



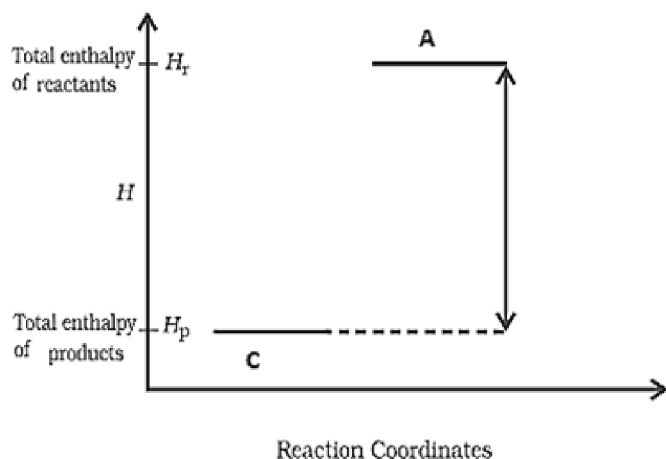
3.



4.



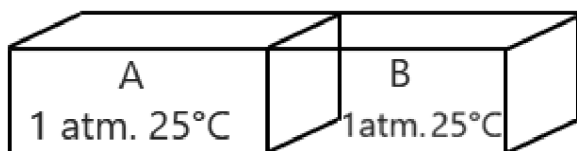
- 27** Consider the following diagram for a reaction  $A \rightarrow C$ :



The nature of the reaction is-

1. Exothermic
2. Endothermic
3. Reaction at equilibrium
4. None of the above

- 28** As an isolated box, equally partitioned, contains two ideal gasses A and B as shown:



When the partition is removed, the gases mix. The changes in enthalpy ( $\Delta H$ ) and entropy ( $\Delta S$ ) in the process, respectively, are

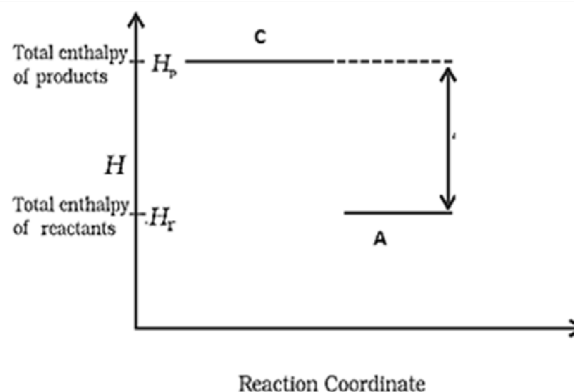
1. Zero, positive
2. Zero, negative
3. Positive, zero
4. Negative, zero

- 29** For the given reaction

$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$ , the heat of formations of  $H_2O_2(l)$  and  $H_2O(l)$  are  $-188 \text{ kJ/mol}$  &  $-286 \text{ kJ/mol}$  respectively. The change in the enthalpy of the reaction will be:

1.  $-196 \text{ kJ/mol}$
2.  $+196 \text{ kJ/mol}$
3.  $+948 \text{ kJ/mol}$
4.  $-948 \text{ kJ/mol}$

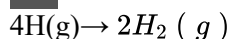
- 30** Consider the following diagram for a reaction .



The nature of the reaction is-

1. Exothermic
2. Endothermic
3. Reaction at equilibrium
4. None of the above

- 31** Given the following reaction:

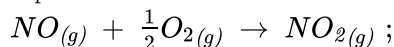


The enthalpy change for the reaction is  $-869.6 \text{ kJ}$ . The dissociation energy of the H-H bond is:

1.  $-869.6 \text{ kJ}$
2.  $+434.8 \text{ kJ}$
3.  $+217.4 \text{ kJ}$
4.  $-434.8 \text{ kJ}$

- 32**  $\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightarrow NO(g)$  ;

$$\Delta_r H^\circ = 90 \text{ kJ mol}^{-1}$$

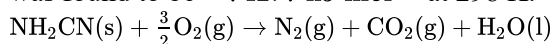


$$\Delta_r H^\circ = -74 \text{ kJ mol}^{-1}$$

The thermodynamic stability of  $NO(g)$  based on the above data is-

1. Less than  $NO_2(g)$
2. More than  $NO_2(g)$
3. Equal to  $NO_2(g)$
4. Insufficient data

**33** The reaction of cyanamide,  $\text{NH}_2\text{CN}$  (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 K.



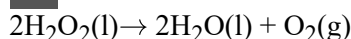
The enthalpy change for the reaction at 298 K would be -

1.  $-741.3 \text{ kJ mol}^{-1}$
2.  $+753.9 \text{ kJ mol}^{-1}$
3.  $+772.7 \text{ kJ mol}^{-1}$
4.  $-845.1 \text{ kJ mol}^{-1}$

**34** The  $\Delta H$  for vaporization of a liquid is  $20 \text{ kJ/mol}$ . Assuming ideal behaviour, the change in internal energy for the vaporization of 1 mol of the liquid at  $60^\circ\text{C}$  and 1 bar is close to:

1.  $13.2 \text{ kJ/mol}$
2.  $17.2 \text{ kJ/mol}$
3.  $19.5 \text{ kJ/mol}$
4.  $20.0 \text{ kJ/mol}$

**35** Change in enthalpy for reaction,



If the heat of formation of  $\text{H}_2\text{O}_2(\text{l})$  and  $\text{H}_2\text{O(l)}$  are  $-188$  and  $-286 \text{ kJ/mol}$  respectively is -

1.  $-196 \text{ kJ/mol}$
2.  $+196 \text{ kJ/mol}$
3.  $+948 \text{ kJ/mol}$
4.  $-948 \text{ kJ/mol}$

**36** For the reaction,  $2\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{N}_2\text{O(g)}$ , at 298K  $\Delta H$  is  $164 \text{ kJ mol}^{-1}$ . The  $\Delta E$  of the reaction is-

1.  $166.5 \text{ kJ mol}^{-1}$
2.  $141.5 \text{ kJ mol}^{-1}$
3.  $104.0 \text{ kJ mol}^{-1}$
4.  $-169 \text{ kJ mol}^{-1}$

**37**  $\Delta_f U^\ominus$  of formation of  $\text{CH}_4(\text{g})$  at certain temperature is  $-393 \text{ kJ mol}^{-1}$ . The value of  $\Delta_f H^\ominus$  is-

