THERMODYNAMICS

1. Standard enthalpy of vapourisation $\Delta_{vap.}H^{\ominus}$ for water at 100°C is 40.66 kJ mol^{-1}. The internal energy of vapourisation of water at 100°C (in kJ mol^{-1}) is:
a) +43.76 b) +40.66 c) +37.56 d) −43.76
2. The factor $\left(\frac{\partial Q}{\partial T}\right)_P - \left(\frac{\partial Q}{\partial T}\right)_V$ is equal to :
a) γ b) R c) $\frac{R}{M}$ d) ΔnRT
3. Heat of combustion of a substance:
a) Is always positive b) Is always negative
c) Is equal to heat of formation d) Nothing can be said without reaction
4. The heat of formations of $CO(g)$ and $CO_2(g)$ are −26.4 kcal and −94.0 kcal respectively. The heat of combustion of carbon monoxide will be
a) −67.6 kcal b) 36.5 kcal c) −36.5 kcal d) −46.5 kcal
5. Which reaction either endothermic or exothermic characteristics has the greater chance of occurring spontaneously?
a) One in which entropy change is positive
b) One is which entropy change is negative
c) One in which Gibbs energy change is negative
d) One in which equilibrium has been established
6. Net work done by the system in a cyclic process is equal to:
a) Zero b) ΔU c) ΔH d) q
7. A thermodynamic quantity is that:
a) Which is used in thermochemistry
b) Which obeys all the laws of thermodynamics
c) Quantity which depends only on the state of the system
d) Quantity which is used in measuring thermal change
8. The Gibbs energy change for a reversible reaction at equilibrium is:
a) Zero b) Small positive c) Small negative d) Large positive
9. If, $S + O_2 \rightarrow SO_2;\ \Delta H = -298.2\ \text{kJ}\quad \ldots(i)$
$SO_2 + \frac{1}{2}O_2 \rightarrow SO_3;\ \Delta H = -98.7\ \text{kJ}\quad \ldots(ii)$
$SO_3 + H_2O \rightarrow H_2SO_4;\ \Delta H = -130.2\ \text{kJ}\quad \ldots(iii)$
$H_2 + \frac{1}{2}O_2 \rightarrow H_2O;\ \Delta H = -227.3\ \text{kJ}\quad \ldots(iv)$
The enthalpy of formation of H_2SO_4 at 298 K will be:
a) − 754.4 kJ
b) + 320.5 kJ
c) − 650.3 kJ
d) − 433.7 kJ
10. The heat required to raise the temperature of a body by 1 K is called

a) Specific heat b) Thermal capacity c) Water equivalent d) None of these

11. A system absorbs 10 kJ of heat and does 4 kJ of work. The internal energy of the system

a) Increases by 6 kJ b) Decreases by 6 kJ c) Decreases by 14 kJ d) Increases by 14 kJ

12. Which of the following statement is true?

a) ΔH is positive for exothermic reactions

b) ΔH is negative for endothermic reactions

c) The enthalpy of fusion is negative

d) The heat of neutralization of strong acid with strong base is always the same

13. The temperature at which the reaction,

$$Ag_2O(s) \rightarrow 2Ag(s) + \frac{1}{2}O_2(g)$$

Is at equilibrium is …;gives, $\Delta H = 30.5 \text{ kJ mol}^{-1}$ and $\Delta S = 0.066 \text{ kJ K}^{-1}\text{mol}^{-1}$.

a) 462.12 K b) 362.12 K c) 262.12 K d) 562.12 K

14. During an adiabatic process:

a) Pressure is maintained constant

b) Gas is isothermally expanded

c) There is perfect heat insulation

d) The system changes heat with surroundings

15. If $CH_3COOH + OH^- = CH_3COO^- + H_2O + q_1$ and $H^+ + OH^- = H_2O + q_2$,then the enthalpy change for the reaction , $CH_3COOH = CH_3COO^- + H^+$ is equal to :

a) $q_1 + q_2$ b) $q_1 - q_2$ c) $q_2 - q_1$ d) $-q_1 - q_2$

16. Which of the following statements is true? The entropy of the universe

a) Increases and tends towards maximum value b) Decreases and tends to be zero

c) Remains constant d) Decreases and increases with a periodic rate

17. The standard change is Gibbs energy for the reaction,

$H_2O \rightleftharpoons H^+ + OH^-$ at 25°C is:

a) 100 kJ b) −90 kJ c) 90 kJ d) −100 kJ

18. Which is not characteristic of thermochemical equation?

a) It indicates physical state of reactants and products

b) It indicates whether the reaction is exothermic or endothermic

c) It indicates allotrope of reactants if present

d) It indicates whether reaction would occur or not

19. For the reaction,

$H_2(g) + Cl_2(g) \rightarrow 2HCl(g) + x_1$ kJ ……… (i)

$2HCl(g) \rightarrow H_2(g) + Cl_2(g) - x_2$ kJ …… (ii)

Which of the following statement is correct?

a) x_1 and x_2 are numerically equal

b) x_1 and x_2 are numerically different

c) $x_1 - x_2 > 0$

d) $x_1 - x_2 < 0$

20. For which process energy will be absorbed?

a) Separation of an electron from an electron

b) Separation of proton from a neutron

c) Separation of a neutron from neutron

d) Separation of an electron from a neutral atom

21. Which of the following is correct for an ideal gas :

a) $\left(\frac{\partial E}{\partial T}\right)_V = 0$ b) $\left(\frac{\partial E}{\partial P}\right)_T = 0$ c) $\left(\frac{\partial E}{\partial T}\right)_P = 0$ d) All of these

22. An exothermic reaction is one in which the reacting substances

a) Have more energy than the products b) Have less energy than the products

c) Are at a higher temperature than the product d) None of the above

23. Heat energy change during the chemical reaction, $CO + \frac{1}{2} O_2 \longrightarrow CO_2$ is known as :
a) Heat of combustion of CO
b) Latent heat of CO_2
c) Latent heat of vaporisation
d) Heat of formation of CO_2

24. Under the same conditions how many mL of $1M$ KOH and $0.5\ M\ H_2SO_4$ solutions, respectively when mixed for a total volume of 100 mL produce the highest rise in temperature?
a) 67 : 33 b) 33 : 67 c) 40 : 60 d) 50 : 50

25. The first law of thermodynamic is expressed as
a) $q - W = \Delta E$ b) $\Delta E = q - W$ c) $q = \Delta E - W$ d) $W = q + \Delta E$

26. One mole of a non-ideal gas undergoes a change of state (2.0 atm, 3.0 L, 95 K)→ (4.0 atm, 5.0 L, 245 K) with a change in internal energy, $\Delta U = 30.0$ L atm. The change in enthalpy (ΔH) of the process in L atm is:
a) 40.0
b) 42.3
c) 44.0
d) Not defined, because pressure is not constant

27. Which one of the following is an exothermic reaction?
a) $N_2(g) + O_2(g) + 180.8kJ \rightarrow 2NO(g)$ b) $N_2(g) + 3H_2(g) - 92kJ \rightarrow 2NH_3(g)$
c) $C(g) + H_2O \rightarrow CO(g) + H_2(g) - 131.1kJ$ d) $C(graphite) + 2S(s) \rightarrow CS_2(l) - 91.9kJ$

28. If liquids A and B from an ideal solution, the:
a) Enthalpy of mixing is zero
b) Entropy of mixing is zero
c) Free energy of mixing is zero
d) Free energy as well as the entropy of mixing are each zero

29. In which of the following cases entropy decreases?
a) Solid changing to liquid
b) Expansion of a gas
c) Crystals dissolve
d) Polymerisation

