

6.THERMODYNAMICS

Single Correct Answer Type

1. Energy can transfer from system to surroundings as work if
 - a) There is thermal equilibrium between system and surrounding
 - b) There is mechanical equilibrium between system and surrounding
 - c) If pressure of system > atmospheric pressure
 - d) None of these
2. Temperature of 1 mol of a gas is increased by 1° at constant pressure. The work done is
 - a) R
 - b) $2R$
 - c) $R/2$
 - d) $3R$
3. Carnot's cycle is said to have 25% efficiency when it operates between T (source) and 300 K (sink). Temperature T is
 - a) 300 K
 - b) 350 K
 - c) 375 K
 - d) 400 K
4. Since the enthalpy of elements in their natural state is taken to be zero, the heat of formation ($\Delta_f H$) of compounds
 - a) Is always negative
 - b) Is always positive
 - c) May be negative or positive
 - d) Is zero
5. The dissolution of NH_4Cl in water is endothermic even though NH_4Cl dissolves in water spontaneously. Which one of the following best explains this behaviour?
 - a) The bonds on solid NH_4Cl are weak
 - b) The entropy-driving force causes dissolution
 - c) Endothermic processes are energetically favourable
 - d) The dissolving process is unrelated to energy
6. Under which of the following condition is the relation $\Delta H = \Delta U + P \Delta V$ valid for a closed system at
 - a) Constant pressure
 - b) Constant temperature
 - c) Constant temperature and pressure
 - d) Constant temperature, pressure and composition
7. When a reaction is carried out in a closed vessel
 - a) $q_P < q_V$
 - b) $q_P > q_V$
 - c) $q_P = q_V$
 - d) $q_V = 0$
8. The products of combustion of an aliphatic thiol (RSH) at 298 K are
 - a) $\text{CO}_2(\text{g}), \text{H}_2\text{O}(\text{g}),$ and $\text{SO}_2(\text{g})$
 - b) $\text{CO}_2(\text{g}), \text{H}_2\text{O}(\text{l}),$ and $\text{SO}_2(\text{g})$
 - c) $\text{CO}_2(\text{l}), \text{H}_2\text{O}(\text{l}),$ and $\text{SO}_2(\text{g})$
 - d) $\text{CO}_2(\text{g}), \text{H}_2\text{O}(\text{l}),$ and $\text{SO}_2(\text{l})$
9. For an endothermic reaction where ΔH represents the enthalpy of the reaction in kJ mol^{-1} , the minimum value for the energy of activation will be
 - a) Less than ΔH
 - b) Zero
 - c) More than ΔH
 - d) Equal to ΔH
10. For the gaseous reaction involving the complete combustion of isobutene
 - a) $\Delta H = \Delta U$
 - b) $\Delta H > \Delta U$
 - c) $\Delta H = \Delta U = 0$
 - d) $\Delta H < \Delta U$
11. If $\text{S} + \text{O}_2 \rightarrow \text{SO}_2; \Delta H = -298.2 \text{ kJ}$
 $\text{SO}_2 + 1/2 \text{O}_2 \rightarrow \text{SO}_3; \Delta H = -98.7 \text{ kJ}$
 $\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4; \Delta H = -130.2 \text{ kJ}$
 $\text{H}_2 + 1/2 \text{O}_2 \rightarrow \text{H}_2\text{O}; \Delta H = -287.3 \text{ kJ}$
 Then the enthalpy of formation of H_2SO_4 at 298 K is
 - a) -814.4 kJ
 - b) -650.3 kJ
 - c) -320.5 kJ
 - d) -233.5 kJ
12. The products of combustion of an aliphatic thiol (RSH) at 298 K are
 - a) $\text{CO}_2(\text{g}), \text{H}_2\text{O}(\text{g}),$ and $\text{SO}_2(\text{g})$
 - b) $\text{CO}_2(\text{g}), \text{H}_2\text{O}(\text{l}),$ and $\text{SO}_2(\text{g})$
 - c) $\text{CO}_2(\text{l}), \text{H}_2\text{O}(\text{l}),$ and $\text{SO}_2(\text{g})$
 - d) $\text{CO}_2(\text{g}), \text{H}_2\text{O}(\text{l}),$ and $\text{SO}_2(\text{l})$
13. For hypothetical reversible reaction
 $1/2 \text{A}_2(\text{g}) + 3/2 \text{B}_2(\text{g}) \rightarrow \text{AB}_3(\text{g}); \Delta H = -20 \text{ kJ}$ if standard entropies of A_2, B_2 and AB_3 are 60, 40, and

- 50J K⁻¹mol⁻¹, respectively. The above reaction will be equilibrium at
 a) 400 K b) 500 K c) 250 K d) 200 K
14. $C_p - C_v = R$. This R is
 a) Change in KE
 b) Change in rotational energy
 c) Work done which system can do on expanding the gas per mol per degree increase in temperature
 d) All correct
15. For which of the following equations, will ΔH be equal to ΔU ?
 a) $H_2(g) + 1/2 O_2(g) \rightarrow H_2O(l)$ b) $H_2(g) + I_2(g) \rightarrow 2HI(g)$
 c) $2NO_2(g) \rightarrow N_2O_4(g)$ d) $4NO_2(g) + O_2(g) \rightarrow 2N_2O_5(g)$
16. A reaction, $A + B \rightarrow C + D + q$, is found to have a positive entropy change, the reaction will be:
 a) Possible at high temperature b) Possible only at low temperature
 c) Not possible at any temperature d) Possible at any temperature
17. For the combustion reaction at 298 K
 $2Ag(s) + 1/2 O_2(g) \rightarrow 2Ag_2O(s)$
 Which of the following alternatives is correct?
 a) $\Delta H = \Delta U$ b) $\Delta H > \Delta U$
 c) $\Delta H < \Delta U$ d) ΔH and ΔU bear no relation with each other
18. Enthalpy of the system is given as
 a) $H + PV$ b) $U + PV$ c) $U - PV$ d) $H - PV$
19. In thermodynamics, a process is called reversible when
 a) The surroundings and system change into each other
 b) There is no boundary between the system and surroundings
 c) The surroundings are always in equilibrium with the system
 d) The system changes into the surroundings spontaneously
20. For the reversible process, the value of ΔS is given by the expression
 a) $\Delta H/\Delta T$ b) $T/q(\text{rev})$ c) $q(\text{rev}) \times T$ d) $q(\text{rev})/T$
21. Which is not intensive property?
 a) Boiling point b) Refractive index c) Molarity d) Volume
22. Which one of the following statements is false?
 a) Work is a state function
 b) Temperature is a state function
 c) Change in the state is completely defined when the initial and final states are specified
 d) Work appears at the boundary of the system
23. The $\Delta_f H^\ominus$ for $CO_2(g)$, $CO(g)$, and $H_2O(g)$ are -393.5 , -110.5 , and -214.8 kJ mol⁻¹, respectively. The standard enthalpy change (in kJ mol⁻¹) for the reaction
 $CO_2(g) + H_2(g) \rightarrow CO(g) + H_2O(g)$ is
 a) 524.1 b) +41.2 c) -262.5 d) -41.2
24. If $\Delta_f H^\ominus$ of $ICl(g)$, $Cl(g)$, and $I(g)$ is 17.57, 121.34 and 106.96 J mol⁻¹ respectively. Then bond dissociation energy of I - Cl bond is
 a) 35.15 J mol⁻¹ b) 106.69 mol⁻¹ c) 210.73 J mol⁻¹ d) 420.9 J mol⁻¹
25. The difference between the heats of reaction at constant pressure and constant volume for the reaction
 $2C_6H_6(l) + 15O_2 \rightarrow 12CO_2(g) + 6H_2O(l)$ at 25°C in kJ is
 a) -7.43 b) +3.72 c) -3.72 d) +7.43
26. The heat of hydrogenation of ethane is x_1 and that of benzene is x_2
 Hence resonance energy of benzene is
 a) $x_1 - x_2$ b) $x_1 + x_2$ c) $3x_1 - x_2$ d) $x_1 - 3x_2$
27. The relationship between the free energy change (ΔG) and entropy change (ΔS) at constant temperature (T) is
 a) $\Delta G = \Delta H - T \Delta S$ b) $\Delta H = \Delta G + T \Delta S$ c) $T \Delta S = \Delta G + \Delta H$ d) $\Delta G = -\Delta H - T \Delta S$

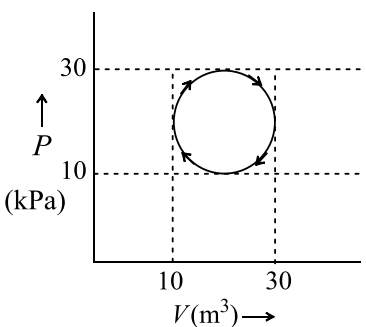
28. Molar heat capacity of water in equilibrium with ice at constant pressure is
 a) Zero b) Infinity (∞) c) $40.45 \text{ kJ K}^{-1}\text{mol}^{-1}$ d) $75.48 \text{ J K}^{-1}\text{mol}^{-1}$
29. $\text{H}_2 + 1/2 \text{O}_2 \rightarrow \text{H}_2\text{O}$; $\Delta H^\ominus = -68 \text{ kcal}$
 $\text{K} + \text{H}_2\text{O} + \text{aq} \rightarrow \text{KOH}(\text{aq}) + 1/2 \text{H}_2$; $\Delta H^\ominus = -48 \text{ kcal}$
 $\text{KOH} + \text{aq} \rightarrow \text{KOH}(\text{aq})$; $\Delta H^\ominus = -14 \text{ kcal}$
 From the above data, the standard heat of formation of KOH in kcal is
 a) $-68 + 48 - 14$ b) $-68 - 48 + 14$ c) $68 - 48 + 14$ d) $68 + 48 + 14$
30. If a gas absorbs 200 J of heat and expands by 500 cm^3 against a constant pressure of $2 \times 10^5 \text{ N m}^{-2}$, then the change in internal energy is
 a) -300 J b) -100 J c) $+100 \text{ J}$ d) $+300 \text{ J}$
31. Which of the following reaction is endothermic?
 a) $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ b) $\text{Fe} + \text{S} \rightarrow \text{FeS}$
 c) $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$ d) $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$
32. Entropy of system depends upon
 a) Volume only b) Temperature only
 c) Pressure only d) Pressure, volume, and temperature
33. When one mole of monoatomic ideal gas at $T \text{ K}$ undergoes adiabatic change under a constant external pressure of 1 atm changes volume from 1 L to 2 L . The final temperature in Kelvin would be
 a) $\frac{T}{2^{2/3}}$ b) $T + \frac{2}{3 \times 0.0821}$ c) T d) $T - \frac{2}{3 \times 0.0821}$
34. $\text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$
 $\text{BE}(\text{H}-\text{H}) = x_1$; $\text{BE}(\text{O}=\text{O}) = x_2$
 $\text{BE}(\text{O}-\text{H}) = x_3$
 Latent heat of vaporization of water liquid into water vapour = x_4 , then $\Delta_f H$ (heat of formation of liquid water) is
 a) $x_1 + \frac{x_2}{2} - x_3 + x_4$ b) $2x_3 - x_1 - \frac{x_2}{2} - x_4$ c) $x_1 + \frac{x_2}{2} - 2x_3 - x_4$ d) $x_1 + \frac{x_2}{2} - 2x_3 + x_4$
35. Evaporation of water is
 a) An exothermic change
 b) An endothermic change
 c) A process where no heat changes occur
 d) A process accompanied by chemical reaction
36. In thermodynamics, a process is called reversible when
 a) The surroundings and the system change into each other
 b) There is no boundary between the system and the surroundings
 c) The surroundings are always in equilibrium with the system
 d) The system changes into the surroundings spontaneously
37. Which of the following equations corresponds to the enthalpy of combustion at 298 K ?
 a) $\text{C}_2\text{H}_6(\text{g}) + 7/2 \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{g})$ b) $2\text{C}_2\text{H}_6(\text{g}) + 7\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{g})$
 c) $\text{C}_2\text{H}_6(\text{g}) + 7/2 \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$ d) $2\text{C}_2\text{H}_6(\text{g}) + 7\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l})$
38. If a certain mass of gas is made to undergo separately adiabatic and isothermal expansions to the same pressure, starting from the same initial conditions of temperature and pressure, then, as compared to that of isothermal expansion, in the case of adiabatic expansion, the final
 a) Volume and temperature will be higher
 b) Volume and temperature will be lower
 c) Temperature will be lower but the final volume will be higher
 d) Volume will be lower but the final volume will be higher
39. Which of the following is an endothermic reaction?
 a) $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ b) $\text{N}_2 + \text{O}_2 \rightarrow 2\text{NO}$

- c) $2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$ d) $\text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}$
40. The expression $\Delta_{\text{subl}}H^\ominus = \Delta_{\text{fus}}H^\ominus + \Delta_{\text{vap}}H^\ominus$ is true at all
a) Temperatures b) Pressures
c) Temperatures and pressures d) Temperatures and 1 atm pressure conditions
41. 1 mol of NH_3 gas at 27°C is expanded under adiabatic condition to make volume 8 times ($\gamma = 1.33$). Final temperature and work done, respectively, are
a) 150 K, 900 cal b) 150 K, 400 cal c) 250 K, 1000 cal d) 200 K, 800 cal
42. Which of the following equations corresponds to the definition of enthalpy of formation at 298 K?
a) $\text{C}(\text{graphite}) + 2\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{l}) \rightarrow \text{CH}_3\text{OH}(\text{g})$ b) $\text{C}(\text{diamond}) + 2\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CH}_3\text{OH}(\text{l})$
c) $2\text{C}(\text{graphite}) + 4\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CH}_3\text{OH}(\text{l})$ d) $\text{C}(\text{graphite}) + 2\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CH}_3\text{OH}(\text{l})$
43. Heat of neutralization of CsOH with all strong acid is $13.4 \text{ kcal mol}^{-1}$. The heat released on neutralization of CsOH with HF (weak acid) is $16.4 \text{ kcal mol}^{-1}$. ΔH^\ominus of ionization of HF is
a) 3.0 kcal b) -3.0 kcal c) 6.0 kcal d) 0.3 kcal
44. Inversion temperature is
a) $\frac{Rb}{2a}$ b) $\frac{2a}{Rb}$ c) $\frac{Rb}{a}$ d) $\frac{a}{Rb}$
45. $\Delta_f H(\text{H}_2\text{O}) = -68 \text{ kcal mol}^{-1}$ and ΔH of neutralization is $-13.7 \text{ kcal mol}^{-1}$, then the heat of formation of OH^\ominus is
a) $-68 \text{ kcal mol}^{-1}$ b) $-54.3 \text{ kcal mol}^{-1}$ c) $54.3 \text{ kcal mol}^{-1}$ d) $-71.7 \text{ kcal mol}^{-1}$
46. For the gaseous reaction: $\text{N}_2\text{O}_4 \rightarrow 2\text{NO}_2$
a) $\Delta H > \Delta U$ b) $\Delta H < \Delta U$ c) $\Delta H = \Delta U$ d) $\Delta H = 0$
47. $\text{A}(\text{l}) \xrightleftharpoons{1 \text{ atm}} \text{A}(\text{g})$, $\Delta_{\text{vap}}H = 460.6 \text{ cal mol}^{-1}$,
boiling point = 50 K. What is the boiling point at 10 atm
a) 150 K b) 75 K c) 100 K d) None is correct
48. The heat of neutralization of oxalic acid is $-25.4 \text{ kcal mol}^{-1}$ using strong base, NaOH . Hence, the enthalpy change of the process is
 $\text{H}_2\text{C}_2\text{O}_4 \rightleftharpoons 2\text{H}^\oplus + \text{C}_2\text{O}_4^{2-}$ is
a) 2.0 kcal b) -11.8 kcal c) 1.0 kcal d) -1.0 kcal
49. For the process, $\text{H}_2\text{O}(\text{s}) \rightarrow \text{H}_2\text{O}(\text{l})$ at 273 K
a) $G(\text{ice}) = G(\text{water}) = 0$ b) $G(\text{ice}) = G(\text{water}) \neq 0$
c) $G(\text{ice}) > G(\text{water})$ d) $G(\text{ice}) < G(\text{water})$
50. For an endothermic reaction where ΔH represents the enthalpy of the reaction in kJ mol^{-1} , the minimum value for the energy of the activation will be
a) Less than ΔH b) Zero c) More than ΔH d) Equal to ΔH
51. Standard molar enthalpy of formation of CO_2 is equal to
a) Zero
b) The standard molar enthalpy of combustion of gaseous carbon
c) The sum of standard molar enthalpies of formation of CO_2 and O_2
d) The standard molar enthalpy of combustion of carbon (graphite)
52. The molar enthalpies of combustion of $\text{C}_2\text{H}_2(\text{g})$, $\text{C}(\text{graphite})$ and $\text{H}_2(\text{g})$ are -1300 , -394 , and -286 kJ mol^{-1} , respectively. The standard enthalpy of formation of $\text{C}_2\text{H}_2(\text{g})$ is
a) -226 kJ mol^{-1} b) -626 kJ mol^{-1} c) 226 kJ mol^{-1} d) 626 kJ mol^{-1}
53. The standard heat of combustion of Al is $-837.8 \text{ kJ mol}^{-1}$ at 25°C . If Al reacts with O_2 at 25°C , which of the following release 250 kJ of heat?
a) The reaction of 0.624 mol of Al b) The formation of 0.624 mol of Al_2O_3
c) The reaction of 0.312 mol of Al d) The formation of 0.150 mol of Al_2O_3
54. The enthalpy of formation of hypothetical MgCl is -125 kJ mol^{-1} and for MgCl_2 is -642 kJ mol^{-1} . What is the enthalpy of the disproportionation of MgCl
a) 392 kJ mol^{-1} b) -392 kJ mol^{-1} c) -767 kJ mol^{-1} d) -517 kJ mol^{-1}