1. Zero
2.  $< \Delta_f U^\ominus$
3.  $> \Delta_f U^\ominus$
4. Equal to  $\Delta_f U^\ominus$

## ENTHALPY & INTERNAL ENERGY - LEVEL II

**38** An ideal gas absorbs 2000 cal of heat from a heat reservoir and does mechanical work equivalent to 4200 J. The change in internal energy of the gas is-

1. 3000 cal
2. 2000 cal
3. 1500 cal
4. 1000 cal

**39**  $\Delta U^\circ$  for combustion of methane is  $-x \text{ kJ mol}^{-1}$ . The value of  $\Delta H^\circ$  for the same reaction would be -

1.  $= \Delta U^\circ$
2.  $> \Delta U^\circ$
3.  $< \Delta U^\circ$
4.  $= 0$

**40**

<b>Assertion (A):</b>	$\text{C}_{\text{diamond}} \rightarrow \text{C}_{\text{graphite}}$ $\Delta H$ and $\Delta U$ are the same for this reaction.
<b>Reason (R):</b>	Entropy increases during the conversion of diamond to graphite.
1.	Both (A) and (R) are true and (R) is the correct explanation of (A).
2.	Both (A) and (R) are true but (R) is not the correct explanation of (A).
3.	(A) is true but (R) is false.
4.	Both (A) and (R) are false.

**41** The difference in  $\Delta H$  and  $\Delta U$  for combustion of benzoic acid at 300 K is equal to-

1.  $-1.247 \text{ kJ}$
2.  $+1.247 \text{ kJ}$
3.  $-1.247 \text{ J}$
4.  $+1.247 \text{ J}$

**42** The standard enthalpy of vaporisation  $\Delta_{\text{vap}} H^\circ$  for water at  $100^\circ\text{C}$  is  $40.66 \text{ kJ mol}^{-1}$ . The internal energy of vaporisation of water at  $100^\circ\text{C}$  (in  $\text{kJ mol}^{-1}$ ) is-

(Assume water vapour to behave like an ideal gas)

1.  $+37.56$
2.  $-43.76$
3.  $+43.76$
4.  $+40.66$

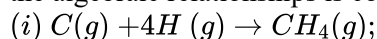
**43** The lattice energy of NaCl is  $780 \text{ kJ mol}^{-1}$ . The enthalpy of hydration of  $\text{Na}^+(\text{g})$  and  $\text{Cl}^-(\text{g})$  ions are  $-406 \text{ kJ mol}^{-1}$  and  $-364 \text{ kJ mol}^{-1}$ . The enthalpy of the solution of NaCl(s) is-

1.  $23 \text{ kJ mol}^{-1}$
2.  $10 \text{ kJ mol}^{-1}$
3.  $-10 \text{ kJ mol}^{-1}$
4.  $-82 \text{ kJ mol}^{-1}$

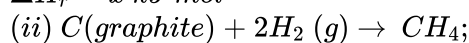
**44** 2.1 g of Fe combines with S evolving 3.77 kJ. The heat of formation of FeS in kJ/mole is-

1. - 3.77
2. - 1.79
3. - 100.5
4. None of the above

**45** Consider the reactions given below. On the basis of these reactions find out which of the algebraic relationships is correct?



$$\Delta H_r = x \text{ kJ mol}^{-1}$$



$$\Delta H_r = y \text{ kJ mol}^{-1}$$

1.  $x=y$
2.  $x=2y$
3.  $x>y$
4.  $x<y$

### CP & CV - LEVEL I

**46** The factor  $\left(\frac{\partial Q}{\partial T}\right)_P - \left(\frac{\partial Q}{\partial T}\right)_V$  is equal to :

1.  $\gamma$
2. R
3.  $\frac{R}{M}$
4.  $\Delta nRT$

**47** Two moles of an ideal gas is heated at a constant pressure of one atmosphere from  $27^\circ\text{C}$  to  $127^\circ\text{C}$ . If  $C_{v,m} = 20 + 10^{-2} T \text{ JK}^{-1} \text{ mol}^{-1}$ , then q and  $\Delta U$  for the process are respectively:

1. 6362.8 J, 4700 J
2. 3037.2 J, 4700 J
3. 7062.8 J, 5400 J
4. 3181.4 J, 2350 J

### FIRST LAW OF THERMODYNAMICS - LEVEL I

**48**

**Assertion (A):** Total enthalpy change of a multistep process is the sum of  $\Delta H_1 + \Delta H_2 + \Delta H_3 + \dots$

**Reason (R):** When heat is absorbed by the system, the sign of q is taken to be negative.

1. Both (A) and (R) are true and (R) is the correct explanation of (A).
2. Both (A) and (R) are true but (R) is not the correct explanation of (A).
3. (A) is true but (R) is false.
4. Both (A) and (R) are false.

**49** A piston filled with 0.04 mol of an ideal gas expands reversibly from 50.0 mL to 375 mL at a constant temperature of  $37.0^\circ\text{C}$ . As it does so, it absorbs 208 J of heat. The values of q and w for the process will be- ( $R = 8.314 \text{ J/mol K}$ ) ( $\ln 7.5 = 2.01$ )

- |                            |                            |
|----------------------------|----------------------------|
| 1. q = +208 J, w = -208 J  | 2. q = -208 J, w = -208 J  |
| 3. q = -208 J, w = + 208 J | 4. q = +208 J, w = + 208 J |

**50** 701 J of heat is absorbed by a system and 394 J of work is done by the system. The change in internal energy for the process is-

- |           |           |
|-----------|-----------|
| 1. 307 J  | 2. -307 J |
| 3. 1095 J | 4. -701 J |

**51**

**Assertion (A):** Work done in an irreversible isothermal process at constant volume is zero.

**Reason (R):** Work is assigned a negative sign during expansion and is assigned a positive sign during compression.

1. Both (A) and (R) are true and (R) is the correct explanation of (A).
2. Both (A) and (R) are true but (R) is not the correct explanation of (A).
3. (A) is true but (R) is false.
4. Both (A) and (R) are false.