30. For the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3 : \Delta H$ is
a) $\Delta E - 2RT$ b) $\Delta E - RT$ c) $\Delta E + RT$ d) $\Delta E + 2RT$

31. When one mole of monoatomic ideal gas at T temperature undergoes adiabatic change under a constant external pressure of 1 atm change in volume is 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) $T - \frac{2}{3 \times 0.0821}$

32. In the combustion of 2.0 g of methane, 25 kcal heat is liberated. Heat of combustion of methane would be
a) 150 kcal b) 200 kcal c) 250 kcal d) 350 kcal

33. 1 mole of an ideal gas at 25°C is subjected to expand reversibly ten times of its initial volume. The change in entropy of expansion is:
a) $19.15\ JK^{-1}mol^{-1}$ b) $16.15\ JK^{-1}mol^{-1}$ c) $22.15\ JK^{-1}mol^{-1}$ d) None of these

34. The heat of formation (ΔH_f) of $H_2O(l)$ is equal to:
a) Zero
b) Molar heat of combustion of $H_2(l)$
c) Molar heat of combustion of $H_2(g)$
d) Sum of heat of formation of $H_2O(g)$ and $O_2(g)$

35. The entropy change for the reaction given below,
$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Is...at 300 K. Standard entropies of $H_2(g), O_2(g)$ and $H_2O(l)$ are 126.6, 201.20 and 68.0 $JK^{-1}mol^{-1}$ respectively.
a) $-318.4\ JK^{-1}mol^{-1}$ b) $318.4\ JK^{-1}mol^{-1}$ c) $31.84\ JK^{-1}mol^{-1}$ d) None of these

36. Heat of combustion ΔH for $C(s), H_2(g)$ and $CH_4(g)$ are $-94, -68$ and -213 kcal/mole then ΔH for $C(s) + 2H_2(g) \rightarrow CH_4(g)$ is :
a) -17 kcal b) -111 kcal c) -170 kcal d) -85 kcal

37. A positive change in enthalpy occurs in :
a) $H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g)$
b) $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
c) $MgCO_3(s) \longrightarrow MgO(s) + CO_2(g)$
d) $H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l)$

38. A gas expands isothermally and reversibly. The work done by the gas is:
a) Zero b) Minimum c) Maximum d) Equal to work done

39. What is Δn for combustion of 1 mole of benzene, when both the reactants and products are gas at 298 K
a) 0 b) 1 c) 0.5 d) 1.5

40. Internal energy and pressure of a gas of unit volume are related as:
a) $P = \frac{2}{3}U$ b) $P = \frac{3}{2}U$ c) $P = \frac{U}{2}$ d) $P = 2U$

41. A reaction accompanied with the absorption of energy is :
a) Burning of a candle b) Rusting of iron c) Electrolysis of water d) Digestion of food

42. The second law of thermodynamics introduced the concept of:
a) Third law of thermodynamics
b) Work
c) Entropy
d) Internal energy

43. The enthalpy change is negative for :
a) $Cl^-(g) + aq \longrightarrow Cl^-(aq)$
b) $Cl(g) \longrightarrow Cl^+(g) + e$
c) $\frac{1}{2}Cl_2(g) \longrightarrow Cl(g)$
d) $Cl_2(l) \longrightarrow Cl_2(g)$

44. Equal volumes of monoatomic and diatomic gases at same initial temperature and pressure are mixed. The ratio of specific heats of the mixture (C_p/C_v) will be
a) 1 b) 2 c) 1.67 d) 1.2

45. If, $C(s) + O_2(g) \rightarrow CO_2(g)$, $\Delta H = -393.5$ kJ and
$CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g), \Delta H = -283.5$ kJ,
then the heat of formation of CO is:
a) -110.0 kJ b) 676.5 kJ c) -676.5 kJ d) 110.5 kJ

46. Hess's law of constant heat summation is an application of :
a) Kirchhoff's law
b) First law of thermodynamics
c) Second law of thermodynamics
d) Entropy

47. The heat of reaction at constant pressure is equal to :
a) $\Sigma U_P - \Sigma U_R$ b) $\Sigma U_R - \Sigma U_P$ c) $\Sigma H_P - \Sigma H_R$ d) $\Sigma H_R - \Sigma H_P$

48. Select the correct limitations of III law of thermodynamics.
a) Glassy solids at zero kelvin has entropy greater than zero
b) Solids having mixture of isotopes do not have entropy zero at zero kelvin

c) Crystals of CO, N_2O, NO, H_2O, etc., do not have zero entropy at zero kelvin
d) All of the above

49. Heat of formation of $H_2O(g)$ at 1 atm and 25°C is – 243 kJ. ΔU for the reaction,
$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g)$ at 25°C is :
a) 241.8 kJ b) -241.8 kJ c) $-$ 243 kJ d) 243 kJ

50. Molar heat capacity of a gas at constant temperature and pressure is :
a) $(3/2)R$
b) $(5/2)R$
c) Infinite
d) Depends upon atomicity of gas

51. If water is formed from H^+ ions and OH^-, the heat of formation of water is :
a) -13.7 kcal b) $+13.7$ kcal c) -63.4 kcal d) $+63.4$ kcal

52. The process of evaporation of a liquid is accompanied by:
a) Increase in enthalpy
b) Decrease in Gibbs energy
c) Increase in entropy
d) All of the above

53. The work done during the process when 1 mole of gas is allowed to expand freely into vacuum is:
a) Zero b) +ve c) −ve d) Either of these

54. The van't Hoff reaction isotherm is:
a) $\Delta G = RT \log_e K_p$ b) $-\Delta G = RT \log_e K_p$ c) $\Delta G = RT^2 \ln K_p$ d) None of these

55. Which species have negative value of specific heat?
a) Ice b) Water c) Vapour d) Saturated vapour

56. The standard heat of formation of sodium ions in aqueous solution from the following data will be:
Heat of solution $NaOH(aq.)$ from $NaOH(s) = -470.7$ kJ
Heat of formation of $OH^-(aq.)$ From $OH^-(s) = -228.8$ kJ
a) $-$ 251.9 kJ b) 241.9 kJ c) $-$ 241.9 kJ d) 251.9 kJ

57. An ideal gas expands in volume from $1 \times 10^3 m^3$ to $1 \times 10^{-2} m^3$ at 300 K against a constant pressure of $1 \times 10^5 Nm^{-2}$. The work done is
a) 270 kJ b) -900 kJ c) -900 J d) 900 kJ

58. For the reaction $3O_2 \longrightarrow 2O_3$; $\Delta H = +ve$, we can say :
a) O_3 is more stable than O_2
b) O_3 is less stable than O_2
c) Both are equally stable
d) Formation of O_3 is exothermic

59. One mole of a gas absorbs 200 J of heat at constant volume. Its temperature rises from 298K To 308K. The change in internal energy is:
a) 200 J b) -200 J c) $200 \times \frac{308}{298}$ J d) $200 \times \frac{298}{308}$ J

60. Which of the following have same units?
(i) work (ii) Heat
(iii) Energy (iv) Entropy
a) (i), (ii) and (iii) b) (i), (ii) and (iv) c) (ii), (iii) and (iv) d) (iii) and (iv)

61. For an ideal gas, the relation between the enthalpy change and internal energy change at constant temperature is given by :
a) $H = U + PV$ b) $\Delta H = \Delta U + \Delta nRT$ c) $\Delta U = \Delta H + P\Delta V$ d) $\Delta H = \Delta G + T\Delta S$

62. Vibrational energy is:
a) Partially potential and partially kinetic b) Only potential c) Only kinetic d) None of the above

63. The relation $\Delta G = \Delta H - T\Delta S$ was given by

a) Boltzmann b) Faraday c) Gibbs-Helmholtz d) Thomson

64. Calculate the free energy change of

$$2\mathrm{CuO}(s) \rightarrow \mathrm{Cu_2O}(s) + \frac{1}{2}\mathrm{O_2(g)}$$

Given, $\Delta H = 145.6$kJ per mol

$\Delta S = 116$ J per mol per K

a) 113.8 kJ per mol b) 221.5 kJ per mol c) 55.4 kJ per mol d) 145.6 kJ per mol

65. The bond dissociation energy of B − F in BF_3 is 646 kJ $\mathrm{mol^{-1}}$, whereas that of C − F in CF_4 is 515 kJ $\mathrm{mol^{-1}}$.