55. At equilibrium state
 a) $\Delta_{\text{total}}S > 0$ b) $\Delta_{\text{total}}S < 0$ c) $\Delta_{\text{total}}S = 0$ d) Unpredictable
56. The relationship between enthalpy and internal energy change is
 a) $\Delta U = \Delta H + P \Delta V$ b) $\Delta H = \Delta U + P \Delta V$ c) $\Delta H = \Delta U - P \Delta V$ d) $P \Delta V = \Delta U + \Delta H$
57. Enthalpy change of a reaction will be equal to
 a) $\Delta U + P \Delta V$ b) $\Delta U + V \Delta P$ c) $\Delta U + \Delta(PV)$ d) $\Delta U + (\Delta n_g) \Delta(PV)$
58. For the given reactions
 $\text{SiO}_2 + 4\text{HF} \rightarrow \text{SiF}_4 + 2\text{H}_2\text{O}, \Delta H = -10.17 \text{ kcal}$
 $\text{SiO}_2 + 4\text{HCl} \rightarrow \text{SiCl}_4 + 2\text{H}_2\text{O}, \Delta H = 36.7 \text{ kcal}$
 It may be concluded that
 a) HF will attack SiO_2 and HCl will not b) HCl will attack SiO_2 and HF will not
 c) HF and HCl both attack SiO_2 d) None attack SiO_2
59. The word 'standard' in standard molar enthalpy change implies
 a) Temperature 298 K b) Pressure 1 atm
 c) Temperature 298 K and pressure 1 atm d) All temperatures and all pressures
60. Which law of thermodynamic, introduces the concept of entropy
 a) First law b) Zeroth law c) Third law d) Second law
61. In which of the following process ΔH and ΔU are of same magnitude
 a) Evaporation of $\text{CCl}_4(l)$ b) $\text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g)$
 c) $\text{NH}_4\text{Cl}(s) \rightarrow \text{NH}_4\text{Cl}(g)$ d) $2\text{HI}(g) \rightarrow \text{H}_2(g) + \text{I}_2(g)$
62. 1 g H_2 gas STP is expanded so that the volume is doubled. Hence, work done is
 a) 22.4 L-atm b) 5.6 L-atm c) 11.2 L-atm d) 44.8 L-atm
63. For the equations
 $\text{C}(\text{diamond}) + 2\text{H}_2(g) \rightarrow \text{CH}_4(g) \quad \Delta H_1$
 $\text{C}(g) + 4\text{H}(g) \rightarrow \text{CH}_4(g) \quad \Delta H_2$
 Predict whether
 a) $\Delta H_1 = \Delta H_2$ b) $\Delta H_1 > \Delta H_2$
 c) $\Delta H_1 < \Delta H_2$ d) $\Delta H_1 = \Delta H_2 + \Delta_{\text{vap}}H(\text{C}) + \Delta_{\text{diss}}H(\text{H}_2)$
64. For the reaction
 $\text{X}_2\text{O}_4(l) \rightarrow 2\text{XO}_2(g)$
 $\Delta U = 2.1 \text{ kcal}, \Delta S = 20 \text{ cal K}^{-1}$ at 300 K
 Hence, ΔG is
 a) 2.7 kcal b) -2.7 kcal c) 9.3 kcal d) -9.3 kcal
65. If $\Delta G = \Delta H - T \Delta S$ and $\Delta G = \Delta H + T \left[\frac{d(\Delta G)}{dT} \right]_P$ then variation of EMF of a cell E , with temperature T , is given by:
 a) $\frac{\Delta H}{nF}$ b) $\frac{\Delta G}{nF}$ c) $\frac{\Delta S}{nF}$ d) $-\frac{\Delta S}{nF}$
66. For an ideal gas Joule-Thomson coefficient is
 a) Positive b) Negative
 c) Zero d) Dependent on molecular weight

Multiple Correct Answers Type

67. Which of the following type of energies are involved in Born Haber's cycle?
 a) $\Delta_{\text{sub}}H$ b) Ionization energy
 c) Bond dissociation energy d) Lattice energy
68. The correct expressions for an adiabatic process are
 a) $\frac{T_2}{T_1} = \left(\frac{V_1}{V_2} \right)^{\gamma-1}$ b) $\frac{P_2}{P_1} = \left(\frac{T_1}{T_2} \right)^{\frac{\gamma-1}{\gamma}}$ c) $P_1 V_1^\gamma = P_2 V_2^\gamma$ d) $P_1 V_1^{\gamma-1} = P_2 V_2^{\gamma-1}$

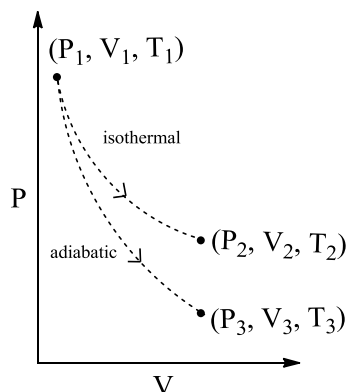
69. The heat of combustion of ethanol was determined in a bomb calorimeter and was found to be $-670.48 \text{ kcal mol}^{-1}$ at 25°C . What will be ΔU for the same reaction at 298K ?
 a) -335.24 kcal b) -669.28 cal c) -670.48 kcal d) $-280.26 \times 10^4 \text{ J}$
70. In which of the following cases entropy increases?
 a) Solid changing to liquid
 b) Expansion of a gas
 c) Crystals dissolve
 d) Boiling of an egg
71. Which of the following statements is/are correct?
 a) The entropy of the universe decreases and increases at a periodic rate
 b) The entropy of the universe increases and tends towards the maximum value
 c) For endothermic spontaneous processes the total entropy change decreases
 d) The entropy of the universe decreases and tends to zero
72. Which of the following are extensive properties?
 a) Elevation in boiling point b) Boiling point
 c) Emf of cell d) E^\ominus of cell
73. The heat of neutralization of a strong acid by a strong base is a constant because :
 a) The strong acid and strong base react completely
 b) The strong acid and strong base dissociate completely and only H^+ and OH^- ions react in every case
 c) The salt formed do not hydrolyse
 d) There is no side reaction during neutralization
74. In the isothermal expansion of an ideal gas:
 a) $\Delta U = 0$ b) $\Delta T = 0$ c) $q = 0$ d) $W = -q$
75. Which one is not correct for a cyclic process as shown in the figure?
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- a) $dU = 0$ b) $q = -w$ c) 314 J d) 31.4 J
76. Identify the intensive quantities from the following
 a) Enthalpy b) Temperature c) Volume d) Refractive index
77. Which of the following affect the heat of reaction?
 a) Physical states of reactants and products
 b) Allotropic forms of elements
 c) Temperature
 d) Reaction carried out at constant pressure or constant volume
78. In which of the following reactions, $\Delta H > \Delta U$?
 a) $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightarrow 2\text{HI}(\text{g})$ b) $\text{PCl}_5(\text{g}) \rightarrow \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$
 c) $2\text{H}_2\text{O}_2(\text{l}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$ d) $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$
79. The criteria for spontaneity of a process is/are
 a) $(dG)_{\text{TP}} < 0$ b) $(dE)_{\text{SV}} < 0$ c) $(dH)_{\text{SP}} < 0$ d) $(dS)_{\text{EV}} < 0$
80. Identify the intensive quantities from the following
 a) Enthalpy b) Temperature c) Volume d) Refractive index
81. For an ideal gas undergoing isothermal irreversible expansion
 a) $\Delta U = 0$ b) $\Delta H = 0$ c) $\Delta S = 0$ d) $w = 0$
82. The enthalpy of formation of $\text{UF}(\text{g})$ is 22 kcal mol^{-1} and that of $\text{U}(\text{g})$ is $128 \text{ kcal mol}^{-1}$. The bond energy of

- the F—F bond is 37.0 kcal mol⁻¹. The bond dissociation energy of UF(g) is/are :
- a) 124.5 kcal mol⁻¹ b) 131.1 kcal mol⁻¹ c) 521 kcal mol⁻¹ d) 623 kcal mol⁻¹
83. The correct relationship is/are :
- Kirchhoff's equation : Kirchhoff's equation :
- a) $\frac{d}{dT}(\Delta H) = \Delta C_P$ b) $\frac{\Delta H_2 - \Delta H_1}{\Delta T} = \Delta C_P$ c) $\Delta H = \Delta U + \Delta nRT$ d) $\frac{q_{\text{rev}}}{T} = \Delta S$
84. Which is not correct relationship?
- a) $\left[\frac{dH}{dT}\right]_P - \left[\frac{dU}{dT}\right]_T = (+ve)$ b) $\left[\frac{dU}{dV}\right]_T = 0$ (for ideal gas)
- c) $\left[\frac{dV}{dT}\right]_P = \frac{nR}{P}$ (for ideal gas) d) All of these
85. For gaseous reactions, if ΔH is the change in enthalpy and ΔU that in internal energy, then
- a) ΔH is always greater than ΔU
- b) ΔH is always less than ΔU
- c) $\Delta H < \Delta U$ only if the number of mole of the products is less than that of the reactants
- d) $\Delta U < \Delta H$ only if the number of mole of the reactants is less than that of the products
86. The intensive property/properties is/are
- a) Temperature b) Pressure c) Internal energy d) Heat capacity
87. Select the correct statement
- a) There is a natural asymmetry between converting work to heat and converting heat to work
- b) No process is possible in which the sole result is the absorption of heat from a reservoir and its complete conversion into work
- c) For every chemical reaction at equilibrium, standard change in Gibbs free energy is zero
- d) At constant temperature and pressure, chemical reactions are spontaneous in the direction of decreasing Gibbs energy
88. Three identical adiabatic containers have helium, neon and oxygen gases at the same pressure. The gases are compressed to half their original volume. Under these conditions
- a) The final temperature of both helium and neon is same
- b) The final pressure of the gas in each container is same
- c) The final temperature of the gas in each container is same
- d) The final pressure of both helium and neon is same
89. Which is intensive property?
- a) Mass b) Mass/volume c) Volume d) Volume/mass
90. In which of the following entropy increases?
- a) Rusting of iron b) Melting of ice
- c) Crystallization of sugar from solution d) Vaporization of camphor
91. In which reaction(s), ΔS is negative?
- a) $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(s)$ b) $3\text{O}_2(g) \rightarrow 2\text{O}_3(g)$
- c) $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$ d) $\text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)$
92. For the adiabatic expansion of an ideal gas
- a) $PV^\gamma = \text{constant}$ b) $TV^{\gamma-1} = \text{constant}$ c) $TP^{1-\gamma} = \text{constant}$ d) None of the above
93. For which process does $\Delta U = 0$ holds true?
- a) Cyclic process b) Isothermal expansion c) Isochoric process d) Adiabatic process
94. Which of the following statements is/are correct?
- a) The evaporation of water is an endothermic change
- b) The conversion of white phosphorus to red phosphorus is an exothermic reaction
- c) The heat of neutralization of a strong acid with a strong base is always the same
- d) ΔH is negative for endothermic reactions
95. For which of the following substances is the heat of formation in the standard state zero?
- a) Sugar b) C_{diamond} c) Zinc d) C_{graphite}
96. The standard heat of formation of a compound is the:

- a) Change in enthalpy for the production of 1 mole of the compound at STP
 b) Change in enthalpy for the formation of 1 mole of the compound from its elements
 c) Change in enthalpy for the formation of 1 mole of the compound from its elements at 298 K and 1 atmospheric pressure
 d) Change in enthalpy for the formation of 1 mole of the compound from its elements at 25°C and a pressure of 760 mm of Hg.
97. In which of the following cases, do you consider the increase in entropy take(s) place?
 a) Pure liquid or liquid solutions are formed from solids
 b) Gases are formed, either from solids or liquids
 c) The number of molecules of gases increase in the course of a chemical reaction
 d) The temperature of a substance is increased
98. Which of the following statements is/are correct?
 a) The heat of neutralization of a strong acid with a strong base is always the same
 b) The enthalpy of combustion is always negative
 c) A spontaneous change involves a lowering of free energy
 d) The enthalpy of an element in the standard state assumed to be unity at 298 K.
99. In which of the following cases is the reaction spontaneous at all temperatures?
 a) $\Delta H > 0, \Delta S > 0$ b) $\Delta H < 0, \Delta S > 0$ c) $\Delta H < 0, \Delta S < 0$ d) $\Delta H < 0, \Delta S = 0$
100. The internal energy (U) of an ideal gas decreases by the same amount as the work done by the system:
 a) The process must be adiabatic
 b) The process must be isothermal
 c) The process must be isobaric
 d) The temperature must decrease
101. Which of the following are applicable for a thermochemical equation? It tells:
 a) About the physical state of reactants and products
 b) About the allotropic form (if any) of the reactants
 c) Whether the reaction is exothermic or endothermic
 d) Whether a particular reaction is spontaneous or not
102. Which of the following statement is/are correct?
 a) Heat, like work is a way of transferring energy
 b) Heat is not a property of the system, whereas the temperature is a property of the system
 c) Heat is manifested only at the boundary of system and surroundings
 d) None of the above
103. Among the following, the intensive property is (properties are):
 a) Molar conductivity b) Electromotive force c) Resistance d) Heat capacity
104. The change in enthalpy for an isobaric gaseous reaction (for an ideal gas system) is/are:
 a) $\Delta H = \Delta U + P\Delta V + V\Delta P$
 b) $\Delta H = \Delta U + P\Delta V$
 c) $\Delta H = \Delta U + \Delta nRT$
 d) $\Delta H = \Delta U + PV$
105. Which of the following statements are correct?
 a) Absolute value of enthalpy cannot be determined
 b) Absolute value of internal energy cannot be determined
 c) Absolute value of entropy can be determined
 d) Internal energy, enthalpy, and entropy are intensive properties
106. The standard heat of formation of U_3O_8 is $-853.5 \text{ kcal mol}^{-1}$ and standard heat of the reaction, $3UO_2 + O_2 \rightarrow U_3O_8$ is -76.01 kcal . The standard, heat of formation of UO_2 is/are :
 a) $-1083 \text{ kJ mol}^{-1}$ b) $-1102 \text{ kJ mol}^{-1}$ c) $-259 \text{ kcal mol}^{-1}$ d) $-302 \text{ kcal mol}^{-1}$
107. Which is/are correct for ideal gas?

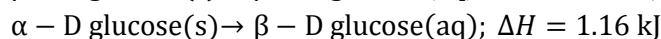
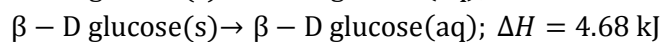
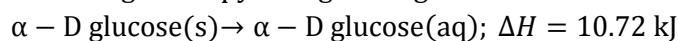
a) $\left(\frac{\delta U}{\delta T}\right)_P = 0$ b) $\left(\frac{\delta T}{\delta P}\right)_H = 0$ c) $\left(\frac{\delta E}{\delta V}\right)_T = 0$ d) $\left(\frac{\delta U}{\delta P}\right)_T = 0$

108. The reversible expansion of an ideal gas under adiabatic and isothermal conditions is shown in the figure. Which of the following statement(s) is (are) correct?

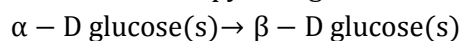


a) $T_1 = T_2$ b) $T_3 > T_1$ c) $w_{\text{isothermal}} > w_{\text{adiabatic}}$ d) $\frac{\Delta U_{\text{isothermal}}}{\Delta U_{\text{adiabatic}}} > 1$

109. Following enthalpy changes are given:



Calculate enthalpy change in

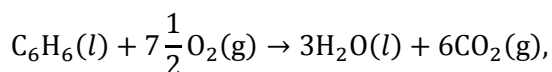


a) 14.24 kJ b) 16.56 kJ c) -7.2 kJ d) 4.88 kJ

110. All natural processes proceed spontaneously in a direction which:

- a) Increases entropy
- b) Increases free energy
- c) Decreases entropy
- d) Decreases free energy

111. The heat evolved in the combustion of benzene is given by



$\Delta H = -781.0 \text{ kcal mol}^{-1}$

When 156 g of C_6H_6 is burnt in an open container, the amount of heat energy released will be

a) 150.2 kcal mol⁻¹ b) 1562.0 kcal mol⁻¹ c) 6528.2 kJ mol⁻¹ d) 2448.5 kJ mol⁻¹

112. Which has/have a positive value(s) of ΔH ?

- a) $\text{H}_2(g) \rightarrow 2\text{H}(g)$
- b) $\text{H}(g) \rightarrow \text{H}^+(g) + e$
- c) $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$
- d) $\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}$

113. When a solid melts, there is/are:

- a) An increase in entropy
- b) An increase in enthalpy
- c) A decrease in internal energy
- d) A decrease in enthalpy

114. Which of the following is/are true in the case of an adiabatic process?

a) $q = W$ b) $\Delta U = q$ c) $q = 0$ d) $\Delta U = W$

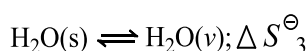
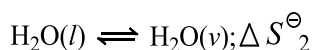
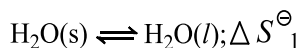
115. The intensive property is/are

a) Temperature b) Heat capacity c) Internal energy d) Pressure

116. Under which of the conditions the process will be spontaneous?

a) $\Delta G = -ve$ b) $\Delta_{\text{Total}}S = +ve$ c) $\Delta_{\text{Total}}S = -ve$ d) $\Delta_{\text{Total}}S = 0$

117. Select the correct statements for the equilibrium under standard conditions



- a) $\Delta S^\ominus_1 > \Delta S^\ominus_2$ b) $\Delta S^\ominus_2 \gg \Delta S^\ominus_1$ c) $\Delta S^\ominus_3 > \Delta S^\ominus_2$ d) $\Delta S^\ominus_3 > \Delta S^\ominus_1$

118. If W is work done by the system (–ve), the mathematical representation of the first law of thermodynamics is/are

- a) $\Delta U = \Delta q - W$ b) $q = \Delta U - W$ c) $dq + dW = 0$ d) $\Delta U = q + W$

119. The second law of thermodynamics states that

- a) All Spontaneous processes are thermodynamically irreversible
b) Entropy of the universe is continuously increasing
c) Energy can neither be created nor destroyed
d) Energy of the universe remain constant

120. Among the following which is/are the endothermic reaction(s)?

- a) Combustion of methane b) Decomposition of water
c) Dehydrogenation of ethane to ethylene d) Conversion of graphite to diamond

121. Average value of poisson's ratio for a mixture of 2 mol of each gas A and B is 1.66, then

- a) Gases are mono-atomic
b) Gases are diatomic
c) Average molar heat capacity at constant volume is 4 cal
d) Average molar heat capacity at constant V is 3 cal

122. Indicate in which case/cases the spontaneity of a change is favoured when

- a) ΔG is +ve b) ΔH is –ve c) ΔS is +ve d) ΔG is –ve

123. In a reaction, ΔH and ΔS both are more than zero. In which of the following cases, the reaction would not be spontaneous?