**52** An ideal gas expands isothermally from  $10^{-3} \text{ m}^3$  to  $10^{-2} \text{ m}^3$  at 300 K against a constant pressure of  $10^5 \text{ Nm}^{-2}$ . The work done by the gas is:

1.	+270 kJ	2.	-900 J
3.	+900 kJ	4.	-900 kJ

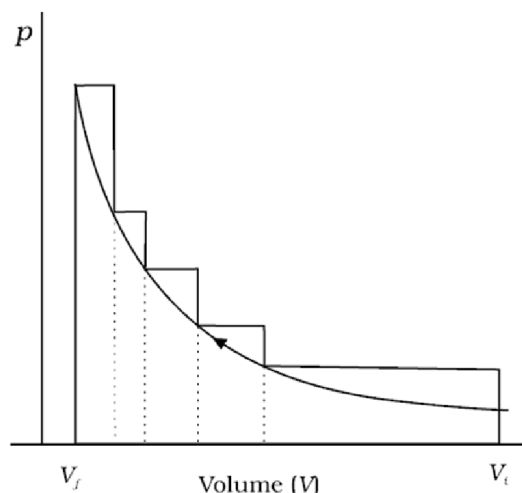
**53** The work done in ergs for the reversible expansion of one mole of an ideal gas from a volume of 10 liters to 20 liters at  $25^\circ\text{C}$  is -

- $-2.303 \times 298 \times 0.082 \log 2$
- $-298 \times 10^7 \times 8.31 \times 2.303 \log 2$
- $2.303 \times 298 \times 0.082 \log 0.5$
- $-8.31 \times 10^7 \times 298 - 2.303 \log 0.5$

**54** The pressure-volume work for an ideal gas can be calculated by using the expression  $W = \int_{V_i}^{V_f} p_{\text{ex}} dV$ . The work can also be calculated from the pV-plot by using the area under the curve within the specified limits. An ideal gas is compressed (a) reversibly or (b) irreversibly from volume  $V_i$  to  $V_f$ . The correct option is -

- $W (\text{reversible}) = W (\text{irreversible})$
- $W (\text{reversible}) < W (\text{irreversible})$
- $W (\text{reversible}) > W (\text{irreversible})$
- $W (\text{reversible}) = W (\text{irreversible}) + p_{\text{ex}} \cdot \Delta V$

**55** For the graph given below, it can be concluded that work done during the process shown will be-



1.	Zero	2.	Negative
3.	Positive	4.	Cannot be determined

**56** Which of the following is correct for isothermal expansion of an ideal gas -

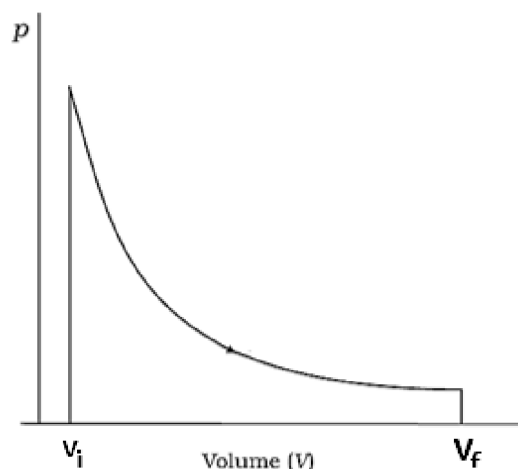
- $W_{\text{rev}} = W_{\text{irr}}$
- $W_{\text{rev}} + W_{\text{irr}} = 0$
- $W_{\text{rev}} > W_{\text{irr}}$
- $q_{\text{rev}} = q_{\text{irr}}$

## FIRST LAW OF THERMODYNAMICS - LEVEL II

**57** To calculate the amount of work done in joules during reversible isothermal expansion of an ideal gas, the volume must be expressed in:

1.	$\text{m}^3$ only	2.	$\text{dm}^3$ only
3.	$\text{cm}^3$ only	4.	All of the above

**58** Consider the following graph.



The work done shown by the above-mentioned graph is-

1.	Positive	2.	Negative
3.	Zero	4.	Cannot be determined

**59** When 1 g  $\text{H}_2$  gas at S.T.P is expanded to twice its initial volume, then the work done is -

- 22.4 L atm
- 5.6 L atm
- 11.2 L atm
- 44.8 L atm

**60** An ideal gas is expanded irreversibly against 10 bar pressure from 20 litres to 30 litres. Calculate 'w' if the process is isenthalpic.

1.	0	2.	+100J
3.	-100 J	4.	-10 kJ

## THERMOCHEMISTRY - LEVEL I

**61** The standard enthalpy of combustion at  $25^\circ\text{C}$  of hydrogen, cyclohexene ( $\text{C}_6\text{H}_{10}$ ), and cyclohexane ( $\text{C}_6\text{H}_{12}$ ) are -241, -3800 and -3920  $\text{kJ mol}^{-1}$ , respectively. Calculate the standard enthalpy of hydrogenation of cyclohexene.

1.	-131 $\text{kJ mol}^{-1}$	2.	-155 $\text{kJ mol}^{-1}$
3.	-167 $\text{kJ mol}^{-1}$	4.	-121 $\text{kJ mol}^{-1}$

**62** The enthalpy of sublimation of a substance is equal to-

1. Enthalpy of fusion + Enthalpy of vaporisation
2. Enthalpy of fusion
3. Enthalpy of vaporisation
4. Twice the enthalpy of vaporisation

**63**

$\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$ ;  $\Delta_r H^\circ = -92.4 \text{ kJ mol}^{-1}$ . The standard enthalpy of formation of  $\text{NH}_3$  gas in the above reaction would be-

1.	-92.4 J (mol) <sup>-1</sup>	2.	-46.2 kJ (mol) <sup>-1</sup>
3.	+46.2 J (mol) <sup>-1</sup>	4.	+92.4 kJ (mol) <sup>-1</sup>

**64** From the following bond energies:

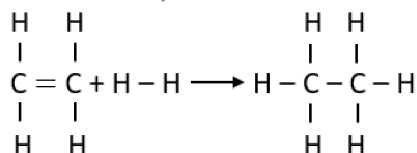
H—H bond energy: 431.37  $\text{kJ mol}^{-1}$