The correct reason for higher B − F bond dissociation energy as compared to that of C − F is :

a) Stronger σ bond between B and F in BF_3 as compared to that between C and F in CF_4

b) Significant $p\pi - p\pi$ interaction between B and F in BF_3 whereas there is no possibility of such interaction between C and F in CF_4

c) Lower degree of $p\pi - p\pi$ interaction between B and F in BF_3 than that between C and F in CF_4

d) Smaller size of B-atom as compared to that of C-atom

66. When water is added to quick lime, the reaction is :

a) Explosive b) Endothermic c) Exothermic d) None of these

67. An engine operating between 150°C and 25°C takes 500 J heat from a higher temperature reservoir if there are no frictional losses, then work done by engine is

a) 165.85 J b) 169.95 J c) 157.75 J d) 147.7 J

68. Least random state of water is:

a) Ice

b) Liquid water

c) Steam

d) All present in same random state

69. Given that standard heat enthalpy of CH_4, C_2H_4 and C_3H_8 are −17.9, 12.5, −24.8 kcal/mol. The ΔH for $CH_4 + C_2H_4 \rightarrow C_3H_8$ is :

a) -55.2 kcal $\mathrm{mol^{-1}}$ b) -30.2 kcal $\mathrm{mol^{-1}}$ c) 55.2 kcal $\mathrm{mol^{-1}}$ d) -19.4 kcal $\mathrm{mol^{-1}}$

70. Absorption of gasses on solid surface is generally exothermic because

a) Enthalpy is positive

b) Entropy decreases

c) Entropy increases

d) Free energy increases

71. Energy equivalent to one erg, one joule and one calorie are in order:

a) 1 erg > 1 J > 1 cal b) 1 erg > 1 cal > 1 J c) 1 cal > 1 J > 1 erg d) 1 J > 1 cal > 1 erg

72. When the change of entropy is greater, then the ability for work is:

a) Maximum b) Minimum c) Medium d) None of these

73. For which change $\Delta H \neq \Delta U$?

a) $H_2 + I_2 \rightleftharpoons 2HI$

b) $HCl + NaOH \rightarrow NaCl + H_2O$

c) $C(s) + O_2(g) \rightleftharpoons CO_2(g)$

d) $N_2 + 3H_2 \rightarrow 2NH_3$

74. Net work done by a system is given by:

a) Decrease in Helmholtz energy (ΔA)

b) Decrease in Gibbs energy (ΔG)

c) Decrease in internal energy

d) Decrease in heat enthalpy

75. Oxidising power of chlorine in aqueous solution can be determined by the parameters indicated below :

$$\frac{1}{2}Cl_2(g) \xrightarrow{\frac{1}{2}\Delta_{diss}H^\ominus} Cl(g) \xrightarrow{\Delta_{eg}H^\ominus} Cl^-(g) \xrightarrow{\Delta_{hyd}H^\ominus} Cl^-(aq)$$

The energy involved in the conversion of

$\frac{1}{2}Cl_2$ (g) to $Cl^-(aq)$

(using the data, $\Delta_{diss}H^\theta_{Cl_2} = 240$ kJ mol^{-1}, $\Delta_{eg}H^\theta_{Cl} = -349$ kJ mol^{-1}, $\Delta_{hyd}H^\theta_{Cl^-} = -381$ kJ mol^{-1}) will be :

a) $+120$ kJ mol^{-1} b) $+152$ kJ mol^{-1} c) -610 kJ mol^{-1} d) -850 kJ mol^{-1}

76. The law of Lavoisier and Laplace is based on :

a) The principle of conservation of energy
b) Equivalence of mechanical and thermal energies
c) The principle of conservation of matter
d) Equivalence of mechanical and chemical energies

77. ΔH for the reaction given below represents,

CO_2 (g) + H_2(g) ⟶ CO(g) + H_2O(g); $\Delta H = 40$ kJ:

a) Heat of formation b) Heat of combustion c) Heat of neutralisation d) Heat of reaction

78. A person requires 2870 kcal of energy to lead normal daily life. If heat of combustion of cane sugar is -1349 kcal, then his daily consumption of sugar is:

a) 728 g b) 0.728 g c) 342 g d) 0.342 g

79. The enthalpy of hydrogenation of cyclohexene is -119.5 kJ mol^{-1}. If resonance energy of benzene is -150.4 kJ mol^{-1}, its enthalpy of hydrogenation would be:

a) -269.9 kJ mol^{-1} b) -358.5 kJ mol^{-1} c) -508.9 kJ mol^{-1} d) -208.1 kJ mol^{-1}

80. The incorrect expression among the following is:

In isothermal process

a) $W_{reversible} = -nRT\ln\frac{V_f}{V_i}$ b) $\ln K = \frac{\Delta H^\circ - T\Delta S^\circ}{RT}$ c) $K = e^{-\Delta G^\circ/RT}$ d) $\frac{\Delta G_{system}}{\Delta S_{total}} = -T$

81. An ideal gas expands in volume from 1×10^{-3} m^3 to 1×10^{-2} m^3 at 300 K against a constant pressure of 1×10^5 Nm^{-2}. The work done is

a) -900 J b) -900 kJ c) 270 kJ d) 900 kJ

82. $C + O_2 \rightarrow CO_2$;

$CO + \frac{1}{2}O_2 \rightarrow CO_2$

Then the heat of formation of CO is

a) $X - Y$ b) $Y - 2X$ c) $X + Y$ d) $2X - Y$

83. The formation water from H_2(g) and O_2(g) is an exothermic reaction because :

a) The chemical energy of H_2(g) and O_2(g) is more than that of water
b) The chemical energy of H_2(g) and O_2(g) is less than that of water
c) Not dependent on energy
d) The temperature of H_2(g) and O_2(g) is more than that of water

84. Which statements are correct?

a) $2.303\log\frac{P_2}{P_1} = \frac{\Delta H_{vap\cdot}}{R}\frac{[T_2 - T_1]}{T_1T_2}$ is Clausius Clapeyron equation
b) $\frac{\Delta H_{vap\cdot}}{\text{Boiling point}} = 88$J mol^{-1}K^{-1} is called Trouton's rule
c) Entropy is a measure of unavailable energy, *i.e.*, unavailable energy = entropy × temperature
d) All of the above

85. The work done in an open vessel at 300 K, when 112 g iron reacts with dil. HCl is:

a) 102 kcal b) 0.6 kcal c) 0.3 kcal d) 0.2 kcal

86. A solution of 500 mL of $2M$ KOH is added to 500 mL of $2M$ HCl and the mixture is well shaken. The rise in temperature T_1 is noted. The experiment is then repeated using 250 mL of each solution and rise in temperature T_2 is again noted. Assume all heat is taken up by the solution:
a) $T_1 = T_2$
b) T_1 is 2 time as larger as T_2
c) T_2 is twice larger as T_1
d) T_1 is 4 time as larger as T_2