- a) $\Delta H > T \Delta S$ b) $T \Delta S > \Delta H$ c) $\Delta H = T \Delta S$ d) $\Delta G < 0$

124. Which is correct about ΔG ?

- a) $\Delta G = \Delta H - T \Delta S$ b) At equilibrium, $\Delta G^\ominus = 0$
c) At equilibrium, $\Delta G = -RT \log K$ d) $\Delta G = \Delta G^\ominus + RT \log K$

125. A reaction attains equilibrium state under standard conditions, then:

- a) Equilibrium constant $K = 0$ b) Equilibrium constant $K = 1$
c) $\Delta G^\ominus = 0$ and $\Delta H^\ominus = T \Delta S^\ominus$ d) $\Delta G = 0$ and $\Delta H = T \Delta S$

126. Which of the following are not correct at 298 K?

- a) $\Delta_f G^\ominus \text{element} = 0$ b) $\Delta_f H^\ominus \text{element} = 0$ c) $\Delta_f S^\ominus \text{element} = 0$ d) $\Delta_f G^\ominus \text{compound} = 0$

127. Which is not correct relationship between ΔG^\ominus and equilibrium constant K_p

- a) $K_p = -RT \log \Delta G^\ominus$ b) $K_p = [e/RT]^{\Delta G^\ominus}$ c) $K_p = -\frac{\Delta G^\ominus}{RT}$ d) $K_p = e^{-\Delta G^\ominus/RT}$

128. The enthalpy change for the process

$\text{C}(\text{graphite}) \rightarrow \text{C}(\text{g})$ is called

- a) Heat of vaporization b) Heat of sublimation
c) Heat of allotropic change d) Heat of atomisation

129. Which of the following are intensive properties?

- a) Heat capacity b) Refractive index c) Specific volume d) Entropy

130. Which of the following are state properties?

- a) Internal energy (U) b) Volume (V) c) Heat (q) d) Enthalpy (H)

131. In which reactions, ΔS is negative?

- a) $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$ b) $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(s)$
c) $3\text{O}_2(g) \rightarrow 2\text{O}_3(g)$ d) $\text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)$

132. Which of the following statements is/are correct?

- a) ΔH is – ve for exothermic reactions

- b) ΔH is +ve for endothermic reactions
 c) The heat of neutralization of strong acid and strong bases is constant
 d) The enthalpy of fusion is +ve
133. Enthalpy change equal internal energy change when
 a) All the reactants and products are in solution
 b) Reaction is carried out in a closed vessel
 c) Number of moles of gaseous reactants and that of products is equal
 d) Reaction is carried out at constant pressure
134. The process of evaporation of a liquid is accompanied by:
 a) Increase in enthalpy
 b) Decrease in free energy
 c) No change in free energy
 d) Increase in entropy
135. During an adiabatic reversible expansion of an ideal gas
 a) Internal energy of the system decreases
 b) Temperature of the system decreases
 c) The value of γ changes
 d) Pressure increases
136. Which of the following are irreversible processes?
 a) Mixing of two gases
 b) Evaporation of water at 373 K and 1 atm in a closed system
 c) Dissolution of NaCl in water
 d) $\text{H}_2\text{O}(s)$ at -4°C
137. If x and y are arbitrary extensive variables, then
 a) $(x + y)$ is an extensive variable
 b) x/y is an intensive variable
 c) dx/dy is an intensive variable
 d) Both (b) and (c)
138. A certain volume of dry air at NTP is expanded reversibly and isothermally to four times of its volume. The final pressure and temperature are respectively
 a) 25 atm and 273 K
 b) 0.25 atm and 0°C
 c) 25 atm and 0°C
 d) 0.25 atm and 273 K
139. Endothermic reactions, having $\Delta S = +ve$ may be spontaneous if
 a) $\Delta H > T \Delta S$
 b) $\Delta H < T \Delta S$
 c) $\Delta H = T \Delta S$
 d) T is very high
140. The poisson's ratio for O_2 is 1.4. Which of the following are correct for O_2 ?
 a) $C_{VM} = 5 \text{ cal}$
 b) $C_V = 0.156 \text{ cal}$
 c) $C_P = \frac{R\gamma}{\gamma - 1}$
 d) $C_V = \frac{R}{(\gamma - 1)}$
141. Which of the following conditions is/are favourable for the feasibility of a reaction?
 a) $\Delta H = -ve; T\Delta S = +ve$
 b) $\Delta H = -ve; T\Delta S = -ve; T\Delta S < \Delta H$
 c) $\Delta H = +ve; T\Delta S = +ve; T\Delta S < \Delta H$
 d) $\Delta H = +ve; T\Delta S = +ve; T\Delta S > \Delta H$
142. If w_1, w_2, w_3 and w_4 are work done in isothermal, adiabatic, isobaric, and isochoric reversible expansion for an ideal gas, respectively, then
 a) $w_3 > w_1$
 b) $w_1 > w_2$
 c) $w_2 > w_4$
 d) $w_4 > w_2$
143. During the isothermal expansion of an ideal gas:
 a) The internal energy remains unaffected
 b) The temperature remains constant
 c) The enthalpy remains unaffected
 d) The enthalpy becomes zero
144. If x and y are arbitrary intensive variables, then
 a) xy is an intensive variable
 b) x/y is an intensive variable
 c) $(x + y)$ is an extensive property
 d) dx/dy is an intensive property
145. If it is necessary to employ electric current (electrolysis) to carry out a chemical reaction, then for that reaction:
 a) $\Delta H > 0$
 b) $\Delta G = \Delta H$
 c) $\Delta G > 0$
 d) $\Delta S > 0$
146. If q is the amount of heat absorbed by the system and W the amount of work done on the system (+ve),

the change in the energy of the system is given by:

- a) $q = \Delta U + W$ b) $q = \Delta U - W$ c) $\Delta U = q - W$ d) $\Delta U = W + q$

147. Hess' law is applicable for determination of enthalpy of

- a) Reaction b) Formation c) Transition d) None of these

148. Which of the options given below are correct?

ΔH ΔS Nature of reaction

- a) (-)(+) Spontaneous at all temperature
b) (+)(-) Nonspontaneous regardless of temperature
c) (+)(+) Spontaneous only at high temperature
d) (-)(-) Spontaneous only at low temperature

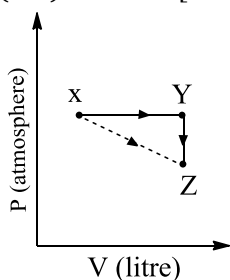
149. The open system(s) is/are which

- a) Can exchange matter with the surroundings
b) Can exchange energy with the surroundings
c) Can exchange both matter and energy with the surroundings
d) Cannot exchange either matter or energy with the surroundings

150. The heat of reaction depends upon:

- a) The manner by which the reaction is carried out
b) Temperature at which the reaction is carried out
c) Physical state of reactants and products
d) Whether the reaction is carried out at constant pressure or at constant volume

151. For an ideal gas, consider only $P - V$ work in going from an initial state X to the final state Z . The final state Z can be reached by either of the two paths shown in the figure. Which of the following choice(s) is (are) correct? [take ΔS as change in entropy and w as work done]



- a) $\Delta S_{x \rightarrow z} = \Delta S_{x \rightarrow y} + \Delta S_{y \rightarrow z}$ b) $w_{x \rightarrow z} = w_{x \rightarrow y} + w_{y \rightarrow z}$ c) $w_{x \rightarrow y \rightarrow z} = w_{x \rightarrow y}$ d) $\Delta S_{x \rightarrow y \rightarrow z} = \Delta S_{x \rightarrow y}$

152. Which of the following are endothermic processes?

- a) Combustion of glucose b) Decomposition of water
c) Dehydrogenation of ethane to ethane d) Conversion of graphite to diamond

153. The heat of neutralization of a strong acid by a strong base is a constant

- a) 57.0 kJ mol^{-1} b) $13.7 \text{ kcal mol}^{-1}$ c) $5.7 \times 10^4 \text{ J mol}^{-1}$ d) $13.7 \times 10^3 \text{ cal mol}^{-1}$

154. Among the following, the state function(s) is (are):

- a) Internal energy
b) Irreversible expansion work
c) Reversible expansion work
d) Molar enthalpy

155. Which of the following are thermodynamically stable?

- a) C(diamond) b) C(graphite) c) P_4 (white) d) P_4 (black)

156. The following is(are) endothermic reaction(s)

- a) Combustion of methane b) Decomposition of water
c) Dehydrogenation of ethane to ethylene d) Conversion of graphite to diamond

157. Which is an irreversible process?

- a) Mixing of two gases by diffusion b) Evaporation of water at 373 K and 1 atm pressure
c) Dissolution of NaCl in water d) All of the above

158. Select the correct statements

- a) The magnitude of work involved in an intermediate irreversible expansion is less than that involved in reversible expansion
- b) Heat absorbed during intermediate reversible expansion is more than in intermediate irreversible expansion
- c) The magnitude of work involved in an intermediate reversible compression is more than that involved in intermediate irreversible compression
- d) Heat released during intermediate irreversible compression is more than that in intermediate reversible compression

159. Which of the following statements is/are false?

- a) Work is a state function
- b) Temperature is a state function
- c) Change in the state is completely defined when the initial and final states are specified
- d) Work appears at the boundary of the system

160. Which of the following are true about resonance energy?

- a) Resonance energy = Experimental heat of formation – Calculated heat of formation
- b) Resonance energy = Calculated heat of formation – Experimental heat of formation
- c) Greater the resonance energy, more the compound will be stable
- d) Lesser the resonance energy, more the compound will be stable

Assertion - Reasoning Type

This section contains 0 questions numbered 161 to 160. Each question contains STATEMENT 1 (Assertion) and STATEMENT 2 (Reason). Each question has the 4 choices (a), (b), (c) and (d) out of which **ONLY ONE** is correct.

- a) Statement 1 is True, Statement 2 is True; Statement 2 **is** correct explanation for Statement 1
- b) Statement 1 is True, Statement 2 is True; Statement 2 **is not** correct explanation for Statement 1
- c) Statement 1 is True, Statement 2 is False
- d) Statement 1 is False, Statement 2 is True

161

Statement 1: Helium has lower entropy than CO₂ gas which has lower entropy than gaseous benzene

Statement 2: The larger the complexity of molecule, the larger is its absolute entropy

162

Statement 1: Heat of neutralization for HF is – 68.552 kJ/eq whereas for HCl it is – 57.26 kJ/eq.

Statement 2: The acid HF is weak acid.

163

Statement 1: Work and heat are not state functions.

Statement 2: The sum of $q + W$ is state function.

164

Statement 1: Internal energy of a system is an extensive property.

Statement 2: The internal energy of a system depends upon the amount and physical state of the substance.

165

Statement 1: $C + \frac{1}{2}O_2 \rightarrow CO; \quad \Delta H = -26.0 \text{ kcal}$

$CO + \frac{1}{2}O_2 \rightarrow CO_2; \quad \Delta H = -68.3 \text{ kcal}$

$\therefore C + O_2 \rightarrow CO_2; \quad \Delta H = -94.3 \text{ kcal}$

Statement 2: This is an experimental proof of Hess's law.

166

Statement 1: The change in entropy during melting of ice is negligible in comparison to change in entropy during vaporization .

Statement 2: The volume occupied by solid and liquid is too less in comparison to volume occupied by gas.

167

Statement 1: Internal energy change in a cyclic process is zero

Statement 2: Internal energy is a state function

168

Statement 1: The enthalpy of formation of HCl is equal to the bond energy of HCl

Statement 2: The enthalpy of formation and the bond energy both involve formation of one mole of HCl from the elements

169

Statement 1: The ratio of heat of vaporization and the normal boiling point of a liquid is approximately 88 J/mol.

Statement 2: This is Trouton's rule derived by experimental data.

170

Statement 1: The entropies of CO, NO, N₂O, Cl₂(s) are not zero at absolute zero.

Statement 2: These are exceptions to III law of thermodynamics.

171

Statement 1: When a real gas is allowed to expand adiabatically through a fine hole from a region of high pressure to a region of low pressure, the temperature of the gas falls

Statement 2: Work is done at the cost of internal energy of the gas

172

Statement 1: The heat of neutralization of perchloric acid, HClO₄, with NaOH is same as that of HCl with NaOH

Statement 2: Both HCl and HClO₄ are strong acids

173

Statement 1: Heat of combustion are always exothermic.

Statement 2: Combustion of N_2 to give NO is exothermic.

174

Statement 1: When hydrogen gas at high pressure and room temperature expands adiabatically into a region of low pressure, there is a decrease in temperature

Statement 2: Hydrogen gas at *room* temperature is above its inversion temperature

175

Statement 1: The enthalpy of formation of $H_2O(l)$ is greater than that of $H_2O(g)$

Statement 2: Enthalpy change is negative for the condensation reaction $H_2O(g) \rightarrow H_2O(l)$

176

Statement 1: Standard heat enthalpy of a compound is its heat of formation of 25°C and 1 atm.

Statement 2: Standard heat enthalpy of pure elements have arbitrarily assumed to be zero.

177

Statement 1: Combustion of N_2 to give NO is endothermic.

Statement 2: Bond energy of N_2 is very high.

178

Statement 1: There is a natural asymmetry between converting work to heat and converting heat to work.

Statement 2: No process is possible in which the sole result is the absorption of heat from a reservoir and its complete conversion into work.

179

Statement 1: Heat of neutralization for both HNO_3 and HCl with NaOH is $53.7 \text{ kJ per mol}^{-1}$

Statement 2: NaOH is a strong electrolyte/base

180

Statement 1: Neither q nor W is state function but $q + W$ is state function.

Statement 2: ΔU is state function.

181

Statement 1: Zeroth law can also be termed as law of thermal equilibrium

Statement 2: Two objects in thermal equilibrium with the third one, are in thermal equilibrium with each other

182

Statement 1: Phase transition involves change in internal energy only.

Statement 2: Phase transition occurs at constant pressure.

183

Statement 1: In a diatomic molecule involving two like atoms covalently bonded with each other, bond energy = $2 \times$ heat of formation of atom.

Statement 2: $\text{H}_2 \rightarrow 2\text{H}; e_{\text{H}-\text{H}} = \Delta H.$

184

Statement 1: The zeroth law of thermodynamics was known before I law of thermodynamics.

Statement 2: The zeroth law concerning thermal equilibrium was appeared after three laws (I, II and III) of thermodynamics and thus, was named as zeroth law.

185

Statement 1: The SI unit of entropy is $\text{JK}^{-1}\text{mol}^{-1}.$

Statement 2: $\Delta S = \frac{q_{rev}}{T}.$

186

Statement 1: The endothermic reactions are favoured at lower temperature and the exothermic reactions are favoured at higher temperature

Statement 2: When a system in equilibrium is disturbed by changing the temperature, it will tend to adjust itself so as to overcome the effect of the change

187

Statement 1: Both work and heat are manifested by an effect in the surroundings.

Statement 2: Work done by/on the system and ΔH appear only at the boundary of system.

188

Statement 1: The dissolution of NaCl in water is endothermic, though it is spontaneous process.

Statement 2: ΔG for the process is -ve because ΔH_{sol} is very low and thus $T\Delta S > \Delta H.$

189

Statement 1: The enthalpy of both graphite and diamond is taken to be zero, being elementary substances

Statement 2: The enthalpy of formation of an elementary substance in any state is taken as zero

190

Statement 1: A reaction which is spontaneous and accompanied by decrease of randomness must be exothermic

Statement 2: All exothermic reactions are accompanied by decrease of randomness

191

Statement 1: Decrease in free energy causes spontaneous reaction

Statement 2: Spontaneous reactions are invariably exothermic

192

Statement 1: An exothermic process, non-spontaneous at high temperature, may become spontaneous at low pressure

Statement 2: With decrease in temperature, randomness (entropy) decreases

193

Statement 1: The change in internal energy and change in heat enthalpy does not depend upon the path by which changes are brought in.

Statement 2: Both ΔU and ΔH are path independent as U and H are state functions.

194

Statement 1: Fall of water as rain drops from clouds is spontaneous.

Statement 2: During the process entropy increases.

195

Statement 1: The enthalpy of formation of gaseous oxygen molecules at 298 K and under a pressure of 1 atm is zero

Statement 2: The entropy of formation of gaseous oxygen molecules under the same condition is zero

196

Statement 1: A non-spontaneous endothermic reaction at room temperature may be spontaneous at high temperature.

Statement 2: At high temperature $T\Delta S$ becomes more than ΔH .

197

Statement 1: When a gas at high pressure expands against vacuum, the work done is maximum

Statement 2: Work done in expansion depends upon the pressure inside the gas and increase in volume

198

Statement 1: For a particular reaction, heat of combustion at constant pressure (q_p) is always greater than that at constant volume (q_v)

Statement 2: Combustion reactions are invariably accomplished by increase in number of moles

199

Statement 1: $C_p - C_v = R$ for an ideal gas

Statement 2: $\left[\frac{\partial E}{\partial V}\right]_T = 0$ for an ideal gas

200

Statement 1: Heat of combustion of C_2H_6 is $-341.1 \text{ kcal mol}^{-1}$ and heat of combustion of C_2H_2 is $-310 \text{ kcal mol}^{-1}$ but C_2H_2 is better fuel.

Statement 2: The better fuel has high calorific value.

201

Statement 1: There is no reaction known for which ΔG is positive, yet it is spontaneous

Statement 2: For photochemical reaction, ΔG is negative

202

Statement 1: Many endothermic reactions that are not spontaneous at room temperature become spontaneous at high temperature

Statement 2: Entropy of the system increases with increase in temperature

203

Statement 1: The Heat of ionization of water is equal to the heat of neutralization of a strong acid with a strong base

Statement 2: Water ionizes to a very small extent while H^{\oplus} ions from an acid combine very rapidly with OH^{\ominus} from a base to form H_2O

204

Statement 1: The thermodynamic factor which determines the spontaneity of a process is the free energy. For a process to be spontaneous the free energy must be - ve

Statement 2: The change in free energy is related to the change in a process must always be positive if it is spontaneous

205

Statement 1: Internal energy is an extensive property

Statement 2: Internal energy depends upon the amount of the system

206

Statement 1: Enthalpy of graphite is lower than that of diamond

Statement 2: Entropy of graphite is lower than that of diamond

207

Statement 1: The Joules-Thomson coefficient for an ideal gas is zero

Statement 2: There are no intermolecular attractive forces in an ideal gas

208

Statement 1: Absolute values of internal energy of substance can't be determined

Statement 2: It is impossible to determine exact values of constituent energies of the substances

209

Statement 1: Bond energy for breaking up a bond is endothermic.

Statement 2: Heat is required to overpower the attractions between two atoms.

210

Statement 1: The heat absorbed during the isothermal expansion of an ideal gas against vacuum is zero.