C=C bond energy: 606.10  $\text{kJ mol}^{-1}$

C—C bond energy: 336.49  $\text{kJ mol}^{-1}$

C—H bond energy: 410.50  $\text{kJ mol}^{-1}$

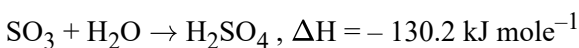
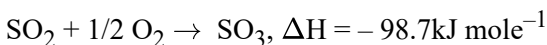
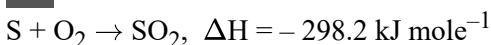
Enthalpy for the reaction,



will be:

1.	1523.6 $\text{kJ mol}^{-1}$	2.	-243.6 $\text{kJ mol}^{-1}$
3.	-120.0 $\text{kJ mol}^{-1}$	4.	553.0 $\text{kJ mol}^{-1}$

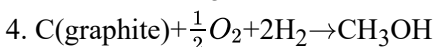
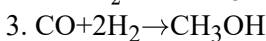
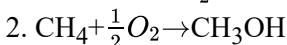
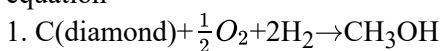
**65** Consider the following reaction,



the enthalpy of formation of  $\text{H}_2\text{SO}_4$  at 298 K will be-

1.	-814.4 $\text{kJ mole}^{-1}$	2.	+814.4 $\text{kJ mole}^{-1}$
3.	-650.3 $\text{kJ mole}^{-1}$	4.	-433.7 $\text{kJ mole}^{-1}$

**66**  $\Delta H_f^\circ$  (298K) of methanol is given by the chemical equation -



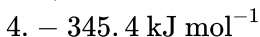
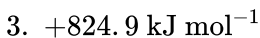
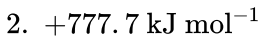
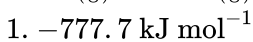
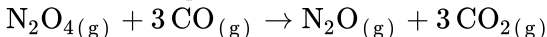
**67**

<b>Assertion (A):</b>	Combustion of all organic compounds is an exothermic reaction.
<b>Reason (R):</b>	The enthalpies of all elements in their standard state are zero.
1.	Both (A) and (R) are true and (R) is the correct explanation of (A).
2.	Both (A) and (R) are true but (R) is not the correct explanation of (A).
3.	(A) is true but (R) is false.
4.	(A) is false but (R) is true.

**68** The enthalpy of formation of  $\text{CO}_{(g)}$ ,  $\text{CO}_{2(g)}$ ,  $\text{N}_2\text{O}_{(g)}$ , and  $\text{N}_2\text{O}_{4(g)}$  are

-110  $\text{kJ mol}^{-1}$ , -393  $\text{kJ mol}^{-1}$ , 81  $\text{kJ mol}^{-1}$ , and 9.7  $\text{kJ mol}^{-1}$  respectively.

The value of  $\Delta_r H$  for the reaction would be-





**69** The enthalpy of combustion of carbon to  $\text{CO}_2$  is  $-393.5 \text{ kJ mol}^{-1}$ . The amount of heat released upon formation of  $35.2 \text{ g}$  of  $\text{CO}_2$  from carbon and dioxygen gas would be-

1.  $-393.5 \text{ kJ mol}^{-1}$
2.  $-314.8 \text{ kJ mol}^{-1}$
3.  $+314.8 \text{ kJ mol}^{-1}$
4.  $-320.5 \text{ kJ mol}^{-1}$

**70** Which of the following is not an endothermic reaction?

1. Combustion of methane
2. Decomposition of water
3. Dehydrogenation of ethane or ethylene
4. Conversion of graphite to diamond

**71** When  $4 \text{ g}$  of iron is burnt to ferric oxide at a constant pressure,  $29.28 \text{ kJ}$  of heat is evolved.

The enthalpy of formation of ferric oxide will be- (At. mass of  $\text{Fe} = 56$ ) ?

1.  $-81.98 \text{ kJ}$
2.  $-819.8 \text{ kJ}$
3.  $-40.99 \text{ kJ}$
4.  $+819.8 \text{ kJ}$

**72**  $\Delta_{\text{vap}}H^\circ (\text{CCl}_4) = 30.5 \text{ kJ mol}^{-1}$

$\Delta_fH^\circ (\text{CCl}_4) = -135.5 \text{ kJ mol}^{-1}$

$\Delta_aH^\circ (\text{C}) = 715.0 \text{ kJ mol}^{-1}$

$\Delta_aH^\circ (\text{Cl}_2) = 242 \text{ kJ mol}^{-1}$

The enthalpy change for the reaction

$\text{CCl}_4(\text{g}) \rightarrow \text{C}(\text{g}) + 4\text{Cl}(\text{g})$  would be -

1.  $326 \text{ kJ mol}^{-1}$
2.  $1304 \text{ kJ mol}^{-1}$
3.  $-328 \text{ kJ mol}^{-1}$
4.  $-1304 \text{ kJ mol}^{-1}$

**73** The enthalpy of combustion of methane, graphite, and dihydrogen at  $298 \text{ K}$  are,

$-890.3 \text{ kJ mol}^{-1}$ ,  $-393.5 \text{ kJ mol}^{-1}$ , and  $-285.8 \text{ kJ mol}^{-1}$  respectively.

The enthalpy of formation of  $\text{CH}_4(\text{g})$  is-

1.  $-74.8 \text{ kJ mol}^{-1}$
2.  $-52.27 \text{ kJ mol}^{-1}$
3.  $+74.8 \text{ kJ mol}^{-1}$
4.  $+52.26 \text{ kJ mol}^{-1}$

**74** The standard enthalpy of the formation of  $\text{CH}_3\text{OH}(\text{l})$  from the following data is:

$\text{CH}_3\text{OH}(\text{l}) + \frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}); \Delta_rH^\circ = -726 \text{ kJ mol}^{-1}$			
$\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}); \Delta_cH^\circ = -393 \text{ kJ mol}^{-1}$			
$\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}); \Delta_fH^\circ = -286 \text{ kJ mol}^{-1}$			
1.	$-239 \text{ kJ mol}^{-1}$	2.	$+239 \text{ kJ mol}^{-1}$
3.	$-47 \text{ kJ mol}^{-1}$	4.	$+47 \text{ kJ mol}^{-1}$

## THERMOCHEMISTRY - LEVEL II

**75** Determine the enthalpy of formation for  $\text{H}_2\text{O}_2(\text{l})$ , using the following enthalpies of reaction :