87. When 1 mole of gas is heated at constant volume. Temperature is raised from 298 to 308 K. Heat supplied to the gas is 500 J. Then which statement is correct?
a) $q = -W = 500\ \text{J}, \Delta U = 0$
b) $q = \Delta U = 500\ \text{J}, W = 0$
c) $q = W = 500\ \text{J}, \Delta U = 0$
d) $\Delta U = 0, q = W = -500\ \text{J}$

88. The heat of formation is the change in enthalpy accompanying the formation of a substance from its elements at 298 K and 1 atm pressure. Since, the enthalpies of elements in their most stable state are taken to be zero, the heat of formation of compounds is :
a) Always negative
b) Always positive
c) Standard heat enthalpy of that compound
d) Zero

89. If enthalpies of methane and ethane are respectively 320 and 360 cal then the bond energy of $C - C$ bond is
a) 80 cal b) 40 cal c) 60 cal d) 120 cal

90. Which correctly represents the physical significance of Gibbs energy change?
a) $-\Delta G = W_{\text{compression}}$
b) $\Delta G = W_{\text{expansion}}$
c) $\Delta G = -W_{\text{expansion}} = W_{\text{non-expansion}}$
d) $-\Delta G = W_{\text{expansion}}$

91. Heat of neutralisation of which acid-base reaction is − 57.32 kJ for?
a) $CH_3COOH + NaOH$ b) $HCl + NH_4OH$ c) $HCOOH + KOH$ d) $HNO_3 + LiOH$

92. Entropy change of vaporisation at constant pressure is given by:
a) $\Delta S_{(v)} = \frac{\Delta H_v}{T}$ b) $\Delta S_v = \frac{\Delta U_v}{T}$ c) $\Delta S_{(v)} = \frac{\Delta H_v}{\Delta T}$ d) None of these

93. Given, $C + O_2 \rightarrow CO_2 + 94.2\ \text{kcal}$ ....(i)
$H_2 + \frac{1}{2}O_2 \rightarrow H_2O + 68.3\ \text{kcal}$ ....(ii)
$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 210.8\ \text{kcal}$ ....(iii)
The heat of formation in kcal will be :
a) 45.9
b) 47.8
c) − 20.0
d) 47.3

94. The enthalpy of formation of HI is 30.4 kJ. Which statement is false according to this observation?
a) HI is an endothermic compound
b) For the reaction, $H_2(g) + I_2(g) \longrightarrow 2HI(g);\ \Delta H = 60.8\ \text{kJ}$
c) HI is a stable compound
d) HI is an unstable compound

95. Mark the correct statement
a) For a chemical reaction to be feasible, ΔG should be zero

b) Entropy is a measure of order in a system
c) For a chemical reaction to be feasible, ΔG should be positive
d) The total energy of an isolated system is constant

96. The entropy change for vaporisation of liquid water to steam at 100°C is...J $K^{-1}mol^{-1}$. Given that heat of vaporisation is 40.8 kJ mol^{-1}.
a) 109.38 b) 100.38 c) 110.38 d) 120.38

97. Given the bond energies of $N \equiv N, H—H$ and $N—H$ bonds as 945, 436 and 39 kJ mol^{-1} respectively, the enthalpy of the reaction, $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ is :
a) -93 kJ b) 102 kJ c) 90 kJ d) 105 kJ

98. When ammonium chloride is dissolved in water, the solution becomes cold. The change is:
a) Endothermic b) Exothermic c) Super cooling d) None of these

99. The Gibbs energy change and standard Gibbs energy change are equal if reaction quotient is equal to:
a) Zero b) 1 c) > 1 d) < 1

100. The bond energy of H_2 is 104.3 kcal mol^{-1}. It means : (N = Av. no.)
a) 104.3 kcal heat is required to break up N bonds in N molecules of H_2
b) 104.3 kcal heat is required to break up N molecules to $2N$ atoms of H
c) 104.3 kcal heat is evolved during combination of $2N$ atoms of H to from N molecules of H_2
d) Heat of formation of H atom $= \frac{1}{2} \times$ bond energy of H—H

101. For the process, $CO_2(s) \longrightarrow CO_2(g)$:
a) Both ΔH and ΔS are +ve
b) ΔH is −ve and ΔS is +ve
c) ΔH is +ve and ΔS is −ve
d) Both ΔH and ΔS are −ve

102. The process, in which no heat enters or leaves the system is termed as
a) Isochoric b) Isobaric c) Isothermal d) Adiabatic

103. The work done during the expansion of a gas from a volume of 4 dm^3 to 6 dm^3 against a constant external pressure of 3 atm is
a) -6 J b) -608 J c) +304 J d) -304 J

104. Which are correct to express work terms?
a) Work = Capacity factor × Intensity factor; where capacity factor is a measure of extent of work done and intensity factor is a measure of force responsible for work
b) Electrical work $= E \times nF$; E is intensity factor; nF is capacity factor
c) Expansion work $= P \times \Delta V$; P is intensity factor; ΔV is capacity factor
d) All of the above

105. Theenthalpy change for the transition of liquid water to steam is $\Delta H_{vap} = 37.3$ kJ mol^{-1} at 373 K. The entropy change for the process is
a) 132.5 J $mol^{-1}K^{-1}$ b) 100 J $mol^{-1}K^{-1}$ c) 135.3 J $mol^{-1}K^{-1}$ d) 75.5 J $mol^{-1}K^{-1}$

106. Which is not a spontaneous process?
a) Expansion of a gas into vacuum
b) Water flowing down hill
c) Heat flowing from colder body to a hotter body
d) Evaporation of water from clothes during drying

107. In a chemical reaction $\Delta H = 150$ kJ and $\Delta S = 100$ JK^{-1} at 300 K. The ΔG for the reaction is:
a) Zero b) 300 kJ c) 330 kJ d) 120 kJ

108. Enthalpy of $CH_4 + \frac{1}{2}O_2 \rightarrow CH_3OH$ is negative. If enthalpy of combustion of CH_4 and CH_3OH are x and y respectively. Then which relation is correct?
a) $x > y$ b) $x < y$ c) $x = y$ d) $x \geq y$

109. The enthalpy and entropy change for the reaction,
$Br_2(l) + Cl_2(g) \rightarrow 2BrCl(g)$
Are $30\ \text{kJ mol}^{-1}$ and $105\ \text{JK}^{-1}\text{mol}^{-1}$ respectively. The temperature at which the reaction will be in equilibrium is:
a) 450 K
b) 300 K
c) 285.7 K
d) 273 K

110. The gas in a refrigerator causes cooling on expansion because:
a) Work done by the gas is converted into heat
b) Heat of the gas is lost as work is done by the gas
c) The heat is spread over a large space
d) None of the above

111. ΔS is positive for the change:
a) Mixing of two gases
b) Boiling of liquid
c) Melting of solid
d) All of these

112. In a flask, colourless $N_2O_4(g)$ is in equilibrium with brown coloured $NO_2(g)$. At equilibrium when the flask is heated to 100°C, the brown colour deepens and on cooling it becomes coloured. Which statement is incorrect about this observation?
a) The ΔH for the reaction $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ is +ve
b) Paramagnetism increases on cooling
c) The $\Delta H - \Delta U$ at 100°C is equal to 200 cal
d) Dimerisation is reduced on heating