Statement 2: The volume occupied by the molecules of an ideal gas is zero.

211

Statement 1: Heat of neutralization can be given as : $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$; $\Delta H = -13.6 \text{ kcal}$

Statement 2: Heat of neutralization can be alternatively defined as heat of formation of water.

212

Statement 1: The mass and volume of a substance are the extensive properties and are proportional to each other

Statement 2: The ratio of mass of a sample to its volume is an intensive property

213

Statement 1: The variation of heat of reaction with temperature are given in terms of Kirchhoff's equation.

Statement 2: The Kirchhoff's equation is : $\Delta H = \Delta U + \Delta nRT$.

214

Statement 1: Pressure, volume, and temperature are all extensive properties

Statement 2: Extensive properties depend upon the amount and nature of the substance

Matrix-Match Type

This section contain(s) 0 question(s). Each question contains Statements given in 2 columns which have to be matched. Statements (A, B, C, D) in **columns I** have to be matched with Statements (p, q, r, s) in **columns II**.

215.

Column-I	Column- II
(A) $\text{CO}_2(s) \rightarrow \text{CO}_2(g)$	(1) Phase transition
(B) $\text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}(g)$	(2) Allotropic change
(C) $2\text{H} \rightarrow \text{H}_2(g)$	(3) $\Delta H = +ve$
(D) $\text{P}_{(\text{White,solid})} \rightarrow \text{P}_{(\text{red,solid})}$	(4) $\Delta S = +ve$
	(5) $\Delta S = -ve$

CODES :

	A	B	C	D
a)	4	5	1,2,3	1,3,4
b)	1,3,4	4	5	1,2,3
c)	1,2,3	5	4	1,3,4
d)	5	1,2,3	4	1,3,4

216.

Column-I	Column- II
(A) Amount of heat required to raise the temperature of 1 mol substance by 1°C	(p) Specific heat \times Molar mass
(B) $\frac{\Delta H}{\Delta T}$ or $\frac{\Delta U}{\Delta T}$	(q) Heat capacity = C, C_p or C_v

- (C) Heat evolved in the combustion of 1 g of a substance
 (D) Heat evolved when an extra electron is added to valence shell of an isolated gaseous atom
- (r) Electron gain enthalpy
 (s) Calorific value

CODES :

	A	B	C	D
a)	Q,p	s	r	p,q
b)	s	r	p,q	q,p
c)	p,q	q,p	s	r
d)	r	p,q	q,p	s

217.

Column-I

Column- II

- (A) Enthalpy
 (B) Temperature
 (C) Free energy
 (D) Work
- (p) Intrinsic property
 (q) Path function
 (r) Function of U, P, V
 (s) State function

CODES :

	A	B	C	D
a)	p	r	q	s
b)	q	r	p	r
c)	r	p	r	q
d)	s	p	q	r

218.

Column-I

Column- II

- (A) $\Delta_{\text{total}}S > 0$
 (B) $\Delta_{\text{total}}S < 0$
 (C) ΔH
 (D) $-\Delta G$
- (p) w_{useful} done by system
 (q) Second law of thermodynamics
 (r) Non-spontaneous
 (s) $\Delta U > \Delta nRT$

CODES :

	A	B	C	D
a)	p	q	r	s
b)	q	r	s	p

- c) s p q r
 d) r s p q

219.

Column-I

Column- II

- (A) Gibbs-Helmholtz reaction
 (B) First law of thermodynamic
 (C) Enthalpy at constant pressure
 (D) Entropy

- (p) Degree of randomness
 (q) $\Delta H = \Delta U + P \Delta V$
 (r) Law of conservation of energy
 (s) $\Delta G = \Delta H - T \Delta S$

CODES :

	A	B	C	D
a)	r	s	q	p
b)	p	q	s	r
c)	s	r	p	q
d)	q	p	r	s

220.

Column-I

Column- II

- (A) Isothermal process (reversible)
 (B) Adiabatic process (reversible)
 (C) Adiabatic free expansion
 (D) Isothermal free expansion

- (p) $w = 2.303nRT \log \left(\frac{P_1}{P_2} \right)$
 (q) $w = 0$
 (r) $w = 2.303nRT \log \left(\frac{V_2}{V_1} \right)$
 (s) $w = \frac{nR}{(\gamma - 1)} (T_2 - T_1)$

CODES :

	A	B	C	D
a)	s	q	p,r	s
b)	p,r	s	q	q
c)	q	p,r	s	q
d)	p	s	q	p,r

221.

Column-I

Column- II

- (A) O_2
 (B) N_2
 (C) CO_2
 (D) 1 mol O_2 + 2 mol O_3

- (p) $\gamma = 1.4$
 (q) $C_p = \frac{7}{2}R$
 (r) $C_v = \frac{17}{6}R$
 (s) $\gamma = 1.33$

CODES :

	A	B	C	D
a)	P,q	p,q	r	r
b)	r	s	p,q	r
c)	p	r	s	p,q
d)	p,q	q	r	s

222.

Column-I

- (A) $N_2(g) + O_2(g) \rightarrow 2NO(g)$
 (B) $2KI(aq) + HgI_2(aq) \rightarrow K_2[HgI_4](aq)$
 (C) $PCl_3(g) + Cl_2(g) \rightarrow PCl_5(g)$
 (D) $NH_3(g) + HCl(g) \rightarrow NH_4Cl(g)$

Column- II

- (p) $\Delta S = 0$
 (q) $\Delta S < 0$
 (r) $\Delta H > 0$
 (s) $\Delta H < 0$

CODES :

	A	B	C	D
a)	Q,r	q,s	q	p,q
b)	q	q,r	p,q	q,s
c)	q,s	p,q	q	q,r
d)	p,q	q	q,r	q,s

223.

Column-I

- (A) $C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$
 (B) $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$
 (C) $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O$
 (D) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$

Column- II

- (p) Combustion
 (q) Neutralization
 (r) Process of formation
 (s) Used in fuel cell

CODES :

	A	B	C	D
a)	P,r	p	q	p,r,s
b)	p	q	p,r	r,s
c)	q	r,s	q	p,r
d)	p,r,s	p,r	p	q

224.

	Column-I
(A)	For spontaneous reaction
(B)	For endothermic reaction
(C)	Bond dissociation energy
(D)	For solids and liquids in a thermochemical reaction

CODES :

	A	B	C	D
a)	r	s	p	q
b)	s	r	q	p
c)	p	q	r	s
d)	q	p	s	r

225.

	Column-I
(A)	Hess' law
(B)	Combustion reaction
(C)	Trouton's law
(D)	Clausius-Cal-peyron equation

CODES :

	A	B	C	D
a)	s	r	q	p
b)	r	s	p	q
c)	p	q	s	r

Column- II

- (p) $\sum (\text{BE})_R - \sum (\text{BE})_P$
 (q) $\Delta H = \Delta U$
 (r) $\Delta G = -ve$
 (s) $\sum H_P > \sum H_R$

Column- II

- (p) $2.303 \log \frac{P_2}{P_1} = \frac{\Delta_{\text{vap}}H}{R} \left(\frac{T_2 - T_1}{T_1 T_2} \right)$
 (q) $\Delta_{\text{vap}}H = 88 \text{ J K}^{-1} \text{ mol}^{-1} \times \text{Boiling point in Kelvin}$
 (r) Exothermic
 (s) ΔH remains the same irrespective steps

d) q p r s

226.

Column-I

Column- II

- (A) Spontaneous process
 (B) Heat flow from high temperature of system towards low temperature of surroundings
 (C) Exergonic process
 (D) Increase in the randomness of system by heating

- (p) $\Delta H = -ve$
 (q) $\Delta G = +ve$
 (r) $\Delta_{total}S = +ve$
 (s) $\Delta G = -ve$

CODES :

	A	B	C	D
a)	P,r,s	q	q,r	r,s
b)	r,s	p,r,s	q	q,r
c)	q	q,r	r,s	p,rs
d)	q,r	r,s	p,rs	q

227.

Column-I

Column- II

- (A) ΔG
 (B) ΔG^\ominus
 (C) w or w_{max}
 (D) ΔS^\ominus

- (p) $-nFE$ or $-nFE^\ominus$
 (q) $\Delta H - T \Delta S$
 (r) $-RT \log_e k$
 (s) $2.303nR \log_{10} \left(\frac{V_2}{V_1} \right)$

CODES :

	A	B	C	D
a)	P,q	p,r	p	s
b)	p	p,q	s	p,r
c)	p,r	s	p,q	p
d)	s	p	p,r	p,q

228.

Column-I

Column- II

- (A) $\left(\frac{\partial U}{\partial V} \right)_T = 0$
 (B) $W = -\Delta U$

- (p) Isothermal process
 (q) $-nFE^\ominus$

(C) $\Delta U = 0$

(D) ΔG^\ominus

(E) $\left(\frac{\partial T}{\partial P}\right)_H \neq 0$

CODES :

	A	B	C	D	E
a)	r	t	q	s	q
b)	t	r	p	q	q
c)	q	p	s	t	q
d)	s	q	t	p	q

(r) Adiabatic reaction

(s) Van der waals gas

(t) Ideal gas

229.

Column-I

Column- II

(A) $\frac{\Delta H_2 - \Delta H_1}{T_2 - T_1} = \Delta C_p$

(B) $\frac{\Delta U_2 - \Delta U_1}{T_2 - T_1} = \Delta C_v$

(C) $\frac{\Delta_{\text{fusion}}H}{T_{\text{mp}}} = \Delta_{\text{fusion}}S$

(D) $\lim_{T \rightarrow 0K} S \rightarrow 0$

(p) Trouton equation

(q) Effect of temperature on the heat of reaction

(r) Kirchhoff's equation

(s) Third law of thermodynamics

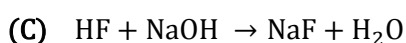
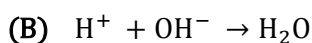
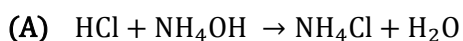
CODES :

	A	B	C	D
a)	Q,r	p,r	s	p
b)	p	s	q,r	p,r
c)	s	p	p,r	q,r
d)	p,r	q,r	p	s

230.

Column-I

Column- II



(1) $\Delta H = -13.7 \text{ kcal}$

(2) $\Delta H > -13.7 \text{ kcal}$

(3) $\Delta H < -13.7 \text{ kcal}$

CODES :

	A	B	C	D
a)	2	3	1	

- b) 3 1 2
 c) 1 2 3
 d) 1 3 2

231.

Column-I

- (A) Joule-Thomson coefficient
 (B) Kirchhoff's equation
 (C) Ideal gas
 (D) Inversion temperature

Column- II

- (p) $\left(\frac{\partial U}{\partial V}\right)_T = 0$
 (q) $\left(\frac{2a}{RT} - b\right) = 0$
 (r) $\left(\frac{\partial \Delta H}{\partial T}\right)_P = \Delta C_P$
 (s) $\left(\frac{\partial T}{\partial P}\right)_H = \frac{-(\partial H/\partial P)_T}{C_P}$

CODES :

	A	B	C	D
a)	r	p	q	s
b)	p	q	s	r
c)	s	r	p	q
d)	q	s	r	p

232.

Column-I

- (A) $(\text{Heat of formation})_{\text{exp}} - (\text{Heat of formation})_{\text{calculated}}$
 (B) $\sum(\text{BE})_{\text{Reactant}} - \sum(\text{BE})_{\text{Products}}$
 (C) $\sum(\text{Heat of combustion})_{\text{Reactants}} - \sum(\text{Heat of combustion})_{\text{Products}}$
 (D) $C_P \Delta T$

Column- II

- (p) Heat of reaction
 (q) Resonance energy
 (r) ΔH
 (s) Used in calorimetric method

CODES :

	A	B	C	D
a)	P,r,s	q	r,s	p,r
b)	q	p,r	p,r	p,r,s
c)	p,r	p	p,r,s	q
d)	p,r	p,r,s	q	r,s

233.

Column-I

Column- II

- (A) O_2
 (B) N_2
 (C) CO_2, CH_4
 (D) 1 mol O_2 + 2 mol O_3

CODES :

	A	B	C	D
a)	s	p,q	p	r
b)	p,q	p,q	s	s
c)	q	s	p,q	s
d)	p,q	p	s	p,q

- (p) $\gamma = 1.4$
 (q) $C_p = \frac{7}{2}R$
 (r) $\frac{23}{6}R$
 (s) $\gamma = 1.33$

234.

Column-I

- (A) $\Delta U = 0, \Delta H = 0$
 (B) $\Delta H = -ve$
 (C) $\Delta_{univ}S = 0$
 (D) $\Delta G_{mix} = -ve$

CODES :

	A	B	C	D
a)	P,r	s	r	q
b)	s	p,r	q	r
c)	r	q	p,r	s
d)	q	r	s	p,r

Column- II

- (p) $CH_2 = CH_2(g) + H_2(g) \rightarrow CH_3CH_3(g)$
 (q) For reversible process
 (r) Non ideal solutions with negative
 (s) For an ideal gas undergoing expansion isothermally

235.

Column-I

- (A) $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$
 (B) $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$
 (C) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
 (D) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

CODES :

Column- II

- (p) $\Delta H = \Delta U + RT$
 (q) $\Delta H = \Delta U$
 (r) $\Delta H = \Delta U - 2RT$
 (s) Forward shift by increasing pressure

	A	B	C	D
a)	q	q	p	r,s
b)	q	p	r,s	s
c)	r,s	s	p	q
d)	p	r,s	q	p

236.

	Column-I	Column- II
(A)	$\Delta_{\text{sys}}S > 0$ (Isolated system)	(p) Spontaneous
(B)	$\Delta G < 0$	(q) Non-spontaneous photochemical reaction
(C)	$\Delta_{\text{Total}}S = 0$	(r) Equilibrium
(D)	$(\Delta G)_{\text{TP}} > 0$	(s) Non-spontaneous

CODES :

	A	B	C	D
a)	p	p,q	r	s
b)	p,q	r	s	p
c)	r	s	p	p,q
d)	s	p	p,q	r

237.

	Column-I	Column- II
(A)	Exothermic	(p) $H_P - H_R$
(B)	Endothermic	(q) ΔH is negative
(C)	Spontaneous	(r) ΔH is positive
(D)	Heat of reaction	(s) ΔG is negative

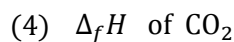
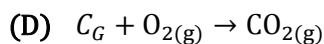
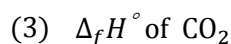
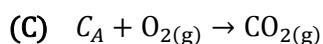
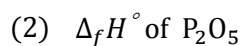
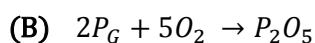
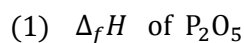
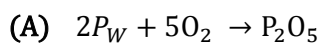
CODES :

	A	B	C	D
a)	s	p	q	r
b)	q	r	s	p
c)	p	s	r	q
d)	r	q	p	s

238.

Column-I

Column- II



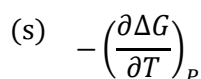
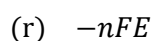
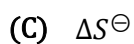
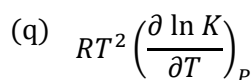
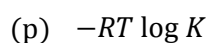
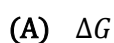
CODES :

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a)	1,4,5	3	2	5,2,4
b)	3,2,1	1	5	4,2,1
c)	1,3,4	4	5	1, 2, 3
d)	5,3,4	2	4	4, 5,3

239.

Column-I

Column- II



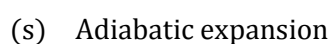
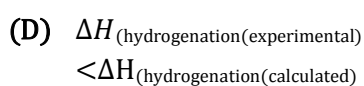
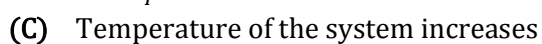
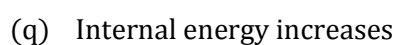
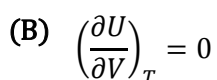
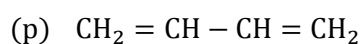
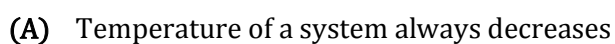
CODES :

	A	B	C	D
a)	p,r	q	s	p,r
b)	q	s	p,r	p
c)	s	p,r	q	p,r
d)	p,r	s	p,r	q

240.

Column-I

Column- II



CODES :

	A	B	C	D
a)	s	r	q	p
b)	p	q	r	s
c)	r	s	p	q
d)	q	p	s	r

241.

Column-I

Column- II

- | | |
|---|---------------------------------|
| (A) The amount of energy which must be invested in reaction to get it started | (p) Chemical equilibrium |
| (B) A state in which the rate of forward reaction is exactly equal to the rate of the reverse reaction | (q) Activation energy |
| (C) A process or reaction which consumes heat | (r) Endothermic |
| (D) A process or reaction that releases heat | (s) Exothermic |
| | (t) Threshold energy |

CODES :

	A	B	C	D
a)	q	p	r	s
b)	r	q	s	p
c)	s	p	q	r
d)	p	r	s	q

242.

Column-I

Column- II

- | | |
|--|---|
| (A) Isothermal process (reversible) | (p) $w = 2.303nRT \log \left(\frac{P_1}{P_2} \right)$ |
| (B) Adiabatic process | (q) $PV^\gamma = \text{constant}$ |
| (C) $w = \frac{nR}{\gamma - 1} (T_2 - T_1)$ | (r) $w = 2.303nRT \log \left(\frac{V_2}{V_1} \right)$ |
| (D) Irreversible isothermal process | (s) $w = -P_{\text{ex}}(V_2 - V_1)$ |

CODES :

A	B	C	D
----------	----------	----------	----------

- a) q s p,r q
 b) s r p p,r
 c) p,r q q s
 d) q p,r s q

243.

Column-I

Column- II

(A) Kirchoff's equation

$$(1) \Delta G = \Delta H + T \left(\frac{\partial \Delta G}{\partial T} \right)$$

(B) Clausius Clapeyron equation

$$(2) \frac{d \ln K_p}{dT} = \frac{\Delta H^\circ}{RT^2}$$

(C) Van't Hoff isochore

$$(3) \frac{d \ln P}{dT} = \frac{\Delta H^\circ}{RT^2}$$

(D) Gibb's Helmholtz equation

$$(4) \frac{\partial(\Delta G)}{\partial T} = \Delta C_p$$

CODES :

	A	B	C	D
a)	4	3	2	1
b)	1	2	3	4
c)	3	4	1	2
d)	2	1	4	3

Linked Comprehension Type

This section contain(s) 48 paragraph(s) and based upon each paragraph, multiple choice questions have to be answered. Each question has atleast 4 choices (a), (b), (c) and (d) out of which **ONLY ONE** is correct.