$\text{N}_2\text{H}_4(\text{l}) + 2\text{H}_2\text{O}_2(\text{l}) \rightarrow \text{N}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l}); \Delta H_1^\circ = -818 \text{ kJ/mol}$

$\text{N}_2\text{H}_4(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}); \Delta H_2^\circ = -622 \text{ kJ/mol}$

$\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}); \Delta H_3^\circ = -285 \text{ kJ/mol}$

1.  $-383 \text{ kJ/mol}$
2.  $-187 \text{ kJ/mol}$
3.  $-49 \text{ kJ/mol}$
4. None of the above

**76** The bond energy of  $\text{H}-\text{H}$  and  $\text{Cl}-\text{Cl}$  is  $430 \text{ kJ mol}^{-1}$  and  $240 \text{ kJ mol}^{-1}$  respectively and  $\Delta H_f^\circ$  for  $\text{HCl}$  is  $-90 \text{ kJ mol}^{-1}$ . The bond enthalpy of  $\text{HCl}$  is:

1.  $290 \text{ kJ mol}^{-1}$
2.  $380 \text{ kJ mol}^{-1}$
3.  $425 \text{ kJ mol}^{-1}$
4.  $245 \text{ kJ mol}^{-1}$

**77**

<b>Assertion (A):</b>	Combustion is an exothermic process.
<b>Reason (R):</b>	Combustion is a spontaneous process.
1.	Both (A) and (R) are true and (R) is the correct explanation of (A).
2.	Both (A) and (R) are true but (R) is not the correct explanation of (A).
3.	(A) is true but (R) is false.
4.	Both (A) and (R) are false.

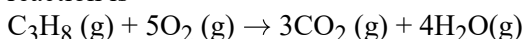
**78** At standard conditions, if the change in the enthalpy for the following reaction is  $-109 \text{ kJ mol}^{-1}$   
 $\text{H}_{2(g)} + \text{Br}_{2(g)} \rightarrow 2\text{HBr}_{(g)}$  and the bond energy of  $\text{H}_2$  and  $\text{Br}_2$  is  $435 \text{ kJ mol}^{-1}$  and  $192 \text{ kJ mol}^{-1}$ , respectively, what is the bond energy (in  $\text{kJ mol}^{-1}$ ) of  $\text{HBr}$ ?

1.	368	2.	736
3.	518	4.	259

**79** The enthalpies of elements in their standard states are taken as zero. The enthalpy of formation of a compound is-

1.	Generally negative	2.	Always positive
3.	Zero	4.	Never negative

**80** The standard heat of combustion of propane is  $-2220.1 \text{ kJ mol}^{-1}$ . The standard heat of vaporisation of liquid water is  $44.0 \text{ kJ mol}^{-1}$ . The enthalpy change for the reaction is-



1.	$-2220.1 \text{ kJ}$	2.	$-2044.1 \text{ kJ}$
3.	$-2396.1 \text{ kJ}$	4.	$-2176.1 \text{ kJ}$

## HESS'S LAW - LEVEL I

**81** The enthalpy of formation of  $\text{CO}_2(g)$ ,  $\text{H}_2\text{O}(l)$  and propene (g) are  $-393.5$ ,  $-285.8$  and  $20.42 \text{ kJ mol}^{-1}$  respectively. The enthalpy change for the combustion of cyclopropane at  $298 \text{ K}$  will be -  
 (The enthalpy of isomerisation of cyclopropane to propene is  $-33.0 \text{ kJ mol}^{-1}$ .)

1.	$-1021.32 \text{ kJ mol}^{-1}$	2.	$-2091.32 \text{ kJ mol}^{-1}$
3.	$-5021.32 \text{ kJ mol}^{-1}$	4.	$-3141.32 \text{ kJ mol}^{-1}$

**82** The enthalpies of formation of  $\text{Al}_2\text{O}_3$  and  $\text{Cr}_2\text{O}_3$  are  $-1596 \text{ kJ}$  and  $-1134 \text{ kJ}$ , respectively.  $\Delta H$  for the reaction  $2\text{Al} + \text{Cr}_2\text{O}_3 \rightarrow 2\text{Cr} + \text{Al}_2\text{O}_3$  is -

1.	$-1365 \text{ kJ}$	2.	$2730 \text{ kJ}$
3.	$-2730 \text{ kJ}$	4.	$-462 \text{ kJ}$

**83**  $4.8 \text{ g}$  of  $\text{C}(\text{diamond})$  on complete combustion evolves  $1584 \text{ kJ}$  of heat. The standard heat of the formation of gaseous carbon is  $725 \text{ kJ/mol}$ . The energy required

for the given process will be:

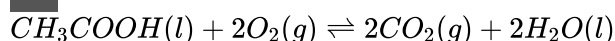
(i)  $\text{C}(\text{graphite}) \rightarrow \text{C}(\text{gas})$

(ii)  $\text{C}(\text{diamond}) \rightarrow \text{C}(\text{gas})$

1.	725, 727	2.	727, 725
3.	725, 723	4.	None of the above

## 2ND & 3RD LAW OF THERMODYNAMICS - LEVEL I

**84** For the reaction



at  $25^\circ\text{C}$  and  $1 \text{ atm}$ . pressure,  $\Delta H = -874 \text{ kJ}$ . The change in internal energy ( $\Delta E$ ) is -

- $-874 \text{ kJ}$
- $-971.53 \text{ kJ}$
- $+971.53 \text{ kJ}$
- $+874 \text{ kJ}$

**85** The maximum work done in expanding  $16 \text{ g}$   $\text{O}_2$  isothermally at  $300 \text{ K}$  and occupying a volume of  $5 \text{ dm}^3$  until the volume becomes  $25 \text{ dm}^3$  is-

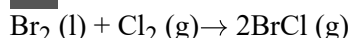
- $-2.01 \times 10^3 \text{ J}$
- $2.01 \times 10^{-3} \text{ J}$
- $+2.81 \times 10^3 \text{ J}$
- $+2.01 \times 10^{-6} \text{ J}$

## SPONTANEITY & ENTROPY - LEVEL I

**86** The unit of entropy is -

- $\text{J mol}^{-1}$
- $\text{JK mol}^{-1}$
- $\text{J mol}^{-1} \text{K}^{-1}$
- $\text{J}^{-1} \text{K}^{-1} \text{mol}^{-1}$

**87** The enthalpy and entropy change for the reaction :



are  $30 \text{ kJ mol}^{-1}$  and  $105 \text{ J K}^{-1} \text{mol}^{-1}$  respectively.