113. A gaseous reaction was carried out, first keeping the volume constant and next keeping the pressure constant. In the second experiment, there was an increase in volume. The heats of reaction were different, because :
a) In the first case the energy was spent to keep the volume constant
b) In the second case energy was spent to expand the gases
c) Specific heats of compressed gases is more
d) Specific heats of rarefied gases is more

114. The maximum work done in expanding 16 g oxygen at 300 K and occupying a volume of 5 dm^3 isothermally until the volume become 25dm^3 is:
a) 2.01×10^3 J
b) $+2.81 \times 10^3$ J
c) 2.01×10^{-3} J
d) $+2.01 \times 10^{-6}$ J

115. The heat of neutralization of strong base and a strong acid is 57 kJ. The heat released when 0.5 mole of HNO_3 solution is added to 0.20 mole of NaOH solution, is :
a) 11.4 kJ
b) 34.7 kJ
c) 23.5 kJ
d) 58.8 kJ

116. For which one of the following equations is $\Delta H^{\circ}_{\text{reaction}}$ equal to ΔH°_f for the product?
a) $N_2(g) + O_3(g) \longrightarrow N_2O_3(g)$
b) $CH_4(g) + 2Cl_2(g) \longrightarrow CH_2Cl_2(l) + 2HCl(g)$
c) $Xe(g) + 2F_2(g) \longrightarrow XeF_4(g)$
d) $2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$

117. If a chemical change is brought about by one or more methods in one or more steps, then the amount of heat absorbed or evolved during the complete course of reaction is same, whichever method was followed. This law is known as :
a) Le-Chatelier's principle
b) Hess's law
c) Joule-Thomson effect
d) Trouton's law

118. When a gas undergoes adiabatic expansion, it gets cooled due to
a) Loss of kinetic energy
b) Fall in temperature
c) Decrease in velocity
d) Energy used in doing work

119. How much heat (in kJ) is evolved when 2.7 g of aluminium reacts with a stoichiometric amount of Fe_2O_3?
$2Al_2 + Fe_2O_3 \rightarrow 2Fe + Al_2O_3;\ \Delta H^o = -852$ kJ

a) 852 kJ b) 426 kJ c) −42.6 kJ d) +42.6 kJ

120. Which reaction is endothermic in nature?
a) $CaCO_3 \rightarrow CaO + CO_2$
b) $C + O_2 \rightarrow CO_2$
c) $NaOH + HCl \rightarrow NaCl + H_2O$
d) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

121. The enthalpy and entropy change for a chemical reaction are -2.5×10^3 cal and 7.4 cal K^{-1} respectively. Predict that nature of reaction at 298 K is:
a) Spontaneous b) Reversible c) Irreversible d) Non-spontaneous

122. If 50 calorie are added to a system and system does work of 30 calorie on surroundings, the change in internal energy of system is:
a) 20 cal b) 50 cal c) 40 cal d) 30 cal

123. If a refrigerator's door is opened then, we get
a) Room heated b) Room cooled
c) More amount of heat is passed out d) No effect on room

124. When a bottle of perfume is opened, odourous molecules mix with air and slowly diffuse throughout the entire room. The incorrect fact about the process is;
a) $\Delta G = -\text{ve}$ b) $\Delta H \simeq 0$ c) $\Delta S = -\text{ve}$ d) $\Delta S = +\text{ve}$

125. Equilibrium constant of a reaction is related to:
a) Standard Gibbs energy change $\Delta G°$
b) Gibbs energy change ΔG
c) Heat enthalpy
d) None of the above

126. 48 g of C (diamond) on complete combustion evolves 1584 kJ of heat. The standard heat of formation of gaseous carbon is 725 kJ/mol. The energies required for the process
(i)C (graphite) → C(gas)
(ii)C (diamond) → C(gas)are:
a) 725, 727 b) 717, 725 c) 725, 723 d) None of these

127. An example of closed system is:
a) Hot water present in an open beaker
b) Some amount of water present in equilibrium with its vapour in a closed and insulated beaker
c) Some amount of hot water enclosed in a closed container which is not insulated
d) None of the above

128. At 1 atm pressure, $\Delta S = 75\ JK^{-1}mol^{-1}; \Delta H = 30\ kJ\ mol^{-1}$. The temperature of the reaction at equilibrium is:
a) 400 K b) 330 K c) 200 K d) 110 K

129. 2.1 g of Fe combines with S evolving 3.77 kJ. The heat of formation of FeS in kJ/mole is
a) −1.79 b) −100.5 c) −3.77 d) None of these

130. In an isochoric process, the increase in internal energy is
a) Equal to the heat absorbed
b) Equal to the heat evolved
c) Equal to the work done
d) Equal to the sum of the heat adsorbed and work done

131. The sublimation energy of $I_2(s)$ is 57.3 kJ/mol and the enthalpy of fusion is 15.5 kJ/mol. The enthalpy of vaporisation of I_2 is
a) 41.8 kJ/mol b) -41.8 kJ/mol c) 72.8 kJ/mol d) -72.8 kJ/mol

132. $\Delta G°$for the reaction $X + Y \rightleftharpoons Z$ is −4.606 kcal. The value of equilibrium constant of the reaction at 227°C is ($R = 2.0$ cal $mol^{-1}K^{-1}$)
a) 100 b) 10 c) 2 d) 0.01

133. The entropy values (in JK^{-1} mol^{-1}) of $H_2(g) = 130.6$ $Cl_2(g) = 223.0$ and $HCl(g) = 186.7$ J $mol^{-1}K^{-1}$ at 298 K and 1 atm pressure, the entropy change for the reaction $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ is

a) +540.3 b) +727.3 c) −166.9 d) +19.8

134. Bond energy of molecule:
- a) Is always negative
- b) Is always positive
- c) Either positive or negative
- d) Depends upon the physical state of the system

135. In which case of mixing of a strong acid and a base each of 1N concentration, temperature increase is highest?
- a) 20 mL acid – 20mL alkali
- b) 10 mL acid – 40mL alkali
- c) 25 mL acid – 25mL alkali
- d) 35 mL acid – 15mL alkali

136. ΔS^o will be highest for the reaction

a) $Ca(s) + \frac{1}{2}O_2(g) \rightarrow CaO(s)$ b) $CaCO_3(g) \rightarrow CaO(s) + CO_2(g)$

c) $C(g) + O_2(g) \rightarrow CO_2(g)$ d) $N_2(g) + O_2(g) \rightarrow 2NO(g)$

137. Bond dissociation enthalpy of H_2, Cl_2 and HCl are 434, 242 and 431 kJ mol^{-1} respectively. Enthalpy of formation of HCl is :

a) – 245 kJ mol^{-1} b) – 93 kJ mol^{-1} c) 245 kJ mol^{-1} d) 93 kJ mol^{-1}

138. Joule-Thomson expansion is

a) Isobaric b) Isoenthalpic c) Isothermal d) None of these

139. The energy absorbed by each molecule (A_2) of a substance is 4.4×10^{-19} J and bond energy per molecule is 4.0×10^{-19} J. The kinetic energy of the molecule per atom will be :

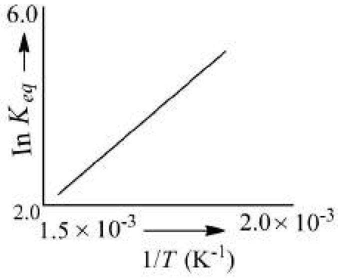
a) 4.0×10^{-20} J b) 2.0×10^{-20} J c) 2.2×10^{-19} J d) 2.0×10^{-19} J