Paragraph for Question Nos. 244 to -244

If the boundary of system moves by an infinitesimal amount, the work involved is given by

$$dW = -p_{\text{ext}}dV$$

For irreversible process, $W = -p_{\text{ext}}\Delta V$

For reversible process, $p_{\text{ext}} = p_{\text{int}} \pm dp \approx p_{\text{int}}$

So, for reversible isothermal process, $W = -nRT \ln \frac{V_j}{V_i}$ 2 moles of an ideal gas undergoes isothermal compression along three different paths

(a) A single stage compression against a constant external pressure of 20 bar

(b) Reversible compression from $p_i = 2$ bar and $V_i = 8$ L to $p_j = 20$ bar

(c) A two stage compression consisting initially of compression against a constant external pressure of 10 bar until $p_{\text{gas}} = p_{\text{ext}}$, followed by compression against a constant pressure of 20 bar until $p_{\text{gas}} = p_{\text{ext}}$

244. Work done on the gas in single stage compression is

- a) 144 bar-L b) 98 bar-L c) 54 bar-L d) 121 bar-L

Paragraph for Question Nos. 245 to - 245

Molar heat capacity is the heat required to raise the temperature of one mole of material by one degree, since heat is not a state function, the amount of heat required to produce a given change in its state depends on the path followed

C_p = Specific heat \times Molecular weight. It is measured at constant pressure

C_v = Specific heat \times Molecular weight. It is measured at constant volume

245. The specific heat of a gas at constant volume is 0.075 cal/g. Predict the atomicity of the gas. Molar mass of gas is 40 g mol⁻¹

- a) 3 b) 2 c) 1 d) None of these

Paragraph for Question Nos. 246 to - 246

Gibbs-Helmholtz equation relates the free energy change to the enthalpy and entropy changes of the process as $(\Delta G)_{PT} = \Delta H - T \Delta S$

The magnitude of ΔH does not change much with the change in temperature but the entropy factor $T \Delta S$ changes appreciably. Thus, spontaneity of a process depends very much on temperature

246. When CaCO₃ is heated to a high temperature, it undergoes decomposition into CaO and CO₂ whereas it is quite stable at room temperature. The most likely explanation of it is

- a) The enthalpy of reaction (ΔH) overweighs the term $T \Delta S$ at high temperature
b) The term $T \Delta S$ overweighs the enthalpy of reaction at high temperature
c) At high temperature, both enthalpy of reaction and entropy change become negative
d) None of these

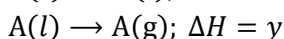
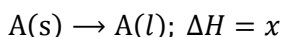
Paragraph for Question Nos. 247 to - 247

The enthalpy change for chemical reaction is denoted as ΔH and $\Delta H^\ominus = H_P^\ominus - H_R^\ominus$. The relation between enthalpy and internal energy is expressed by equation:

$$\Delta H = \Delta U + \Delta nRT$$

Where ΔU = change in internal energy, Δn = change in number of moles, R = gas constant

247. Given that:

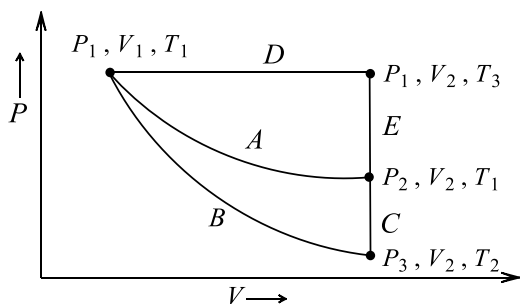


The heat of sublimation of A will be

- a) $x - y$ b) $x + y$ c) x or y d) $-(x + y)$

Paragraph for Question Nos. 248 to - 248

For an ideal gas, an illustration of three different paths A, (B + C) and (D + E) from an initial state P_1, V_1, T_1 to a final state P_2, V_2, T_1 is shown in the given figure



Path A represents a reversible isothermal expansion from P_1, V_1 to P_2, V_2 . Path (B + C) represents a reversible adiabatic expansion (B) from P_1, V_1, T_1 to P_3, V_2, T_2 followed by reversible heating the gas at constant volume (C) from P_3, V_2, T_2 to P_2, V_2, T_1 . Path (D + E) represents a reversible expansion at constant pressure P_1 (D) from P_1, V_1, T_1 to P_1, V_2, T_3 followed by a reversible cooling at constant volume V_2 (E) from P_1, V_2, T_3 to P_2, V_2, T_1

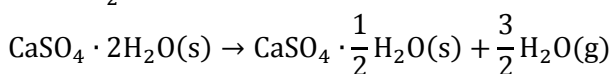
248. What is q_{rev} , for path (A)?

- a) $P(V_2 - V_1)$ b) $-nRT_1 \ln \frac{V_2}{V_1}$ c) $-nR \ln \frac{V_2}{V_1}$ d) $-nRT_1 \ln \frac{V_2}{V_1}$

Paragraph for Question Nos. 249 to - 249

Concrete is produced from a mixture of cement, water and small stones. Small amount of gypsum, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ is added in cement production to improve the subsequent hardening of concrete. The elevated temperature during the production of cement may lead to the formation of unwanted hemihydrate

$\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$ according to reaction,



The $\Delta_f H^\ominus$ of $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}(\text{s})$, $\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}(\text{s})$, $\text{H}_2\text{O}(\text{g})$ are $-2021.0 \text{ kJ mol}^{-1}$, $-1575.0 \text{ kJ mol}^{-1}$ and $-241.8 \text{ kJ mol}^{-1}$ respectively. The respective values of their standard entropies are 194.0, 130.0 and $188.0 \text{ J K}^{-1} \text{ mol}^{-1}$. The values of $R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1} = 0.0831 \text{ L bar mol}^{-1} \text{ K}^{-1}$

Answer the following questions on the basis of above information

249. Heat change occurring during conversion of 1 kg of $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}(\text{s})$ (molar mass 172 g mol^{-1}) of $\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}(\text{s})$ is equal to

- a) 484 kJ mol^{-1} b) 400 kJ c) $-484.0 \text{ kJ mol}^{-1}$ d) -1000 kJ

Paragraph for Question Nos. 250 to - 250

A sample of ideal gas undergoes isothermal expansion in a reversible manner from volume V_1 to volume V_2 . The initial pressure is P_1 and the final pressure is P_2 . The same sample is then allowed to undergo reversible expansion under adiabatic conditions from volume V_1 to V_2 . The initial pressure being same but final pressure is P_2

250. The work of expansion in adiabatic process (w_{adi}) is related to work of expansion in isothermal process (w_{iso}) as

- a) $w_{adi} = w_{iso}$ b) $w_{adi} < w_{iso}$ c) $w_{adi} = 2w_{iso}$ d) $w_{adi} > w_{iso}$

Paragraph for Question Nos. 251 to - 251

Free energy, $G = H - TS$, is a state function that indicates whether a reaction is spontaneous or non-spontaneous. If you think of TS as the part of the system's energy that is disordered already, then $(H - TS)$ is the part of the system's energy that is still ordered and therefore free to cause spontaneous change by becoming disordered

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

From the second law of thermodynamics, a reaction is spontaneous if $\Delta_{\text{total}}S$ is positive, non-spontaneous if $\Delta_{\text{total}}S$ is negative, and at equilibrium if $\Delta_{\text{total}}S$ is zero.

Since, $-T \Delta S = \Delta G$ and since ΔG and ΔS have opposite signs, we can restate the thermodynamic criterion for the spontaneity of a reaction carried out at a constant temperature and pressure

If $\Delta G < 0$, the reaction is spontaneous

If $\Delta G > 0$, the reaction is non-spontaneous

If $\Delta G = 0$, the reaction is at equilibrium

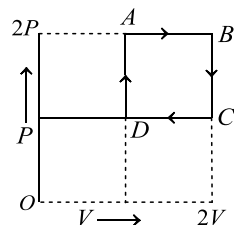
Read the above paragraph carefully and answer the following questions based on the above comprehension

251. For the spontaneity of a reaction, which statement is true?

- a) $\Delta G = +ve$; $\Delta H = +ve$
- b) $\Delta H = +ve$; $\Delta S = -ve$
- c) $\Delta G = -ve$; $\Delta S = -ve$
- d) $\Delta H = -ve$; $\Delta S = +ve$

Paragraph for Question Nos. 252 to - 252

The state of a mole of an ideal gas changed from state A at pressure $2P$ and volume V follows four different processes and finally returns to initial state A reversibly as shown below in the graph. By interpreting the graph, answer the following questions



252. Which is the kind of process followed from state A to state B ?

- a) Isochoric expansion
- b) Isobaric expansion
- c) Isothermal reversible expansion
- d) Isothermal irreversible compression

Paragraph for Question Nos. 253 to - 253

The second law of thermodynamics is a fundamental law of science. In this problem, we consider the thermodynamics of an ideal gas, phase transition, and chemical equilibrium

Three moles of CO_2 gas expands isothermally (in thermal contact with the surroundings; temperature = 15°C) against a fixed external pressure of 1.00 bar. The initial and final volumes of the gas are 10.0 L and 30.0 L, respectively

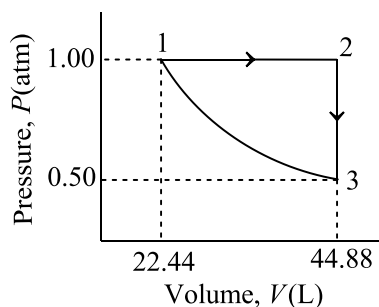
253. Select the correct order of the entropy change

- a) $\Delta_{\text{sys}}S > 0, \Delta_{\text{surr}}S = 0$
- b) $\Delta_{\text{sys}}S < 0, \Delta_{\text{surr}}S > 0$
- c) $\Delta_{\text{sys}}S > 0, \Delta_{\text{surr}}S < 0$

$$d) \Delta_{\text{sys}}S > 0, \Delta_{\text{surr}}S = 0$$

Paragraph for Question Nos. 254 to - 254

A sample consisting of 1 mol of a mono-atomic perfect gas ($C_V = \frac{3}{2}R$) is taken through the cycle as shown



254. Temperature at points (1), (2), and (3), respectively is

- a) 273 K, 546 K, 273 K b) 546 K, 273K, 273K c) 273K, 273K, 273K d) 546K, 546K, 273K

Paragraph for Question Nos. 255 to - 255

Chemical reactions are invariably associated with the transfer of energy either in the form of heat or light. In the laboratory, heat changes in physical and chemical processes are measured with an instrument call calorimeter. Heat change in the process is calculated as

$$q = ms \Delta T \quad s = \text{Specific heat}$$

$$= c\Delta T \quad c = \text{Heat capacity}$$

Heat of reaction at constant volume is measured using bomb calorimeter

$$q_V = \Delta U = \text{Internal energy change}$$

Heat of reaction at constant pressure is measured using simple or water calorimeter

$$q_P = \Delta H$$

$$q_P = q_V + P \Delta V$$

$$\Delta H = \Delta U + \Delta nRT$$

255. Match **List I** with **List II** and select the answer from the given codes

List I		List II	
A	$C(s) + O_2(g) \rightarrow CO_2(g)$	p	$\Delta H = \Delta U + RT$
B	$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$	q	$\Delta H = \Delta U$
C	$NH_4HS(s) \rightarrow NH_3(g) + H_2S(g)$	r	$\Delta H = \Delta U - 2RT$
D	$PCl_5(g) \rightarrow PCl_3(g) + Cl_2(g)$	s	$\Delta H = \Delta U + 2RT$
E	$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$	t	$\Delta H = \Delta U - RT$

a) A → 1, B → 2, C → 3, D → 4, E → 5

c) A → 1, B → 3, C → 4, D → 2, E → 5

b) A → 5, B → 2, C → 3, D → 4, E → 1

d) A → 2, B → 3, C → 4, D → 1, E → 5

Paragraph for Question Nos. 256 to - 256

Bond energies can be obtained by using the following relation:

$$\Delta H(\text{reaction}) = \sum \text{Bond energy of bonds, broken in the reactants} - \sum \text{Bond energy of bonds, formed in the products}$$

Bond energy depends on three factors:

- Greater is the bond length, lesser is the bond energy
- Bond energy increases with the bond multiplicity
- Bond energy increases with the electronegativity difference between the bonding atoms

256. Arrange N – H, O – H and F – H bonds in the decreasing order of bond energy

- F – H > O – H > N – H
- N – H > O – H > F – H
- O – H > N – H > F – H
- F – H > N – H > O – H

Paragraph for Question Nos. 257 to - 257

A change in the free energy of a system at constant temperature and pressure will be:

$$\Delta_{\text{sys}}G = \Delta_{\text{sys}}H - T\Delta_{\text{sys}}S$$

At constant temperature and pressure

$$\Delta_{\text{sys}}G < 0 \text{ (spontaneous)}$$

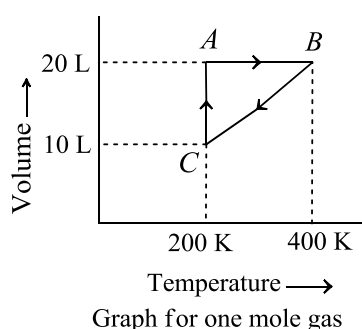
$$\Delta_{\text{sys}}G = 0 \text{ (equilibrium)}$$

$$\Delta_{\text{sys}}G > 0 \text{ (non-spontaneous)}$$

257. The free energy for a reaction having $\Delta H = 31400 \text{ cal}$, $\Delta S = 32 \text{ cal K}^{-1}\text{mol}^{-1}$ at 1000°C is

- 9336 cal
- 7006 cal
- 2936 cal
- +9006 cal

Paragraph for Question Nos. 258 to - 258



258. Process A → B represents

- Isobaric
- Isochoric
- Isothermal
- Isoentropic

Paragraph for Question Nos. 259 to - 259

The thermodynamic property that measures the extent of molecular disorder is called entropy. Entropy change of phase transformation can be calculated using Trouton's formula ($\Delta S = \Delta H/T$). In the reversible adiabatic process, however, ΔS will be zero. The rise in temperature in isobaric and isochoric conditions is found to

increase the randomness or entropy of the system

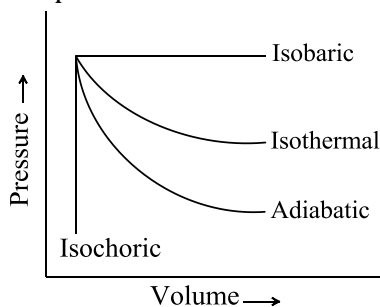
$$\Delta S = 2.303 \log(T_1/T_2)$$

259. The entropy change in an adiabatic process is

- a) Zero b) Positive c) Negative d) Remains true

Paragraph for Question Nos. 260 to - 261

The pressure-volume of various thermodynamic processes is shown in graphs:



Work is the mole of transference of energy. It has been observed that reversible work done by the system is the maximum obtainable work

$$w_{rev} > w_{irr}$$

The works of isothermal and adiabatic processes are different from each other

$$w_{isothermal\ reversible} = 2.303nRT \log_{10} \left(\frac{V_2}{V_1} \right)$$

$$= 2.303nRT \log_{10} \left(\frac{P_2}{P_1} \right)$$

$$w_{adiabatic\ reversible} = C_V(T_1 - T_2)$$

260. If w_1, w_2, w_3 and w_4 are work done in isothermal, adiabatic, isobaric, and isochoric reversible processes, respectively then the correct sequence (for expansion) would be

- a) $w_1 < w_2 < w_3 < w_4$ b) $w_3 = w_2 = w_1 = w_4$ c) $w_3 < w_2 < w_4 < w_1$ d) $w_3 > w_1 > w_2 > w_4$

Paragraph for Question Nos. 261 to - 261

The change in internal energy (U) can be brought about in two ways:

- (i) Either by allowing the heat to flow into the system or out of the system.
- (ii) By doing work on the system or the work done by the system.

Using the symbol q to represent heat transferred to system and using work done by the system $-W$, we can represent the internal energy change of a system, Δu , as

$$q = \Delta U + (-W) \quad (\text{First law of thermodynamics})$$

If the reaction is carried out in a closed container with constant volume, so that $\Delta V = 0$.

Hence, $q_v = \Delta U$

On the other hand, if a reaction is carried out in open vessel that keeps the pressure constant and allows the volume of the system to change freely. In such case, $\Delta V \neq 0$ and $-W = P \cdot \Delta V$.

Hence, $q_p = \Delta U + P\Delta V$

Also $q_p = q_v + \Delta n_g RT$

As reactions carried out at constant pressure are so common, the heat change for such process is given a special symbol ΔH , called the enthalpy change of the reaction. The enthalpy (H) of the system is the name given to the quantity $(U + PV)$.

261. In which of the following cases ΔH and ΔU are not equal to each other?
- The reaction involves no gaseous reactant and product
 - The number of moles of gaseous reactants and gaseous product is not equal to each other
 - The number of moles of gaseous reactants and gaseous products is equal to each other
 - The process is carried out in closed vessel

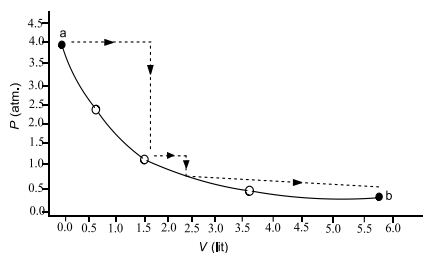
Paragraph for Question Nos. 262 to - 262

Chemical reactions are usually exothermic or endothermic. A balanced thermochemical equation involving physical states of reactants and products expresses the chemical changes as well as heat of reaction. Heat changes are usually expressed in terms of ΔH (at constant P) or ΔU (at constant V). The heats of reactions varies with physical state of reactants and products, conditions of constant pressure or volume and temperature. Heats of combustion and heat of neutralization, heat of condensation are always exothermic. Standard heat enthalpy of a compound is its heat of formation at 1 atm P and 25°C.