The temperature at which the reaction will be in equilibrium is :

- $285.7 \text{ K}$
- $273.4 \text{ K}$
- $450.9 \text{ K}$
- $300.1 \text{ K}$

**88** Which of the following is not correct?

1.  $\Delta G$  is zero for a reversible reaction.
2.  $\Delta G$  is positive for a spontaneous reaction.
3.  $\Delta G$  is negative for a spontaneous reaction.
4.  $\Delta G$  is positive for a non-spontaneous reaction.

**89**

<b>Assertion (A):</b>	A liquid crystallizes into a solid and is accompanied by a decrease in entropy.
<b>Reason (R):</b>	In crystals, molecules organize in an ordered manner.
1.	Both (A) and (R) are true and (R) is the correct explanation of (A).
2.	Both (A) and (R) are true but (R) is not the correct explanation of (A).
3.	(A) is true but (R) is false.
4.	(A) is false but (R) is true.

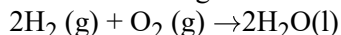
**90** Entropy decreases during:

1. Crystallization of sucrose from solution
2. Rusting of iron
3. Melting of ice
4. Vaporization of camphor

**91** The entropy change in the fusion of one mole of a solid melting at  $27^\circ\text{C}$  is -  
(the latent heat of fusion is  $2930\text{ J mol}^{-1}$ )

1.  $9.77\text{ JK}^{-1}\text{mol}^{-1}$
2.  $19.73\text{ JK}^{-1}\text{mol}^{-1}$
3.  $2930\text{ JK}^{-1}\text{mol}^{-1}$
4.  $108.5\text{ JK}^{-1}\text{mol}^{-1}$

**92** At a temperature of  $300\text{K}$ , what is the entropy change for the reaction given below?



Standard entropies of  $\text{H}_2(\text{g})$ ,  $\text{O}_2(\text{g})$  and  $\text{H}_2\text{O}(\text{l})$  are  $126.6$ ,  $201.20$  and  $68.0\text{ JK}^{-1}\text{mol}^{-1}$  respectively.

1.  $-318.4\text{ JK}^{-1}\text{mol}^{-1}$
2.  $318.4\text{ JK}^{-1}\text{mol}^{-1}$
3.  $31.84\text{ JK}^{-1}\text{mol}^{-1}$
4. None of the above

**93** Match the following process with entropy change

Reaction	Entropy change
A. A liquid vaporizes	1. $\Delta S = 0$
B. Reaction is non-spontaneous at all temperatures and $\Delta H$ is positive	2. $\Delta S = \text{positive}$
C. Reversible expansion of an ideal gas	3. $\Delta S = \text{negative}$

**Codes**

	A	B	C
1.	2	3	1
2.	1	2	3
3.	1	3	2
4.	3	2	1

**94** For the reaction,  $2\text{Cl}(\text{g}) \rightarrow \text{Cl}_2(\text{g})$ , the value of  $\Delta H$  and  $\Delta S$  are respectively -

1.  $\Delta H = 0$ ,  $\Delta S = -\text{ve}$
2.  $\Delta H = 0$ ,  $\Delta S = 0$
3.  $\Delta H = -\text{ve}$ ,  $\Delta S = -\text{ve}$
4.  $\Delta H = +\text{ve}$ ,  $\Delta S = +\text{ve}$

**95** Match the following parameters with a description for spontaneity.

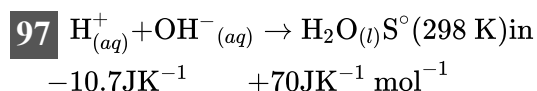
$\Delta$ (Parameters) $\Delta_r H^\ominus$ $\Delta_r S^\ominus \Delta_r G^\ominus$	Description
A. + - +	1. Non-spontaneous at high temperature
B. - - + at high T	2. Spontaneous at all temperatures
C. - + -	3. Non-spontaneous at all temperatures

**Codes**

	A	B	C
1.	3	1	2
2.	1	2	3
3.	1	3	2
4.	3	2	1

**96** Which of the following thermodynamic quantities is an outcome of the second law of thermodynamics?

1. Work
2. Enthalpy
3. Internal energy
4. Entropy



Standard entropy change for the above reaction is:

1.  $60.3 \text{ JK}^{-1} \text{ mol}^{-1}$
2.  $80.7 \text{ JK}^{-1} \text{ mol}^{-1}$
3.  $-70 \text{ JK}^{-1} \text{ mol}^{-1}$
4.  $+10.7 \text{ JK}^{-1} \text{ mol}^{-1}$

**98** Which of the following conditions will always lead to a non-spontaneous change?

1. Positive  $\Delta H$  and positive  $\Delta S$
2. Negative  $\Delta H$  and negative  $\Delta S$
3. Positive  $\Delta H$  and negative  $\Delta S$
4. Negative  $\Delta H$  and positive  $\Delta P$

**99** The entropy change in the isothermal reversible expansion of 2 moles of an ideal gas from 10 to 100 L at 300 K is

1.  $42.3 \text{ J K}^{-1}$
2.  $35.8 \text{ J K}^{-1}$
3.  $38.3 \text{ J K}^{-1}$
4.  $32.3 \text{ J K}^{-1}$

**100** Change in entropy is negative for:

1. Bromine (l)  $\rightarrow$  Bromine (g)
2.  $\text{C(s)} + \text{H}_2\text{O(g)} \rightarrow \text{CO(g)} + \text{H}_2\text{(g)}$
3.  $\text{N}_2\text{(g, 10 atm)} \rightarrow \text{N}_2\text{(g, 1 atm)}$
4.  $\text{Fe (1 mol, 400 K)} \rightarrow \text{Fe (1 mol, 300 K)}$

**101** The occurrence of a reaction is impossible if -

1.  $\Delta H$  is +ve ;  $\Delta S$  is also +ve but  $\Delta H < T\Delta S$
2.  $\Delta H$  is -ve ;  $\Delta S$  is also -ve but  $\Delta H > T\Delta S$
3.  $\Delta H$  is -ve ;  $\Delta S$  is +ve
4.  $\Delta H$  is +ve ;  $\Delta S$  is -ve