140. Which one of the following is not a state function?

a) Enthalpy b) Entropy c) Work d) Free energy

141. Temperature and heat are:
- a) Extensive properties
- b) Intensive properties
- c) Intensive and extensive properties respectively
- d) Extensive and intensive properties respectively

142. A schematic plot of In K_{eq} *versus* inverse of temperature for a reaction is shown below



The reaction must be

a) Exothermic b) Endothermic

c) One with negligible enthalpy change d) Highly spontaneous at ordinary temperature

143. For a reaction at 25°C, enthalpy and entropy changes are -11.7×10^3J mol^{-1} and −105 J $mol^{-1}K^{-1}$ respectively. What is the Gibbs free energy?

a) 15.05 kJ b) 19.59 kJ c) 2.55 kJ d) 22.55 kJ

144. Which of the following equations correctly represents the standard heat of formation (ΔH_f^o) of methane?

a) C(diamond)+ $2H_2(g) \rightarrow CH_4(g)$ b) C(graphite) + $2H_2(g) \rightarrow CH_4(l)$
c) C(graphite) + $2H(g) \rightarrow CH_4(g)$ d) C(graphite) + $4H \rightarrow CH_4(g)$

145. Combustion of glucose takes place according to the equation
$C_6H_{12}O_6 + CO_2 \rightarrow 6CO_2 + 6H_2O; \Delta H = -72$kcal. How much energy will be required for the production of 1.6 g of glucose (Molecular mass of glucose = 180)?
a) 0.064 kcal b) 0.64 kcal c) 6.4 kcal d) 64 kcal

146. In an irreversible process taking place at constant T and P and in which only pressure volume work is being done, the change in Gibbs free energy (dG) and change in entropy
a) $(dS)_{V,E} < 0 (dG)_{T,P} < 0$ b) $(dS)_{V,E} > 0 (dG)_{T,P} < 0$
c) $(dS)_{V,E} > 0 (dG)_{T,P} = 0$ d) $(dS)_{V,E} = 0 (dG)_{T,P} > 0$

147. Hess's law states that
a) The standard enthalpy of an overall reaction is the sum of the enthalpy changes in individual reactions.
b) Enthalpy of formation of a compound is same as the enthalpy of decomposition of the compound into constituent elements, but with opposite sign.
c) At constant temperature the pressure of a gas is inversely proportional to its volume.
d) The mass of a gas dissolved per litre of a solvent is proportional to the pressure of the gas in equilibrium with the solution.

148. Internal energy is sum of
a) Kinetic energy and potential energy b) All types of energy of the system
c) Energy of internal system d) None of the above

149. Heat given to a system under isochoric process is equal to:
a) W b) q_p c) ΔU d) ΔH

150. All the naturally occurring processes, *i. e.*, spontaneous proceed spontaneously in a direction which leads to:
a) Decrease of free energy
b) Increase of free energy
c) Decrease of entropy
d) Increase of enthalpy

151. Which phenomena cannot be described as oxidation?
a) Oxidation of coal in air
b) Burning of magnesium in nitrogen
c) Reaction of antimony with chlorine
d) Lighting of an electric lamp

152. Heat of formation of H_2O is −188 kJ/mol and H_2O_2 is −286 kJ/mol. The enthalpy change for the reaction
$2H_2O_2 \rightarrow 2H_2O + O_2$ is
a) 196 kJ b) −196 kJ c) 984 kJ d) −984 kJ

153. When 1g atom of carbon is converted into 1g molecule of CO_2, the heat liberated is same:
a) Irrespective of whether the volume is kept constant or pressure is kept constant
b) Irrespective of the temperature at which the reaction is carried out
c) Whether the carbon is in the form of diamond or graphite
d) None of the above

154. For the gas phase reaction,
$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$
Which of the following conditions are correct?
a) $\Delta H > 0$ and $\Delta S > 0$ b) $\Delta H < 0$ and $\Delta S < 0$ c) $\Delta H > 0$ and $\Delta S < 0$ d) $\Delta H = 0$ and $\Delta S < 0$

155. A system provided 50 joule of heat and work done on the system is 10 J. The change in internal energy during the process is:

a) 40 J b) 60 J c) 80 J d) 50 J

156. The correct relationship between free energy change in a reaction and the corresponding equilibrium constant K_c is

a) $\Delta G = RT\ \text{In}K_c$ b) $-\Delta G = RT\ \text{In}K_c$ c) $\Delta G^\circ = RT\ \text{In}K_c$ d) $-\Delta G^\circ = RT\ \text{In}K_c$

157. Identify the correct statement for change of Gibb's energy for a system (ΔG_{system}) at constant temperature and pressure:

a) If $\Delta G_{\text{system}} = 0$, the system is still moving in a particular direction

b) If $\Delta G_{\text{system}} = -$ve, the process is not spontaneous

c) If $\Delta G_{\text{system}} = +$ve, the process is spontaneous

d) If $\Delta G_{\text{system}} = 0$, the system has attained equilibrium

158. The standard heat of combustion of solid boron is :

a) $\Delta_f H^\circ(B_2O_3)$ b) $\frac{1}{2}\Delta_f \Delta H^\circ(B_2O_3)$ c) $-\Delta_f H^0\ (B_2O_3)$ d) $-\frac{1}{2}\Delta_f \Delta H^\circ(B_2O_3)$

159. During an isothermal expansion of an ideal gas its:

a) Enthalpy decreases

b) Internal energy decreases

c) Internal energy increases

d) Internal energy remains constant

160. The exchange of heat energy during chemical reaction at constant temperature and pressure occurs in form of:

a) Free energy b) Internal energy c) Enthalpy d) Bond energy

161. Using the data provided, calculate the multiple bond energy (kJ mol^{-1}) of a $C \equiv C$ bond in C_2H_2. Given that the heat of formation of C_2H_2 is 225 kJ mol^{-1} (take the bond energy of C—H bond as 350 kJ mol^{-1}.) :

$2C(s) \rightarrow 2C(g)$ $\Delta H = 1410$ kJ mol^{-1}

$2C(s) \rightarrow 2C(g)$ $\Delta H = 1410$ kJ mol^{-1}

$H_2(g) \rightarrow 2H(g)$ $\Delta H = 330$ kJ mol^{-1}

a) 1165 b) 837 c) 865 d) 815

162. The molar heat capacity of water at constant pressure is 75 JK^{-1} mol^{-1}. When 1.0 kJ of heat is supplied to 100 g of water which is free to expand the increase in temperature of water is

a) 2.4 K b) 3.6 K c) 4.8 K d) 1.2 K

163. For which of the process, ΔS is negative?

a) $H_2(g) \rightarrow 2H(g)$ b) $2SO_3(g) \rightarrow 2SO_2(g) + O_2(g)$

c) N_2(1 g atom) $\rightarrow N_2$(8 g atom) d) $C_{(\text{diamond})} \rightarrow C_{(\text{graphite})}$

164. Given that, $C(g) + 4H(g) \rightarrow CH_4(g)$; $\Delta H = -166$ kJ. The bond energy C − H will be

a) −416 kJ/mol b) −41.6 kJ/mol c) 832 kJ/mol d) None of these

165. In an isobaric process, the ratio of heat supplied to the system ($d\theta$) and work done by the system (dW) for diatomic gas is

a) 1:1 b) 7:2 c) 7:5 d) 5:7

166. An adiabatic process occurs in

a) Open system b) Closed system

c) Isolated system d) In all the given system

167. The heat of formation of water is exothermic in nature because :

a) $H_2(g)$ and $O_2(g)$ have higher temperature than water

b) $H_2(g)$ and $O_2(g)$ have lower temperature than water

c) $H_2(g)$ and $O_2(g)$ have higher internal energy than water

d) None of the above

168. In an adiabatic process

a) $p.\Delta V = 0$ b) $q = +W$ c) $\Delta E = q$ d) $q = 0$

169. A well stoppered thermos flask containing some ice cubes is an example of:

a) Closed system b) Open system c) Isolated system d) None of these

170. Internal energy is

a) Partly potential and partly kinetic b) Totally kinetic
c) Totally potential d) None of the above