262. The heat energy released during neutralization of 1 eq. of NaOH and with 1 eq. of CH_3COOH and 1 eq. of NaOH are respectively :
- 16.4 kcal, –12.0 kcal
 - 12.0 kcal, –10 kcal
 - 13.7 kcal in both
 - 12.0 kcal, –16.4 kcal

Integer Answer Type

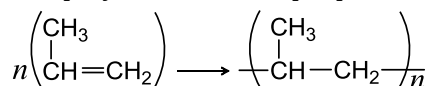
263. A reaction becomes spontaneous only at 500 K. If ΔH at 500 K is 3.0 kJ, the change in entropy at 500 K.
264. Standard molar heat enthalpy for converting carbon from graphite and diamond form into vapour state directly are 725 and 713 kJ respectively. The standard enthalpy change for polymorphism of 1g carbon $C_{G \rightarrow D}$ is ..
265. A certain amount of gas ($P = 5 \text{ atm}, V = 2\text{L}, T = 500\text{K}$) in state A is compressed to state B ($P = 2 \text{ atm}, T = 100 \text{ K}, V = ?$). The final volume of state B in litre is:
266. The combustion of 5.0 g of coke raised the temperature of 1.0 g of water from 10°C to 55°C. If specific heat of water of 1 cal/g, the calorific value of coke is ...
267. At 0°C, if enthalpy of fusion of ice is 1365 kcal/mol. The molar entropy change for melting of ice at 0°C is...kcal.
268. ' a ' moles of Na reacts with sufficient water in an open vessel at 300 K. The work done by the liberated gas H_2 is equivalent to 4988.4 J. What is a ?
269. Heat of solution of anhydrous CuSO_4 is –15.9 kcal and heat of hydration of anhydrous CuSO_4 is – 18.9 kcal. The heat of solution of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ is
270. A reversible reaction is carried out at 500 K, where its equilibrium constant is unity. If ΔH° at 500 K is 4.0 kJ, the value of ΔS° is equal to ...
271. The enthalpy changes of some processes are given below .
- $\alpha - D - \text{glucose}_{(s)} + \text{H}_2\text{O} \rightarrow \alpha - D - \text{glucose}_{(aq)}$; Heat of dissolution = 10.84 kJ
- $\beta - D - \text{glucose}_{(s)} + \text{H}_2\text{O} \rightarrow \beta - D - \text{glucose}_{(aq)}$; Heat of dissolution = 4.68 kJ
- $\alpha - D - \text{glucose}_{(aq)} + \text{H}_2\text{O} \rightarrow \beta - D - \text{glucose}_{(aq)}$; Heat of mutarotation = 1.16 kJ
- The ΔH° for $\alpha - D - \text{glucose} \rightarrow \beta - D - \text{glucose}$ is
272. One mole of an ideal gas is taken from a to b along two paths denoted by the solid and dashed line as shown in graph below. If the work done along the solid line is w_s



and that along the dotted line path is w_d , then the integer closest to the ratio w_d/w_s is...

273. Heat of combustion of a fuel mol. wt 86 is 688 kJ/mole. The calorific value of fuel is... ..

274. The polymerization of propene to linear polypropene is represented by the reaction



Where n has large integral value, the average enthalpies of bond dissociation for (C = C) and (C - C) at 298 K are +590 and +331 kJ mol⁻¹, respectively. The enthalpy of polymerization is -360 kJ mol⁻¹. Find the value of n

275. $\Delta_f H^\ominus$ of hypothetical MgCl is -125 kJ mol⁻¹ and for MgCl₂ is -642 kJ mol⁻¹. The enthalpy of disproportionation of MgCl is -49x. Find the value of x

276. One mole of monoatomic ideal gas at $P = 2$ bar and $T = 273$ K is compressed to 4 bar pressure following a reversible path obeying $P/V = \text{constant}$. Assume $C_v = 12.5$ J mol⁻¹K⁻¹. The value of $\frac{\Delta U}{W}$ for this process is minus.

277. The enthalpy of transition of crystalline boron to amorphous boron at 1500°C is 0.4 kcal mol⁻¹. Assuming at. wt. of boron 10, the change in enthalpy of transition 50 g boron from crystalline to amorphous form is...

278. Bond dissociation energy of XY, X₂, and Y₂ (all diatomic molecules) are in the ratio of 1:1:0.5 and ΔH_f of XY is -100 kJ mol⁻¹. The bond dissociation energy of X₂ is 100x. Find the value of x

279. Amongst the following, the total number of reactions/processes in which the entropy increases are:

1. $2\text{NaHCO}_3(\text{s}) \rightarrow \text{Na}_2\text{CO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$
2. A liquid crystalline into a solid
3. Temperature of crystalline solid is raised from zero K to 100 K
4. Hard boiling of an egg
5. Devitrification of glass
6. Stretching of a rubber band
7. Desalination of water
8. $\text{NH}_3(\text{g}, 10 \text{ atm}) \rightarrow \text{NH}_3(\text{g}, 1 \text{ atm})$

280. Out of the following properties how many are path functions?

Heat enthalpy, internal energy, Temperature, Work, Heat, Specific heat

281. If heat of formation of CaCl₂ and NaCl are 191 and 97.5 kcal, the heat of reaction for $\text{CaCl}_2 + 2\text{Na} \rightarrow 2\text{NaCl} + \text{Ca}$ is ...

282. The lattice energy of solid KCl is 181 kcal mol⁻¹ and the enthalpy of solution of KCl in H₂O is 1.0 kcal mol⁻¹.

If the hydration enthalpies of K[⊕] and Cl[⊖] ions are in the ratio of 2:1 then the enthalpy of hydration of K[⊕] is -20x kcal mol⁻¹. Find the values of x

283. A heated iron block at 127°C loses 300 J of heat to the surroundings which are at a temperature of 27°. This process is 0.05x J K⁻¹. Find the value of x

284. Amongst the following, the total number of physical properties which are extensive are:

- | | | |
|---|--------------|---------------------|
| a. Density | b. Viscosity | c. Surface tension |
| d. Dipole moment | e. Volume | f. Refraction index |
| g. ΔG h. ΔH i. ΔU j. ΔS | | |

285. Amongst the following the total number of intensive physical properties

- | | | |
|---|--------------|---------------------|
| a. Density | b. Viscosity | c. Surface tension |
| d. Dipole moment | e. Volume | f. Refraction index |
| g. ΔG h. ΔH i. ΔU j. ΔS | | |

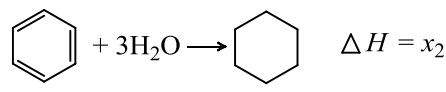
286. $\Delta_f H^\ominus$ of Cyclohexene (*l*) and benzene at 25°C is -156 and $+46$ kJ mol^{-1} , respectively. $\Delta_{\text{hydrogenation}} H^\ominus$ of cyclohexene (*l*) at 25°C is -119 kJ mol^{-1}
 Resonance energy of benzene is found to be $-38x$ kJ mol^{-1} . Find the value of x
287. In the process:
 $\text{H}_2\text{O}(\text{s}, -10^\circ\text{C}, 1 \text{ atm}) \rightarrow \text{H}_2\text{O}(\text{l}, 10^\circ\text{C}, 1 \text{ atm})$
 C_p for ice = $9 \text{ cal deg}^{-1} \text{ mol}^{-1}$, C_p for $\text{H}_2\text{O} = 18 \text{ cal deg}^{-1} \text{ mol}^{-1}$. Latent heat of fusion of ice = $1440 \text{ cal mol}^{-1}$ at 0°C . The entropy change for the above process is $6.258 \text{ cal deg}^{-1} \text{ mol}^{-1}$
 Give the total number of steps in which the third law of thermodynamics is used
288. How much of the following are intensive properties? Vapour pressure, Molarity, Refractive index, Dielectric constant, Osmotic pressure, Molarity, Specific gravity, Molar volume
289. The difference of $\Delta H - \Delta U$ for the given reaction is nRT . The value of n is ...
 $\text{C}_6\text{H}_{12}\text{O}_6(\text{s}) + 6\text{O}_2(\text{g}) \rightarrow 6\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l})$
290. If ΔG and ΔS for the reaction $A(\text{g}) + B(\text{g}) \rightarrow P(\text{g})$ at 300 K are -600 cal and -10 cal/K . The ΔU for the reaction in kcal is...
291. 10 mL of dissolution of a strong acid (HA) on mixing with 10 mL of strong alkali (BOH) at the same temperature shows a temperature rise of 4°C . if 50 mL of same acid are mixed with 50 mL of same alkali; the temperature rise will be ... $^\circ\text{C}$. Assume all the heat produced is used up in increasing temperature of mixture only.
292. ΔH_f^\ominus of Fe_2O_3 and Al_2O_3 are -189 kJ mol^{-1} and -405 kJ mol^{-1} respectively. How much heat (in kJ) is given out during reaction of 1g Al according to $2\text{Al} + \text{Fe}_2\text{O}_3 \rightarrow \text{Al}_2\text{O}_3 + 2\text{Fe}$?
293. A certain number of moles of gas is allowed to heat from 300 K to 500 K at constant P . By doing so gas is expanded to do work on boundaries equivalent to 9.9768 kJ. How many moles of gas were used?
294. The enthalpy change involved in the oxidation of glucose is $-2880 \text{ kJ mol}^{-1}$ 25% of this energy is available for muscular work. If 80 kJ of muscular work is needed to walk one km, what is the maximum distance that a person will walk after eating 120g glucose.
295. For a gas the numerical value (in cal) of $\left[\left(\frac{\partial H}{\partial T}\right)_P - \left(\frac{\partial U}{\partial T}\right)_V\right]$ is equal to ...

: ANSWER KEY :

1) c	2) b	3) d	4) c	a, b, d
5) b	6) b	7) c	8) b	65) b,c,d 66) a, b, c, d 67) a,b,c
9) d	10) b	11) a	12) b	68) a, b, d
13) b	14) c	15) b	16) d	69) a,b 70) a,c,d 71) a,b,c 72)
17) c	18) b	19) c	20) d	b,d
21) d	22) a	23) b	24) c	73) b,d 74) a,b,c,d 75) a, b, d 76)
25) a	26) c	27) b	28) b	a,b,c
29) b	30) c	31) a	32) d	77) a, b 78) a,b,c 79) c, d 80)
33) a	34) c	35) b	36) c	b, d
37) c	38) b	39) a	40) d	81) a,b,c 82) a,b,c,d 83) a,b,c 84)
41) a	42) d	43) b	44) b	b, c, d
45) b	46) a	47) c	48) a	85) a, c 86) b,c,d 87) a,b,c,d 88)
49) a	50) c	51) d	52) c	a, d
53) d	54) b	55) c	56) b	89) a,b 90) b,c,d 91) a,c 92)
57) c	58) a	59) c	60) d	a,b,c,d
61) d	62) c	63) b	64) b	93) a 94) a,c 1) a 2) d
65) c	66) c	1) a,b,c,d	2) a,b,c,d	3) d 4) c
a,c	3) c, d, e	4) a, b, c, d	5) a	6) c 7) a 8) e
5) a,c,d	6) a,c	7) b, d	8) b	9) c 10) d 11) a 12) a
a, b, d				13) a 14) d 15) a 16) c
9) a,b,c	10) b,d	11) a,b,c,d	12) a	17) c 18) c 19) b 20) c
b,c				21) a 22) b 23) c 24) b
13) a,b,c	14) b,d	15) a,b	16) a	25) c 26) d 27) c 28) c
a, c				29) e 30) c 31) c 32) b
17) a, b, c, d	18) a,c	19) c,d	20) a,b	33) c 34) c 35) c 36) c
21) b,d	22) a,d	23) b,d	24) a,b,d	37) e 38) e 39) b 40) c
a,b,d				41) e 42) b 43) b 44) c
25) a,b,d	26) a,b,c	27) a,b	28) a,b,d	45) a 46) b 47) a 48) a
a, b, c				49) c 50) d 51) a 52) b
29) c, d, e	30) c, d	31) a, b, c, d	32) a, b, c	53) a 54) e 1) b 2) c
33) b, d	34) a, d	35) a, b, c	36) a,b,c	3) c 4) b
a,b,c				5) c 6) b 7) a 8) d
37) a, b	38) b, c	39) a,b	40) a,b	9) a 10) a 11) a 12) b
a, c				13) a 14) b 15) d 16) b
41) b, c, d	42) a, c, d	43) d	44) a,b,c	17) c 18) b 19) b 20) b
a, d				21) a 22) a 23) b 24) c
45) b,c	46) a, b, c	47) a, b	48) a,b,c	25) a 26) a 27) a 28) c
c, d				29) a 1) a 2) c 3) b
49) a,d	50) a,b	51) b,c,d	52) a,b,c	4) b
b,d				5) d 6) a 7) b 8) d
53) a,b	54) b,c,d	55) a,c	56) a,b,c	9) b 10) c 11) a 12) d
b,c,d				13) a 14) a 15) b 16) a
57) a,c	58) a,b,c,d	59) b,c,d	60) a,b,c	17) d 18) b 19) a 1) 6
b,c,d				2) 1 3) 1 4) 9
61) a,b,c	62) b,c	63) b,c,d	64) a,b,c	5) 5 6) 4 7) 3 8) 8
				9) 5 10) 2 11) 8 12) 5
				13) 8 14) 3 15) 2 16) 4

17)	4	18)	3	19)	4	20)	6	29)	4	30)	4	31)	6	32)	6
21)	5	22)	5	23)	5	24)	4	33)	2						
25)	2	26)	7	27)	6	28)	3								

: HINTS AND SOLUTIONS :

- 1 **(c)**
 $w = P_{\text{ex}}(\Delta V)$
- 2 **(b)**
 $W = P(\Delta V)$
 $PV = RT$
 $P(V + \Delta V) = R(T + 1)$
 $\therefore P\Delta V = R$
- 3 **(d)**
 $\eta = \frac{T_2 - T_1}{T_2}$; T_1 = temperature of sink
 T_2 = temperature of the source
 $0.25 = \frac{T_2 - T_1}{T_2}$
- 7 **(c)**
 $q_P = q_V$
- 8 **(b)**
 $\text{C}_2\text{H}_5\text{SH} + \frac{9}{2}\text{O}_2(\text{g})$
 $\rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l}) + \text{SO}_2(\text{g})$
- 10 **(b)**
 $\Delta H > \Delta E$
 $\text{C}_4\text{H}_{10} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
- 12 **(b)**
 $\text{CO}_2(\text{g}), \text{H}_2\text{O}(\text{l}),$ and $\text{SO}_2(\text{g})$
- 13 **(b)**
 At equilibrium, $\Delta S_{\text{Total}} = 0$
 $\therefore S_P^\ominus - S_R^\ominus = 0$
- 14 **(c)**
 $PV = RT$ at temperature T for one mole
 $P(V + \Delta V) = R(T + 1)$ at temperature $(T + 1)$ for one mol
 $\therefore P\Delta V = R$
- 15 **(b)**
 $\Delta H = \Delta U + \Delta nRT$
 In (b), $\Delta n = 0$
 $\therefore \Delta H = \Delta U$
- 17 **(c)**
 $\Delta H = \Delta U + \Delta nRT$
 $\Delta n = n_P - n_R$
 $\therefore \Delta n = 2 - \frac{5}{2} = -\frac{1}{2}$
 $\therefore \Delta H = \Delta U - \frac{1}{2}RT$
 $\therefore \Delta U = \Delta H + \frac{1}{2}RT$
 Or $\therefore \Delta U > \Delta H$
- 18 **(b)**
 $H = U + PV$
- 20 **(d)**
 $\Delta S = \frac{q_{\text{rev}}}{T}$
- 23 **(b)**
 $\text{CO}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{CO}(\text{g}) + \text{H}_2\text{O}(\text{g})$
 $\Delta H^\ominus = \left\{ \begin{array}{l} [\Delta_f H^\ominus \text{CO}(\text{g}) + \Delta_f H^\ominus \text{H}_2\text{O}(\text{g})] \\ -\Delta_f H^\ominus \text{CO}_2(\text{g}) \end{array} \right\}$
 $= [(-110.5) + (-241.8 - (-393.5 + 0))]$
 $= +41.2 \text{ kJ}$
 $\Delta_f H^\ominus \text{H}_2(\text{g}) = 0$
 $\Delta_f H^\ominus = 0$ in elementary state
- 25 **(a)**
 $2\text{C}_6\text{H}_6(\text{l}) + 15\text{O}_2 \rightarrow 12\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l})$
 $\Delta H = \Delta E + \Delta n_g RT$
 $\Delta H - \Delta E = \Delta n_g RT$
 $\Delta n_g = 12 - 15 = -3$
 $\Delta H - \Delta E = -3 \times 8.314 \times 298 = -7432.716 \text{ J}$
- 26 **(c)**
 $\text{CH}_2 = \text{CH}_2 + \text{H}_2 \rightarrow \text{CH}_3 - \text{CH}_3 \Delta H = x_1$
 $\Delta H = x_2$
 $\text{RE} = 3x_1 - x_2$
- 27 **(b)**
 $\Delta G = \Delta H - T \Delta S$
- 28 **(b)**
 By definition $C_{P,m} = \frac{dq_p}{dT}$
 For $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{O}(\text{s})$,
 Temperature does not change if some heat is given to the system. Hence
 $C_{P,m} = \frac{+\text{ve}}{\text{Zero}} = \infty$
- 30 **(c)**
 $\Delta U = q - w = q - P \Delta V$
 $= 200 - 2 \times 10^5 \times 500 \times 10^{-6}$
 $= 100 \text{ J}$
- 31 **(a)**
 Dissociation of CaCO_3 required energy
- 32 **(d)**
 Entropy of system depends upon Pressure, volume, and temperature
- 33 **(a)**
 $TV^{\gamma-1} = \text{constant}$
 $\frac{T}{T_{\text{final}}} = \left(\frac{V_2}{V_1}\right)^{\gamma-1}$

$$\frac{T}{T_{\text{final}}} = \left(\frac{2}{1}\right)^{(5/3-1)} = 2^{(2/3)}$$

$$\frac{T}{T_{\text{final}}} = \frac{T}{2^{(2/3)}}$$

34 (c)

$$\Delta H = (\text{BE})_{\text{reactant}} - (\text{BE})_{\text{products}}$$

[But all the species must be in gaseous state. In product, $[\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)]\Delta H$ must be added

Hence, $\text{H}_2(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{H}_2\text{O}(l)$

$$\Delta H = \left[(\text{BE})_{\text{H-H}} + \frac{1}{2}(\text{BE})_{\text{O=O}} \right]$$

$$= [(\Delta H)_{\text{vap}} + 2(\text{BE})_{\text{O-H}}]$$

$$= x_1 + \frac{x_2}{2} - [x_4 + 2x_3]$$

$$= x_1 + \frac{x_2}{2} - x_4 - 2x_3$$

35 (b)

Evaporation of water required heat energy to proceed the reaction

36 (c)

In thermodynamics, a process is called reversible when the surroundings are always in equilibrium with the system

37 (c)