**102** For exothermic reaction to be spontaneous ( $\Delta S$  = negative) temperature must be -

1. High
2. Zero
3. Constant
4. Low

**103** 1 mole of an ideal gas at  $25^\circ\text{C}$  is subjected to expand reversibly ten times of its initial volume. The change in entropy of expansion is:

1.  $19.15 \text{ JK}^{-1} \text{ mol}^{-1}$
2.  $16.15 \text{ JK}^{-1} \text{ mol}^{-1}$
3.  $22.15 \text{ JK}^{-1} \text{ mol}^{-1}$
4. None of the above

**104** The molar entropy of the vapourization of acetic acid is  $14.4 \text{ cal K}^{-1} \text{ mol}^{-1}$  at its boiling point  $118^\circ\text{C}$ . The latent heat of vapourization of acetic acid is-

1.  $49 \text{ cal g}^{-1}$
2.  $64 \text{ cal g}^{-1}$
3.  $94 \text{ cal g}^{-1}$
4.  $84 \text{ cal g}^{-1}$

**105** Assertion: Dissolution of sugar in water proceeds via an increase in entropy.

Reason: Entropy decreases, when an egg is boiled hard.

1. Both assertion & reason are true and the reason is the correct explanation of the assertion.
2. Both assertion & reason are true but the reason is not the correct explanation of the assertion.
3. Assertion is a true statement but reason is false.
4. Both assertion and reason are false statements.

**106** For the conversion  $\text{C}_{(\text{graphite})} \rightarrow \text{C}_{(\text{diamond})}$  the  $\Delta S$  is:

1. Zero
2. Positive
3. Negative
4. Unknown

## SPONTANEITY & ENTROPY - LEVEL

### II

**107** Which of the following statements is correct for the spontaneous adsorption of a gas?

1.	$\Delta S$ is negative and therefore, $\Delta H$ should be highly positive
2.	$\Delta S$ is negative and therefore, $\Delta H$ should be highly negative
3.	$\Delta S$ is positive and therefore, $\Delta H$ should be negative
4.	$-\Delta S$ is positive and therefore, $\Delta H$ should also be highly positive

**108** For the reaction,  $2\text{Cl(g)} \rightarrow \text{Cl}_2\text{(g)}$ , the correct option is:

1.  $\Delta_r H > 0$  and  $\Delta_r S < 0$
2.  $\Delta_r H < 0$  and  $\Delta_r S > 0$
3.  $\Delta_r H < 0$  and  $\Delta_r S < 0$
4.  $\Delta_r H > 0$  and  $\Delta_r S > 0$

**109** The entropy change in the conversion of one mole of liquid water at 373 K to vapour at the same temperature would be: (Latent heat of vaporization of water,  $\Delta H_{\text{vap}} = 2.257 \text{ kJ/g}$ )

1.  $105.9 \text{ JK}^{-1}\text{mol}^{-1}$
2.  $107.9 \text{ JK}^{-1}\text{mol}^{-1}$
3.  $108.9 \text{ JK}^{-1}\text{mol}^{-1}$
4.  $109.9 \text{ JK}^{-1}\text{mol}^{-1}$

**110** The entropy change can be calculated by using the expression  $\Delta S = \frac{q_{\text{rev}}}{T}$ . When water freezes in a glass beaker, the correct statement among the following is:

1.	$\Delta S$ (system) decreases but $\Delta S$ (surroundings) remains the same
2.	$\Delta S$ (system) increases but $\Delta S$ (surroundings) decreases
3.	$\Delta S$ (system) decreases but $\Delta S$ (surroundings) increases
4.	$\Delta S$ (system) decreases but $\Delta S$ (surroundings) also decreases

**111** For  $A \rightarrow B$ ,  $\Delta H = 4 \text{ kcal mol}^{-1}$ ,  $\Delta S = 10 \text{ cal mol}^{-1} \text{ K}^{-1}$ , the reaction is spontaneous when the temperature is:

1. 400 K
2. 300 K
3. 500 K
4. None of the above

**112**

<b>Assertion (A):</b>	Spontaneous process is an irreversible process and may be reversed by some external agency.
<b>Reason (R):</b>	A decrease in enthalpy is a contributing factor to spontaneity.

1.	Both (A) and (R) are true and (R) is the correct explanation of (A).
2.	Both (A) and (R) are true but (R) is not the correct explanation of (A).
3.	(A) is true but (R) is false.
4.	(A) is false but (R) is true.

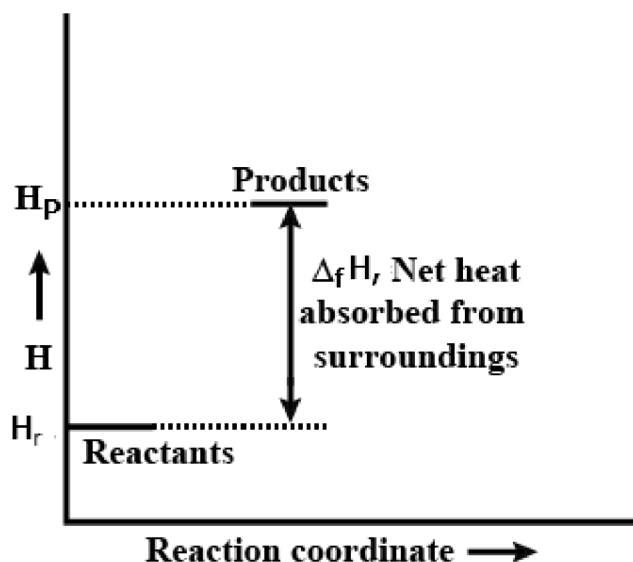
**113** Entropy decreases in which of the following reactions?