171. The enthalpy of formation of ammonia gas at 298 K is $-46.11\ \text{kJ mol}^{-1}$. The equation to which this value relates is:

a) $\frac{1}{2}N_2(g)+\frac{3}{2}H_2(g)\longrightarrow NH_3(g)$
b) $N(g)+3H(g)\longrightarrow NH_3(g)$
c) $N_2(g)+3H_2(g)\longrightarrow 2NH_3(g)$
d) $\frac{1}{2}N_2(g)+\frac{3}{2}H_2(g)\longrightarrow NH_3(l)$

172. $C_{\text{diamond}}+O_2\rightarrow CO_2;\ \Delta H=-395.3\ \text{kJ/mol}$
$C_{\text{graphite}}+O_2\rightarrow CO_2;\ \Delta H=-393.4\ \text{kJ/mol}$
$C_{\text{graphite}}\rightarrow C_{\text{diamond}};\ \Delta H=?$

a) -3.8 b) -1.9 c) $+3.8$ d) $+1.9$

173. The heat evolved in the combustion of methane is given by the following equations
$CH_4(g)+2O_2(g)\rightarrow CO_2(g)+H_2O(l);\ \Delta H=-890.3\ \text{kJ}$
How many grams of methane would be required to produce 44.15 kJ of heat of combustion?

a) 4 g b) 8 g c) 12 g d) 16 g

174. The calorific value of fat is :

a) Lesser than that of carbohydrate and protein
b) Lesser than that of protein but more than carbohydrates
c) Lesser than that of carbohydrate but more than protein
d) More than that of carbohydrate and protein

175. The standard heat of formation of $NO_2(g)$ and $N_2O_4(g)$ are 8.0 and 2.0 kcal mol^{-1} respectively. The heat of dimerization of NO_2 in kcal is :

a) 10.0 b) -6.0 c) -12.0 d) -14.0

176. Given that,
$3C(s)+2Fe_2O_3(s)\rightarrow 4Fe(s)+3CO_2(g)\Delta H^o=-93657\ \text{cal}$
$C(s)+O_2(g)\rightarrow CO_2(g);\ \Delta H^o=-94050\ \text{cal}$
If both the values of ΔH are at 25°C then calculate$\Delta H^o_{Fe_2O_3}$

a) 16.750 kcal b) -16.750 kcal c) -94.25 kcal d) -393 kcal

177. 16 g oxygen gas expands at STP to occupy double of its original volume. The work done during the process is:

a) 260 kcal b) 180 kcal c) 130 kcal d) 272.8 kcal

178. One mole of an ideal gas at 300K is expanded isothermally from an initial volume of 1 litre to 10 litre. The ΔU for this process is ($R=2\ \text{cal K}^{-1}\text{mol}^{-1}$):

a) 163.7 cal b) 1381.1 cal c) $1.0\ \text{J mol}^{-1}\text{K}^{-1}$ d) $0.1\ \text{J mol}^{-1}\text{K}^{-1}$

179. Gibbs energy G, enthalpy H and entropy S are related by:

a) $G=H+TS$ b) $G=H-TS$ c) $G-TS=H$ d) $S=H-G$

180. In a closed insulated container a liquid is stirred with a paddle to increase the temperature, which of the following is true?

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

181. The ionization energy of solid NaCl is 180 kcal per mole. The dissolution of the solid in water in the form of ions is endothermic to the extent of 1 kcal per mol. If the solvation energies of Na^+ and Cl^- ions are in the ratio 6:5, what is the enthalpy of hydration of sodium ion?

a) -85.6 kcal/mol b) -97.5 kcal/mol c) 82.6 kcal/mol d) $+100$ kcal/mol

182. Enthalpy of solution of NaOH (solid) in water is -41.6 kJ mol^{-1}, when NaOH is dissolved in water, the temperature of water

a) Increases
b) Decreases
c) Does not change
d) Fluctuates indefinitely

183. Mark out the enthalpy of formation of carbon monoxide (CO)

Given, $C(s) + O_2(g) \rightarrow CO(g), \Delta H = -393.3$ kJ/mol

$CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g), \Delta H = -282.8$ kJ/mol

a) 110.5 kJ/mol
b) 676.1 kJ/mol
c) 282.8 kJ/mol
d) 300.0 kJ/mol

184. Heat of neutralization of a strong acid and a strong base is equal to ΔH of :

a) $H^+ + OH^- = H_2O$
b) $H_2O + H^+ = H_3O^+$
c) $2H_2 + O_2 = 2H_2O$
d) $CH_3COOH + NaOH \longrightarrow CH_3COONa + H_2O$

185. The temperature of the system increases during an:

a) Isothermal expansion
b) Adiabatic compression
c) Adiabatic expansion
d) Isothermal compression

186. For an adiabatic expansion of a perfect gas $\Delta P/P$ is equal to:

a) $\Delta V/V$
b) $\gamma \frac{\Delta V}{V}$
c) $-\gamma \frac{\Delta V}{V}$
d) $-\gamma^2 \frac{\Delta V}{V}$

187. The heat evolved during neutralisation is maximum in the reaction of :

a) NH_4OH and CH_3COOH
b) NH_4OH and HCl
c) NaOH and CH_3COOH
d) NaOH and HCl

188. In which reaction ΔS is positive?

a) $H_2O(l) \rightarrow H_2O(s)$
b) $3O_2(g) \rightarrow 2O_3(g)$
c) $H_2O(l) \rightarrow H_2O(g)$
d) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

189. It is impossible to attain the lowest temperature known as zero degree absolute. This is a simple statement of:

a) First law of thermodynamics
b) Second law of thermodynamics
c) Third law of thermodynamics
d) None of the above

190. The total entropy change for a system and its surroundings increases, if the process is:

a) Reversible
b) Irreversible
c) Exothermic
d) Endothermic

191. Heat of formation of compound is defined as :

a) Heat evolved to form one mole of the compound from its elements
b) Heat required to form one molecule of a compound
c) Change in heat content of the system when one molecule of a compound is formed
d) None of the above

192. At absolute zero, the entropy of a perfect crystal is zero. This is ... of thermodynamics.

a) First law
b) Second law
c) Third law
d) None of these

193. Standard state Gibbs free energy changes for the isomerization reaction,

cis-2-pentene $\rightleftharpoons$ *trans*-2-pentene is -3.67 kJ/mol at 400K. If more *trans*-2-pentene is added to the reaction vessel, then:

a) Equilibrium remains unaffected

b) Additional *trans*-2-pentene is formed

c) More *cis*-2-pentene is formed

d) Equilibrium is shifted in forward direction

194. If $S°$ for H_2, Cl_2 and HCl are 0.13, 0.22 and $0.19 kJK^{-1}mol^{-1}$ respectively. The total change in standard entropy for the reaction, $H_2 + Cl_2 \rightarrow 2HCl$ is:

a) 30 $JK^{-1}mol^{-1}$ b) 40 $JK^{-1}mol^{-1}$ c) 60 $JK^{-1}mol^{-1}$ d) 20 $JK^{-1}mol^{-1}$