Combustion of one mole of reactants gives products in standard state or most stable state

41 (a)

$$T_2 = T_1 \left(\frac{V_1}{V_2}\right)^{\gamma-1} = 150\text{K}$$

$$w = -C_V\Delta T = -C_V(T_2 - T_1)$$

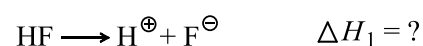
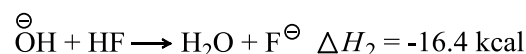
$$= -3 \times 2 \times (150 - 300)$$

$$= 900 \text{ cal}$$

42 (d)

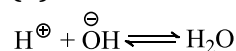
In enthalpy of formation, reactants and products must be in most stable standard state

43 (b)



$$\Delta H_3 = \Delta H_2 - \Delta H_1 = -16.4 - (-13.4) = -3.0 \text{ kcal}$$

45 (b)



$$\Delta H = -13.7 \text{ kcal}$$

$$\Delta H^{\ominus} = \Delta_f H^{\ominus}(\text{H}_2\text{O}) - [\Delta_f H^{\ominus}(\text{H}^{\oplus}) + \Delta_f H^{\ominus}(\text{OH}^{\ominus})]$$

$$= -68 - (0 - 13.7) = -54.3 \text{ kcal mol}^{-1}$$

$$[\Delta_f H^{\ominus}(\text{H}^{\oplus}) = 0 \text{ (convention)}]$$

46 (a)

$$\Delta H = \Delta U + \Delta nRT$$

$$\Delta n = n_p - n_R = 2 - 1 = 1$$

$$\therefore \Delta H = \Delta U + (1)RT$$

$$\therefore \Delta H > \Delta U$$

47 (c)

$$\Delta H = 2.303 R \left[\frac{1}{T_1} - \frac{1}{T_2} \right] \log \frac{P_2}{P_1}$$

$$460.6 = 2.303 \times 2 \left[\frac{1}{50} - \frac{1}{T_2} \right] \log 10$$

$$\therefore T_2 = 100 \text{ K}$$

48 (a)

Oxalic acid has two ionisable H^{\oplus} . Hence, expected heat of neutralization, if it behaves as a strong acid would have been

$$= -13.7 \times 2 = -27.4 \text{ kcal mol}^{-1}$$

But experimental value = $-25.4 \text{ kcal mol}^{-1}$

$$\therefore \text{Heat of ionization} = 2.0 \text{ kcal mol}^{-1}$$

49 (a)

$$G_{(\text{ice})} = G_{(\text{water})} = 0$$

50 (c)

Activation energy: E_a is the energy that must be possessed by the molecules in excess to the average energy at a given temperature to enter a chemical reaction

Relation between activation energy and enthalpy of a reversible reaction

If the reaction is endothermic in forward direction, then

$$E_{a(\text{backward})} = E_{a(\text{forward})} + \Delta H$$

If the reaction is exothermic in forward direction

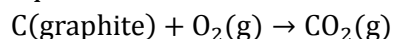
$$E_{a(\text{backward})} = E_{a(\text{forward})} + \Delta H$$

For an endothermic reaction, where

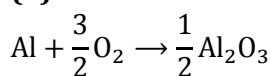
ΔH represents the enthalpy of the reaction in kJ mol^{-1} , the minimum value for the energy of activation will be slightly more than ΔH

51 (d)

Standard molar enthalpy of formation CO_2 and the standard molar enthalpy of combustion of carbon (graphite) refer to the same chemical equation:



53 (d)



$$27 \text{ g} \qquad \qquad \qquad 51 \text{ g}$$

$$\Delta H = -837.8 \text{ kJ mol}^{-1}$$

1. 0.624 mol of Al = 837.8×0.624 on combustion gave = 523 kJ

Hence false

2. Formation of 0.624 mol of Al_2O_3 gave

$$= 837.8 \times 2 \times 0.624$$

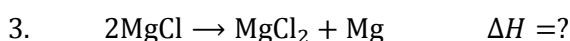
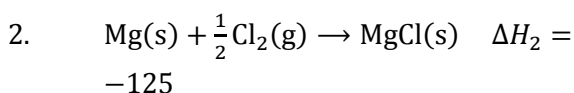
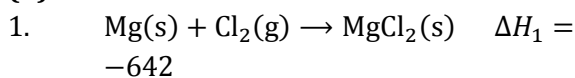
$$= 1045 \text{ kJ}$$

Hence, false

3. 0.312 mol of Al on combustion gave = 261 kJ. Hence, false

4. Formation of 0.150 mol of Al_2O_3 gave = 251.3 kJ. Hence, true

54 (b)



$$\Delta H = \Delta H_1 - 2\Delta H_2$$

$$= -642 - 2 \times (-125) = -392 \text{ kJ mol}^{-1}$$

55 (c)

$$\Delta S_{\text{Total}} = 0$$

56 (b)

$$\Delta H = \Delta U + P\Delta V$$

57 (c)

$$\Delta H = \Delta U + \Delta(PV)$$

$$= \Delta U + \Delta P \Delta V$$

58 (a)

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

All exothermic reactions are spontaneous, hence HF will attack SiO_2

59 (c)

Standard state implies to 1 atm pressure and 298 K temperature

61 (d)

$$\Delta H = \Delta U + \Delta nRT$$

$$\Delta n = 0 = n_P - n_R$$

$$\therefore \Delta H = \Delta U$$

62 (c)

$$V_1 (\text{volume of 1 g } \text{H}_2) = 11.2 \text{ L at STP}$$

$$V_2 (\text{volume of 1 g } \text{H}_2) = 22.4 \text{ L}$$

$$\therefore W = P \Delta V = 11.2 \text{ L atm}$$

63 (b)

$$\Delta H_1 > \Delta H_2$$

64 (b)

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

$$= 2.1 \times 2 \times 0.002 \times 300 = 3.3 \text{ kcal}$$

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

$$= 3.3 \times -300 \times (0.02) = -2.7 \text{ kcal}$$

65 (c)

On comparison:

$$\Delta S = - \left[\frac{d(\Delta G)}{dT} \right]$$

$$\Delta S = - \frac{d(-nFE)}{dT} = nF \left(\frac{dE}{dT} \right) \left[\text{Here } E \text{ is the } E_{mf} \text{ of cell} \right]$$

$$\therefore \left(\frac{dE}{dT} \right) = \frac{\Delta S}{nF}$$

66 (c)

In an ideal gas, there is no force of attraction between molecules. Hence, no heat is lost in the expansion

67 (a,b,c,d)

Born-Haber cycle

72 (a,c)

Mass-independent properties are extensive

75 (a,b,c)

For a cyclic process, $dU = 0$

$$\therefore q = \Delta U + (-w) \Rightarrow q = -w$$

Also, w = Area covered by sphere

$$= \pi r^2 = \pi \times \left[\frac{(V_2 - V_1)}{2} \right]^2$$

$$= \frac{\pi \times (20)^2}{2^2}$$

$$= 100\pi = 100 \times 3.14 = 314 \text{ J}$$

76 (b,d)

The quantities that do not depend upon the quantity or mass of a substance are called intensive properties

77 (a,b,c,d)

In all these process heat of reaction changes

78 (b,c)

$$\Delta H = \Delta U + \Delta nRT, \text{ where } \Delta n = n_P -$$

$$n_R \text{ (} n = \text{number of moles)}$$

81 (a,b)

In isothermal process, $\Delta T = 0$ and $\Delta U = \Delta H = 0$

84 (a,c)

$$\left(\frac{\partial H}{\partial T} \right)_P - \left(\frac{\partial U}{\partial T} \right)_V = R$$

87 (b,d)

$\Delta G = -ve$ for spontaneous process and $\Delta G^\ominus \neq 0$ at equilibrium. Statement (b) represents Carnot theorem

88 (a,d)

$$pV^\gamma = \text{constant ad } pV^{\gamma-1} = \text{constant}$$

Helium and neon are monoatomic gases

($\gamma = 1.67$), while oxygen is diatomic gas ($\gamma = 1.4$).

- Therefore, final temperature and pressure of both helium and neon are same
- 89 **(b,d)**
Intensive property does not depend upon the mass of substance
- 90 **(a,b,d)**
In conversion of solid ice to liquid and in rusting, the entropy increases
- 91 **(a,b,d)**
In all these conversion, the degree of randomness decreases
- 92 **(a,b,c)**
Mathematical relation
- 93 **(a,b)**
 $\Delta U = 0$ for cyclic and isothermal process
- 102 **(a,b,c)**
Heat is not a state function, while temperature is a state function. Heat flowing into the system is positive and heat flowing out of the system is negative
- 109 **(d)**
Use Hess' law
- 115 **(a,d)**
Temperature and pressure are intensive properties
- 116 **(a,b)**
For spontaneous process, $\Delta G = -ve$ and $\Delta S_{\text{Total}} = +ve$
- 117 **(b,c,d)**
 $\text{H}_2\text{O(s)}$ has more ordered arrangement. Also, $\Delta S_2^\ominus = S^\ominus_{\text{H}_2\text{O}(l)} - S^\ominus_{\text{H}_2\text{O}(s)}$
 $\Delta S_1^\ominus = S^\ominus_{\text{H}_2\text{O}(l)} - S^\ominus_{\text{H}_2\text{O}(s)}$
 $\therefore S^\ominus_{\text{H}_2\text{O}(v)}$ is maximum and thus $\Delta S^\ominus_2 \gg \Delta S^\ominus_1$.
- 118 **(b,d)**
The first law of thermodynamics states that $\Delta U = q + w$, i.e., the internal energy of system is equal to the sum of heat and work
- 120 **(b,c,d)**
All of these three are endothermic processes
- 121 **(a,c)**
 γ for mono-atomic is 1.66
 $\therefore C_V = \frac{3R}{2}, C_P = \frac{5}{2}R$
 $C_{VM} = \frac{2 \times \frac{3R}{2} + 2 \times \frac{3R}{2}}{4} \therefore C_V = 3 \text{ cal}$
 $C_{PM} = \frac{2 \times \frac{5R}{2} + 2 \times \frac{5R}{2}}{4} \therefore C_P = 5 \text{ cal}$
- 122 **(b,c,d)**
In spontaneous process, $\Delta G = -ve$, $\Delta H =$
- $-ve$, and $\Delta S = +ve$
- 125 **(b,c,d)**
 $\Delta G^\ominus = 0$ at equilibrium under standard state.
Also, at equilibrium, $\Delta G = 0$
 $\therefore \Delta H^\ominus - T\Delta S^\ominus = 0$
Also $\Delta G^\ominus = 2.303RT \ln K$
 $\therefore K = 1$
- 127 **(a,b,c)**
 $\Delta G^\ominus = -RT \ln K_p$ or $K_p = e^{-\Delta G^\ominus/RT}$
- 128 **(b,c)**
 $\text{C(graphite)} \rightarrow \text{C(g)}$. The process is sublimation as well conversion of crystalline allotropic form (graphite) to amorphous form (gas carbon)
- 129 **(b,c,d)**
Mass-independent properties are intensive
- 131 **(b,c,d)**
Entropy is greater in the liquid than in the solid state because molecules are held tightly in the solid. No. of molecules are decreasing on the product side in option (c) and (d)
- 133 **(a,b,c)**
In (a), volume change is negligible while in (b) and (c), there is no volume change
- 135 **(a,b)**
In adiabatic process, $q = 0$. Hence, in expansion process, temperature decreases and hence internal energy also decreases
- 137 **(a,b,c)**
The sum of two extensive properties is always extensive while the ratio of extensive properties is intensive and the derivation is also intensive
- 138 **(b,d)**
Let V_1 be the initial volume of dry air at NTP
We know, $p_1V_1 = p_2V_2$
 $p_2 = \frac{p_1V_1}{V_2} = \frac{1 \times V_1}{4V_1} = 0.25 \text{ atm}$
During isothermal expansion, the temperature remains the same throughout. Hence, final temperature will be 273 K (0°C)
- 139 **(b,d)**
 $\Delta G = \Delta H - T\Delta S$ for spontaneous process ΔG must be negative. When $\Delta H = +ve$ or endothermic then to make $\Delta G = -ve$. Temperature should be high
- 140 **(a,b,c,d)**
 $\frac{C_P}{C_V} = 1.4 = \gamma$
 $C_P - C_V = R$
 $\therefore C_V = \frac{R}{\gamma-1}$ and similarly $C_P - \frac{C_P}{\gamma} = R$
 $\therefore C_V = \frac{2}{0.4} = 5 \text{ cal} \therefore C_P = \frac{\gamma \cdot R}{(\gamma-1)}$

$$\text{Also, } C_V = \frac{2}{0.4 \times 32} = 0.156 \text{ cal}$$

142 **(a,b,c)**

The correct order is $w_3 > w_1 > w_2 > w_4$

144 **(a,b,c)**

The multiplication of two intensive properties is an intensive property and their ratio and derivative is also an intensive property

147 **(a,b,c)**

Hess' law can be used for all these process or reaction

149 **(a,b,c)**

In open system, heat and matter can be exchanged between system and surrounding

152 **(b,c,d)**

In endothermic reaction, heat is absorbed by the system from the surroundings

153 **(a,b,c,d)**

The heat of neutralization for strong acid and base is equal to $-57.3 \text{ kJ mol}^{-1}$ or $-13.7 \text{ kcal mol}^{-1}$

158 **(a,b,c,d)**

These are mathematically established facts

159 **(a)**

Work is not a state function, but it is a path function

161 **(a)**

Entropy is randomness. So, more the complexity of molecule, more will be its absolute entropy. That's why benzene shows more entropy than CO_2 which has also. More entropy than He

162 **(d)**

No doubt HF is weak acid but the higher values are due to extensive hydration of F^- ion being smallest anion (only H^- is smaller than F^-).

163 **(d)**

Work and heat are not state functions whereas $\Delta U = q + W$ is state function.

164 **(c)**

Reason is correct explanation for statement.

165 **(a)**

This is an application of Hess's law and not an experimental proof. Hess's law has been verified theoretically.

166 **(c)**

Reason is correct explanation for statement.

$$\Delta S_{\text{vap}} = S_{\text{vapour}} - S_{\text{liq}} \text{ and } \Delta S_{\text{fus}} = S_{(L)} - S_{(S)}$$

$\Delta S_{\text{vap}} \gg \Delta S_{\text{fusion}}$. Gaseous state has more disordered system.

167 **(a)**

For cyclic process, $\Delta U = 0$. Energy is state function

169 **(c)**

This is Trouton's rule derived from exp. data.

170 **(d)**

Both are facts.

171 **(a)**

In adiabatic process, $q = 0$ hence is expansion process temperature of the system decreases

172 **(a)**

Heat of neutralization for strong acid and strong base is equal to $-57.3 \text{ kJ mol}^{-1}$. HClO_4 and HCl both are strong acid

173 **(a)**

$1/2 \text{ N}_2 + 1/2 \text{ O}_2 \rightarrow \text{NO}$; $\Delta H = +ve$. It is due to high bond energy of N_2 . Note this equation does not represent heat of combustion of N_2 (partial combustion).

174 **(d)**

Assertion is incorrect

176 **(c)**

Explanation is correct reason for statement.

177 **(c)**

Explanation is correct reason for statement.

178 **(c)**

Explanation represents II law of thermodynamics.

179 **(b)**

The heat of neutralization of strong acid and strong base is $-57.3 \text{ kJ mol}^{-1}$

180 **(c)**

Explanation is correct reason for statement.

181 **(a)**

Zeroth law of temperature can also be summarized as two objects at different temperature in thermal contact with each other tend to move towards the same temperature

182 **(b)**

Phase transition occurs at constant P and thus,

change is referred as ΔH .

183 (c)

$$\text{Heat of formation of H atom} = \Delta H = \frac{e_{\text{H-H}}}{2}$$

184 (b)

The reason is also an explanation for the name zeroth law.

185 (c)

$$\Delta S = \frac{q_{rev}}{T}, \text{ thus, unit is } \text{JK}^{-1}\text{mol}^{-1}.$$

186 (d)

Endothermic reaction requires heat energy to proceed in forward direction

187 (c)

Reason is correct explanation for statement.

188 (c)

Explanation is correct reason for statement.

189 (e)

Enthalpy of element in most stable state is zero.
Diamond is not a stable of carbon

190 (c)

Exothermic reaction may also be accompanied with increase of entropy

191 (c)

Spontaneous reaction may be negative also

193 (c)

Reason is correct explanation is correct.

194 (c)

Explanation is correct reason for statement.

195 (c)

The entropy of formation cannot be zero

196 (c)

For a spontaneous $\Delta G = -ve$ and $\Delta G = \Delta H - T\Delta S$, if $\Delta H = +ve$, ΔG will be negative when $T\Delta S > \Delta H$.

199 (b)

$$C_p - C_v = R \text{ (Ideal gas) and } \left(\frac{dE}{dV}\right)_T = 0 \text{ (ideal gas)}$$

Both are correct but reason is not the correct explanation

200 (c)

Calorific value is defined as the heat given out by burning 1 g fuel.

$$\text{Calorific value for } \text{C}_2\text{H}_6 = \frac{-341.1}{30} = -11.37 \text{ kcal/g}$$

$$\text{Calorific value for } \text{C}_2\text{H}_2 = \frac{-310}{26} = -11.92 \text{ kcal/g}$$

202 (b)

On increasing temperature, the endothermic reaction becomes spontaneous and entropy also increases

203 (b)

Dissociation of water is $\text{H}_2\text{O} \rightleftharpoons \text{H}^{\oplus} + \text{OH}^{\ominus}$ is reverse of the heat of neutralization and the value of heat is equal but sign is reverse

204 (c)

For a spontaneous process, ΔG must be negative

205 (a)

The properties whose magnitude depends upon the quantity of matter present in the system are called extensive properties eg, internal energy

206 (b)

Graphite is most stable state of carbon, hence its energy is lower than that of diamond and entropy of graphite is also lower than that of diamond

208 (a)

It is fact that absolute values of internal energy of substance can not be determined. It is also true that to determine exact values of constituent energies of the substance is impossible

209 (c)

Explanation is correct reason for statement.

210 (d)

$q = \Delta U - W$. For isothermal expansion $\Delta U = 0$. Also $W = 0$, because $W = P \times \Delta V$ and $P = 0$; Also volume occupied by molecules is zero for ideal gas. Thus, both are assertion and reason are correct, but this is not an explanation for assertion.

211 (a)

Heat of formation of water is $\text{H}_2 + 1/2 \text{O}_2 \rightarrow \text{H}_2\text{O}$; ΔH_{for} .