1. Boiling of egg
2. Combustion of benzene at  $127^\circ\text{C}$
3. Stretching of rubber band
4. Dissolution of sugar in water

**114** For the reaction  $A + B \rightarrow C + D + q \text{ (kJ/mol)}$ , entropy change is positive. The reaction will be

1. Possible only at high temperature
2. Possible only at low temperature
3. Not possible at any temperature
4. Possible at any temperature

**115** An enthalpy diagram for a particular reaction is given below:



The correct statement among the following is-

1. Reaction is spontaneous
2. Reaction is non-spontaneous
3. Cannot predict spontaneity of the reaction from the graph given above
4. None of the above

**116** The entropy change in the surroundings when 1.00 mol of  $\text{H}_2\text{O(l)}$  is formed under standard conditions is-

$$\Delta_f H^\circ = -286 \text{ kJ mol}^{-1}$$

1.  $952.5 \text{ J mol}^{-1}$
2.  $979.7 \text{ J mol}^{-1}$
3.  $949.7 \text{ J mol}^{-1}$
4.  $959.7 \text{ J mol}^{-1}$

## GIBBS ENERGY CHANGE - LEVEL I

**117** The values of  $\Delta H$  and  $\Delta S$  for the reaction,  $\text{C}_{(\text{graphite})} + \text{CO}_2(\text{g}) \rightarrow 2\text{CO}(\text{g})$  are 170 kJ and  $170 \text{ JK}^{-1}$ , respectively. This reaction will be spontaneous at:

1. 710 K
2. 910 K
3. 1110 K
4. 510 K

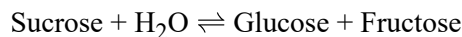
**118** The free energy change due to a reaction is zero when-

1. The reactants are initially mixed.
2. A catalyst is added
3. The system is at equilibrium
4. The reactants are completely consumed

**119** Which of the following statements is correct for a reversible process in a state of equilibrium?

1.  $\Delta G = -2.30RT \log K$
2.  $\Delta G = 2.30RT \log K$
3.  $\Delta G^\circ = -2.30RT \log K$
4.  $\Delta G^\circ = 2.30RT \log K$

**120** Hydrolysis of sucrose is given by the following reaction



If the equilibrium constant ( $K_c$ ) is  $2 \times 10^{13}$  at 300 K, the value of  $\Delta_r G^\circ$  at the same temperature will be:

1.  $8.314 \text{ J mol}^{-1} \text{ K}^{-1} \times 300 \text{ K} \times \ln(2 \times 10^{13})$
2.  $8.314 \text{ J mol}^{-1} \text{ K}^{-1} \times 300 \text{ K} \times \ln(3 \times 10^{13})$
3.  $-8.314 \text{ J mol}^{-1} \text{ K}^{-1} \times 300 \text{ K} \times \ln(4 \times 10^{13})$
4.  $-8.314 \text{ J mol}^{-1} \text{ K}^{-1} \times 300 \text{ K} \times \ln(2 \times 10^{13})$

**121** For a given reaction,  $\Delta H = 35.5 \text{ kJ mol}^{-1}$  and  $\Delta S = 83.6 \text{ JK}^{-1} \text{ mol}^{-1}$ . The reaction is spontaneous at: (Assume that  $\Delta H$  and  $\Delta S$  do not vary with temperature)

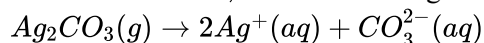
1.  $T < 425 \text{ K}$
2.  $T > 425 \text{ K}$
3. All temperatures
4.  $T > 298 \text{ K}$

**122** The equilibrium constant for a reaction is 10. The value of  $\Delta G^\circ$  will be-

$$(R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}; T = 300 \text{ K})$$

1.  $-5.74 \text{ kJ mol}^{-1}$
2.  $-5.74 \text{ J mol}^{-1}$
3.  $+4.57 \text{ kJ mol}^{-1}$
4.  $-57.4 \text{ kJ mol}^{-1}$

**123** Given the Gibbs free energy change,  $\Delta G^\circ = +63.3 \text{ kJ}$ , for the following reaction,



$K_{sp}$  of  $\text{Ag}_2\text{CO}_3(\text{s})$  in water at  $25^\circ \text{C}$  is ( $R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}$ )

1.  $3.2 \times 10^{26}$
2.  $8.0 \times 10^{-12}$
3.  $2.9 \times 10^{-3}$
4.  $7.9 \times 10^{-2}$

**124** The free energy change ( $\Delta G^\circ$ ) is negative when -

1. The surroundings do no electrical work on the system.
2. The surroundings do electrical work on the system.
3. The system does electrical work on the surroundings.
4. The system does no electrical work on the surroundings.

**125**

The standard free energies of formation (in kJ/mol) at 298 K are -237.2, -394.4, and -8.2 for  $\text{H}_2\text{O(l)}$ ,  $\text{CO}_2(\text{g})$ , and pentane (g), respectively. The value of  $E_{\text{cell}}$  for the pentane-oxygen fuel cell is:

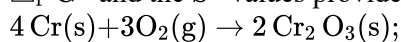
1. 1.968 V
2. 2.0968 V
3. 1.0968 V
4. 0.0968 V

## GIBBS ENERGY CHANGE - LEVEL II

**126** In the reaction,  $\Delta H$  and  $\Delta S$  both are positive. The condition under which the reaction would not be spontaneous is -

1.  $\Delta H > T\Delta S$
2.  $\Delta S = \Delta H/T$
3.  $\Delta H = T\Delta S$
4. All of the above

**127** Calculate  $\Delta_f H^\circ$  (in kJ/mol) for  $\text{Cr}_2\text{O}_3$  from the  $\Delta_r G^\circ$  and the  $S^\circ$  values provided at  $27^\circ\text{C}$



$$\Delta_r G^\circ = -2093.4 \text{ kJ/mol}$$

$$S^\circ (\text{J/K}) \text{ mol} : S^\circ (\text{Cr, s}) = 24;$$

$$S^\circ (\text{O}_2, \text{g}) = 205; S^\circ (\text{Cr}_2\text{O}_3, \text{s}) = 81$$

1. -2258.1 kJ/mol
2. -1129.05 kJ/mol
3. -964.35 kJ/mol
4. None of the above

**128** For the reaction  $2\text{A(g)} + \text{B(g)} \rightarrow 2\text{D(g)}$ ;  $\Delta U^\circ = -10.5 \text{ kJ}$  and  $\Delta S^\circ = -44.1 \text{ J K}^{-1}$ , the value of  $\Delta G^\circ$  for the given reaction would be-

1. 1.6 J
2. -0.16 kJ
3. 0.16 kJ
4. 1.6 kJ

**129** Equilibrium is represented by -

1.  $\Delta H = 0$
2.  $\Delta G_{\text{Total}} = 0$
3.  $\Delta S_{\text{Total}} = 0$
4.  $\Delta E = 0$