212 (b)

The mass and volume depend upon the quantity of matter so these are extensive properties while ratio of mass to its volume does not depend upon the quantity of matter so this ratio is an extensive property

213 (a)

Kirchhoff's equation is $\Delta H_2 - \Delta H_1 = \Delta C_p(T_2 - T_1)$.

216 (c)

(a→p,q) Specific heat = Amount of heat required to raise temperature of 1 mol substance

$$(b \rightarrow q, p) C_p \text{ or } C_v = \frac{\Delta H}{\Delta T} \text{ or } \frac{\Delta U}{\Delta T}$$

(c→s)

(d→r)

217 (c)

(a→r) Enthalpy is a states function and is a function of internal energy and pressure and volume

(b→p) Temperature is independent of the bulk matter of the system and hence it is a intrinsic property but it is a state variable

(c→r) Free energy is a state function as it depends on the state of the system as well as it depends on $H, T,$ and S . It is also a function of U, P, V

(d→q) Work is a path function as the amount of work

218 (b)

(a→q) $\Delta_{\text{total}}S > 0$ is according to second law of thermodynamics. It states entropy of universe is continuously increasing due to spontaneous processes taking place in it

(b→r) if $\Delta_{\text{total}}S < 0$, the process is non-spontaneous. $\Delta_{\text{total}}S$ mean $\Delta_{\text{sys}}S + \Delta_{\text{surr}}S > 0$, for a spontaneous process

$$(c \rightarrow s) \Delta H = \Delta U + \Delta nRT$$

(d→p) $-\Delta G = W_{\text{useful}}$, i.e., useful work done by the system

i.e., decrease in free energy is measure of useful work done by the system

219 (c)

(a→s) $\Delta G = \Delta H - T\Delta S$ (Gibbs-Helmholtz reaction)

(b→r)

$$(c \rightarrow q) \Delta H = \Delta U + P\Delta V$$

(d→p)

220 (b)

(a→p,r) In isothermal process, $-w = 2.303RT \log \frac{V_2}{V_1}$ or $2.303RT \log \frac{P_1}{P_2}$

(b→s) In adiabatic process, $-w = \frac{nR}{(\gamma-1)}(T_2 - T_1)$

(c→q) In free expansion, $w = 0$

(d→q) In free expansion, $w = 0$

222 (d)

(a→p,q) At equilibrium $\Delta S = 0$ and for spontaneous process

(b→q) Entropy decreases ($\Delta S < 0$) because in formation of product number of moles decrease

(c→q,r)

(d→q,s)

224 (a)

(a→r) $\Delta G = -ve$, Gibbs free energy is negative and it favours the reaction to occur spontaneously

(b→s) Endothermic reaction $\Delta H = H_p - H_R > 0$

Therefore, $H_p > H_R$

(c→p) Bond enthalpy = $\sum(\text{BE})_R - \sum(\text{BE})_P$

(d→q) For solid and liquid $\Delta H = \Delta U = 0$

225 (a)

(a→s) Hess' law states that enthalpy change in a reaction remains the same whether the reaction takes place in one step or in several steps

(b→r) Combustion reactions are exothermic

$$(c \rightarrow q) \frac{\Delta_{\text{vap}}H}{\text{Boiling point in K}} = 88 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$(d \rightarrow p) 2.303 \log \frac{P_2}{P_1} = \frac{\Delta_{\text{vap}}H}{R} \left(\frac{T_2 - T_1}{T_1 T_2} \right)$$

It is Clausius-Clapeyron equation

227 (a)

(a→p,q) $\Delta G = \Delta H - T\Delta S$ and $\Delta G = -nFE$

(b→p,r) $\Delta G^\ominus = -nFE^\ominus$ or and $\Delta G = -RT \log K$

(c→p)

(d→s) $\Delta S^\ominus = 2.303nR \log \left(\frac{V_2}{V_1}\right)$

228 (b)

(a→t) For ideal gas, $\left(\frac{\partial U}{\partial V}\right)_T = 0$

(b→r) In adiabatic process, $q = 0$. Therefore, from the first law of thermodynamic, $\Delta U = q - w$ and $q = 0$

$\therefore \Delta U = -w$

(c→p) In isothermal process, $\Delta U = 0$, because internal energy is the function of temperature

(d→q) $\Delta G^\ominus = -nFE^\ominus$

(e→s) $\left(\frac{\partial T}{\partial P}\right)_H \neq 0$ for real gases

229 (d)

(a→p,r) Kirchoff's equation $\Delta H = \Delta C_p (T_2 - T_1)$

(b→q,r) $\Delta U = \Delta C_v (T_2 - T_1)$

(c→p) $\frac{\Delta H}{T} = \Delta S$

(d→s) The third law of thermodynamics states that at absolute zero the entropy of a perfectly crystalline substance is zero

231 (c)

(a→s)

(b→r) Kirchoff's equation relate the variation of enthalpy with temperature

(c→p) For ideal gas, $\left(\frac{\partial U}{\partial V}\right)_T = 0$

(d→q) Inversion temperature, $T_i = \frac{2a}{Rb}$

232 (b)

(a→q)

(b→p,r)

(c→p,r)

(d→p,r,s) Kirchoff's relation $\Delta H = C_p \Delta T$

234 (b)

(a→s) $dU = \left(\frac{dU}{dV}\right)_T dV + \left(\frac{dU}{dT}\right)_T dT$

For an ideal gas $\left(\frac{dU}{dV}\right)_T = 0$

$\Rightarrow dU \left(\frac{dU}{dT}\right)_T dT$

But for isothermal expansion/compression of ideal gas, $dT = 0$

$\Rightarrow dU = 0$

$H = U + PV$

$\Rightarrow \Delta H = \Delta U + \Delta(PV); \Delta U = 0$

and $PV = \text{constant}$ for isothermal change

$\Rightarrow \Delta(PV) = 0 \Rightarrow \Delta H = 0$

(b→p,r) $\Delta H = -ve$ for hydrogenation and also for non-ideal solutions with negative deviation

(c→q) For a reversible process

$dq_{(Rev)sys} = -dq_{(Rev)surr}$

For any process

$dS_{univ} = dS_{sys} + dS_{surr}$

$= \frac{-dq_{sys}}{T} + \frac{dq_{surr}}{T}$

$= \frac{dq_{surr}}{T} + \frac{dq_{surr}}{T}$

$\Rightarrow dS_{univ} = 0$

(d→r) $\Delta G_{mix} = \Delta H_{mix} - T\Delta S_{mix}$

$\Delta H_{mix} = +ve$ during mixture

$\Rightarrow \Delta G_{mix} = -ve$

235 (a)

(a→q) $\Delta H = \Delta U + \Delta nRT, \Delta n = n_p - n_R = 0$

$\therefore \Delta H = \Delta U$ ($n = \text{number of moles}$)

(b→q) $\Delta H = \Delta U$ ($\because \Delta n = 0$)

(c→p) $\Delta H = \Delta U + RT$ ($\Delta n = n_p - n_R = 1$)

(d→r,s) $\Delta H = \Delta U - 2RT$ ($\Delta n = -2$)

236 (a)

(a→p) For non spontaneous process, photochemical reaction $\Delta_{\text{sys}}S > 0$

(b→p,q) For spontaneous process, $\Delta G < 0$ or -ve but in photochemical reaction (non-spontaneous), $\Delta G < 0$

(c→r) At equilibrium, $\Delta S = 0$

(d→s) For non spontaneous process, $\Delta G > 0$

237 (b)

(a→q) For exothermic reaction, $\Delta H = -ve$

(b→r) For endothermic reaction, $\Delta H = +ve$

(c→s) For spontaneous process, $\Delta G = -ve$

(d→p) $\Delta H = H_p - H_R$

240 (a)

(a→s) Whenever internal energy of a system decreases, then the temperature always decreases

In an adiabatic expansion work is done by the gas

⇒ w is -ve

⇒ $\Delta U = -ve$

⇒ Temperature decreases

(b→r) For an ideal gas, $\left(\frac{\partial E}{\partial V}\right)_T = 0$

(c→q) Whenever internal energy increases temperature of the system increases

(d→p) During resonance,

$$\Delta H_{\text{Combustion(experimental)}} < \Delta H_{\text{Combustion (calculated)}}$$

In $\text{CH}_2 - \text{CH} = \text{CH} = \text{CH}_2$ where resonance takes place

241 (a)

(a→q) To initial reaction, activation energy must be invested

(b→p)

(c→r) In endothermic reaction (ΔH or $\Delta U = +ve$) heat absorbed by system

(d→s) In exothermic reaction (ΔH or $\Delta U = -ve$)

heat releases out from system to surroundings

242 (c)

(a→p,r) In isothermal process ($\Delta T = 0$), $-w = q$

$$\text{Or } -w = 2.303 RT \log \frac{V_2}{V_1} \text{ or } 2.303 RT \log \frac{P_1}{P_2}$$

(b→q) Adiabatic process ($q = 0$) ($PV^\gamma = \text{constant}$)

(c→q) Mathematical relationship

(d→s) $w = -P_{\text{ext}}(V_2 - V_1)$

244 (a)

$$pV = nRT$$

$$2 \times 8 = 2 \times 0.080 \times T$$

$$W_{\text{irr}} = -p_{\text{ext}}(V_2 - V_1)$$

$$= -20 \left(\frac{nRT}{p_2} - \frac{nRT}{p_1} \right) = 144 \text{ bar-L}$$

245 (c)

$$C_V = 0.075 \times 40 = 3 \text{ cal}$$

$$C_p - C_V = 4$$

$$C_p - 3 = 2$$

$$C_p = 5 \text{ cal}$$

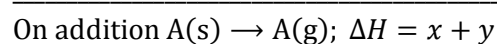
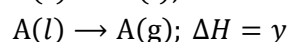
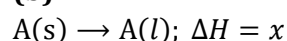
$$\gamma = \frac{C_p}{C_v} = \frac{5}{3} = 1.66$$

Thus, gas is monatomic

246 (b)

$\Delta G = \Delta H - T \Delta S$, at high temperature $T \Delta S$ factor dominates of ΔH and hence ΔG becomes negative and reaction occurs spontaneously

247 (b)



248 (d)

According to the first law of thermodynamics

$$\Delta U = q - w$$

In isothermal process, $\Delta U = 0$

$$\therefore q = w$$

$$\text{Or } w = nRT_1 \ln \frac{V_2}{V_1}$$

249 (a)

$$\Delta H^\ominus = \Delta H_P^\ominus - \Delta H_R^\ominus; \text{ (for 1 mol)}$$

$$= \left[-1575.0 \text{ kJ mol}^{-1} - \frac{3}{2} \times 241.8 \right]$$

$$-[-2021.0 \text{ kJ mol}^{-1}]$$

$$= +83.3 \text{ kJ mol}^{-1}$$

For 1 kg $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$

$$\text{Number of moles} = \frac{1000}{172} = 5.81$$

$$\begin{aligned} \therefore \text{Heat change for 5.81 mol of CaSO}_4 \cdot 2\text{H}_2\text{O} \\ = 5.81 \times 83.3 \text{ kJ mol}^{-1} \\ = 484 \text{ kJ mol}^{-1} \end{aligned}$$

250 (b)

Work in reversible isothermal expansion is greater than work done in adiabatic expansion

251 (d)

$\Delta H = -ve$ and $\Delta S = +ve$, both favour the process

252 (b)

Expansion from state A to state B occurs at constant pressure (isobaric expansion) = $2P$

253 (c)

In expansion of CO_2 gas molecule, entropy of system $\Delta_{\text{sys}}S$ increases or $\Delta_{\text{sys}}S > 0$ and entropy of surrounding decreases $\Delta_{\text{surr}}S < 0$

256 (a)

Fluorine is more electron-negative than oxygen and oxygen is more electro-negative than nitrogen and hence bond energy between F – H in greater than O – H which is greater than N – H

257 (a)

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

258 (b)

Isochoric process ($\Delta V = 0$)

259 (a)

In adiabatic process, $q = 0$

$$\text{and } \Delta S = \frac{q}{T}$$

$$\therefore \Delta S = 0$$

260 (d)

$w = \text{Work done} = \text{Area under curve}$

261 (b)

If $\Delta n_g = 0$, then only $\Delta H = \Delta U$.

262 (a)

Heat released during neutralization of weak acid < -13.7 in HF, hydration of F^- ion is responsible for higher value.

274 (5)

Energy released = Energy due to formation of two single bonds

$$= 2 \times 331 = 662 \text{ kJ mol}^{-1} \text{ of propene}$$

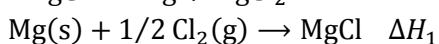
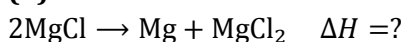
$$\Delta H \text{ polymerization/mol} = 590 - 662 =$$

$$-72 \text{ kJ mol}^{-1}$$

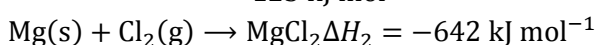
$$\Delta H \text{ polymerisation} = -72 \times n = -360$$

$$n = 5$$

275 (8)



$$= -125 \text{ kJ mol}^{-1}$$

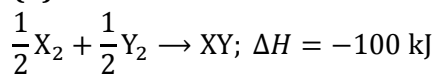


$$\begin{aligned} \Delta H &= \Delta H_2 - 2\Delta H_1 = -642 - (2 \times -125) \\ &= -392 \text{ kJ mol}^{-1} \end{aligned}$$

$$\therefore -49x = -392$$

$$x = 8$$

278 (4)



Let the bond dissociation energy of X_2, Y_2 and XY be: $\frac{a}{2} : a$ (the given ration) kJ mol^{-1} , respectively

$$\therefore \frac{a}{2} + \frac{a}{4} - a = -100$$

$$\therefore a = 400$$

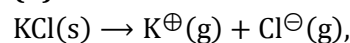
$$100x = 400$$

$$x = 4$$

279 (4)

Entropy increases in (a), (c), (d) and (h)

282 (6)

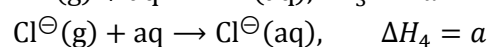


$$\Delta H_1 = 181 \text{ kcal mol}^{-1}$$



$$\Delta H_2 = 1.0 \text{ kcal mol}^{-1}$$

Let the enthalpy of hydration of K^{\oplus} is 2Cl kcal mol^{-1}



$$\therefore \Delta H_3 = -\Delta H_1 + \Delta H_2 - \Delta H_4$$

$$2a = -181 + 1 - a$$

$$3a = -180, a = -60$$

$$\therefore \Delta_{\text{hyd}}H^{\ominus} \text{ of } \text{K}^{\oplus} = 2a = -60 \times 2 = -120$$

$$\therefore -20x = -120$$

$$x = 6$$

283 (5)

$$\Delta_{\text{sys}}S = \frac{q_{\text{sys}}}{T_{\text{sys}}} = -\frac{300}{273 + 127}$$

$$= \frac{-300}{400} = \frac{-3}{4} \text{ J K}^{-1}$$

$$\Delta_{\text{surr}}S = \frac{-q_{\text{sys}}}{T_{\text{surr}}} = -\frac{300}{273 + 27}$$

$$= \frac{300}{300} = +1 \text{ J K}^{-1}$$

$$\Delta_{\text{total}}S \text{ or } \Delta_{\text{universe}}S = \Delta_{\text{sys}}S + \Delta_{\text{surr}}S$$

$$= \frac{-3}{4} + 1 = \frac{1}{4} = 0.25 \text{ J K}^{-1}$$

$$\therefore 0.05x = 0.25$$

$$x = 5$$

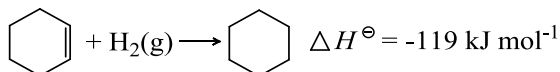
284 (5)

Extensive properties which depends on mass are (e), (g), (h), (i) and (j)

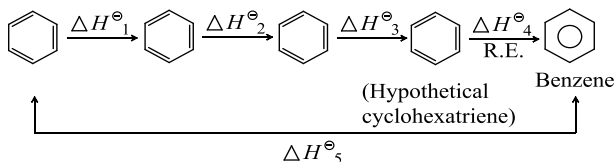
285 (5)

Intensive properties are independent on mass.

(a), (b), (c), (d) and (f)
286 (4)



This shows that the generation of one (C = C) bond in cyclohexane requires 119 kJ mol^{-1} of enthalpy. To calculate RE resonance energy



$$\Delta H_1^\ominus = \Delta H_2^\ominus = \Delta H_3^\ominus = 119 \text{ kJ mol}^{-1}, \Delta H_4^\ominus = \text{RE}$$

$$\Delta H_5^\ominus = \Delta_f H^\ominus(\text{benzene}) - \Delta_f H^\ominus(\text{cyclohexane})$$

$$= 46 - (-156) = 205 \text{ kJ mol}^{-1}$$

∴ From Hess' law

$$\Delta H_4^\ominus = \Delta H_5^\ominus - (\Delta H_1^\ominus + \Delta H_2^\ominus + \Delta H_3^\ominus)$$

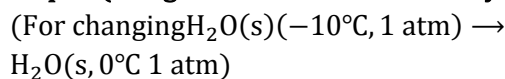
$$= 205 - 3 \times 119 = -152 \text{ kJ mol}^{-1}$$

$$\therefore -38x = -152$$

$$x = 4$$

287 (2)

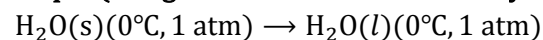
Step 1. (using the third law of thermodynamics):



$$\Delta S_1 = \int_{-10}^0 n \frac{C_p}{T} dT = 1 \times 9 \times 2.3 \times \log \frac{273}{263}$$

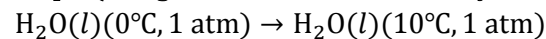
$$= 0.336 \text{ cal deg}^{-1} \text{ mol}^{-1}$$

Step 2 (using the second law of thermodynamics):



$$\Delta S_2 = \frac{q_{\text{rev}}}{T} = \frac{1440}{273} = 5.258 \text{ cal deg}^{-1} \text{ mol}^{-1}$$

Step 3 (using the third law of thermodynamics):



$$\Delta S_3 = \int_0^{10} n \frac{C_p}{T} dT = 1 \times 18 \times 2.3 \times \log \frac{283}{273}$$

$$= 0.647 \text{ cal deg}^{-1} \text{ mol}^{-1}$$

$$\Delta S = \Delta S_1 + \Delta S_2 + \Delta S_3 = 0.336 + 5.258 + 0.647$$

$$= 6.258 \text{ cal deg}^{-1} \text{ mol}^{-1}$$