195. The total internal energy change for a reversible isothermal cycles is

a) Always 100 calories per degree b) Always negative

c) Zero d) Always positive

196. One mole of a gas occupying $3dm^3$ expands against constant external pressure of 1 atm to a volume of 13 dm^3. The work done is:

a) -10 atm dm^3 b) -20 atm dm^3 c) -39 atm dm^3 d) -48 atm dm^3

197. The heat of combustion of methane is $-$ 880 kJ mol^{-1}. If 3.2 g of methan is burnt:

a) $-$ 176 kJ of heat is evolved

b) 176 *kJ* of heat is absorbed

c) 88 *kJ* of heat is evolved

d) None of above

198. Heat of solution is defined as :

a) Heat required to dissolved one mole in excess of water

b) Heat evolved, when one mole is dissolved in excess of water

c) Change in heat content of the system when one mole of the solute is dissolved un excess of water, so that further dilution of solution does not bring any heat change

d) None of the above

199. A mixture of two moles of carbon monoxide and one mole of oxygen, in a closed vessel is ignited to convert the carbon monoxide to carbon dioxide. If ΔH is the enthalpy change and ΔE is the change in internal energy, then

a) $\Delta H > \Delta E$

b) $\Delta H < \Delta E$

c) $\Delta H = \Delta E$

d) The relationship depends on the capacity of the vessel

200. Molar heat capacity is given by :

a) $\frac{dQ}{dT}$ b) $dQ \times dT$ c) $\Sigma Q \frac{1}{dT}$ d) None of these

201. Two mole of an ideal gas is expanded isothermally and reversibly from 1 L to 10 L at 300 K. The enthalpy change (in kJ) for the process is

a) 11.4 b) -11.4 c) 0 d) 4.8

202. A gaseous system changes from state $A(P_1, V_1, T_1)$ to $B(P_2, V_2, T_2)$, B to $C(P_3, V_3, T_3)$ and finally from C to A. The whole process may be called:

a) Reversible process b) Cyclic process c) Isobaric process d) Spontaneous process

203. One mole of ice is converted into water at 273 K. The entropies of $H_2O(s)$ and $H_2O(l)$ are 38.20 and 60.01 J $mol^{-1}K^{-1}$ respectively. The enthalpy change for the conversion is:

a) 59.54 J mol^{-1} b) 5954 J mol^{-1} c) 595.4 J mol^{-1} d) 320.6 J mol^{-1}

204. For a diatomic molecule AB, the electronegativity difference between A and $B = 0.2028\sqrt{\Delta}$. [Where Δ= bond energy of AB Geometric mean of the bond energies of A_2 and B_2] The electronegativities of fluorine and chlorine are 4.0 and 3.0 respectively and the bond energies are of F $-$ F: 38 kcal mol^{-1} and Cl $-$ Cl : 58 kcal mol^{-1}. The bond energy of Cl $-$ F is :

a) ~ 71 kcal/mol b) ~ 61 kcal/mol c) ~ 48 kcal/mol d) ~ 75 kcal/mol

205. Any series of operation so carried out that at the end, the system is back to its state is called

a) Boyle's cycle b) Reversible process c) Adiabatic process d) Cyclic process

206. The heat of neutralisation of a strong acid and a strong alkali is 57.0 kJ mol^{-1}. The heat released when 0.5 mole of HNO_3 solution is mixed with 0.2 mole of KOH is

a) 57.0 kJ b) 11.4 kJ c) 28.5 kJ d) 34.9 kJ

207. The Kirchhoff's equation gives the effect of .....on heat of reaction.

a) Pressure b) Temperature c) Volume d) Molecularity

208. Δn values in $\Delta H = \Delta U + \Delta nRT$ may have:

a) Integer nature b) Fractional value c) Positive or negative d) All of these

209. AB, A_2 and B_2 are diatomic molecules. If the bond enthalpies of A_2, AB and B_2 are in the ratio 1:1:0.5 and the enthalpy of formation of AB from A_2 and B_2 is -100 kJ mol^{-1}, what is the bond enthalpy of A_2?

a) 400 kJ mol^{-1} b) 200 kJ mol^{-1} c) 100 kJ mol^{-1} d) 300 kJ mol^{-1}

210. Which of the following is an intensive property?

a) Temperature b) Viscosity c) Surface tension d) All of these

211. The temperature of the system decreases in an

a) Adiabatic compression b) Isothermal compression

c) Isothermal expansion d) Adiabatic expansion

212. If a refrigerator door is kept open, then we get:

a) Room cooled

b) Room heated

c) More heat is passed out

d) No effect on room

213. The enthalpy of vaporization of a liquid is 30 kJ mol^{-1} and entropy of vaporization is 75 J mol^{-1}. The boiling point of liquid at 1 atm is :

a) 250 K b) 400 K c) 450 K d) 600 K

214. Which is correct about the heat of combustion?

a) The combustion be exothermic in some cases and endothermic in other cases

b) Heat of combustion is always exothermic

c) Its value change with temperature

d) All of the above

215. In an isothermal process

a) $q = 0$ and $\Delta E = 0$ b) $q \neq 0$ and $\Delta E = 0$ c) $q = 0$ and $\Delta E \neq 0$ d) $q \neq 0$ and $\Delta E \neq 0$

216. The enthalpy of combustion of H_2, cyclohexane (C_6H_{10}) and cyclohexane (C_6H_{12}) are $-241, -3800$ and -3920 kJ per mol respectively. Heat of hydrogenation of cyclohexane is

a) 121 kJ/mol b) -121 kJ/mol c) +242 kJ/mol d) -242 kJ/mol

217. For the isothermal expansion of an ideal gas

a) E and H increases b) E increases but H decreases

c) H increases but E decreases d) E and H are unaltered

218. Heat evolved in the reaction, $H_2 + Cl_2 \rightarrow 2HCl$ is 182 kJ. Bond energies of H—H and Cl—Cl are 430 and 242 kJ/mol respectively. The H—Cl bond energy is :

a) 245 kJ mol^{-1} b) 427 kJ mol^{-1} c) 336 kJ mol^{-1} d) 154 kJ mol^{-1}

219. Which is not correct?

a) In an exothermic reaction, the enthalpy of products is less than that of reactants

b) $\Delta H_{fusion} = \Delta H_{sublimation} - \Delta H_{vaporisation}$

c) A reaction for which $\Delta H^\circ < 0$ and $\Delta S^\circ > 0$ is possible at all temperatures

d) ΔH is less than ΔU for the reaction, $C(s) + (1/2)O_2 (g) \rightarrow CO_2(g)$

220. A cylinder of gas is assumed to contain 11.2 kg of butane (C_4H_{10}). If a normal family needs 20000 kJ of energy per day. The cylinder will last (Given that ΔH for combustion of butane is -2658 kJ)

a) 20 days b) 25 days c) 26 days d) 24 days

221. The enthalpy of formation of water from hydrogen and oxygen is -286 kJ mol^{-1}. The enthalpy of decomposition of water into hydrogen and oxygen is:

a) -286 kJ mol^{-1} b) -143 kJ mol^{-1} c) $+286$ kJ mol^{-1} d) $+143$ kJ mol^{-1}

222. An ideal gas is allowed to expand both reversibly and irreversibly in an isolated system. If T_i is the initial temperature and T_f is the final temperature, which of the following statements is correct?

a) $(T_f)_{\text{irrev}} > (T_f)_{\text{rev}}$

b) $T_f > T_i$ for reversible process but $T_f = T_i$ for irreversible process

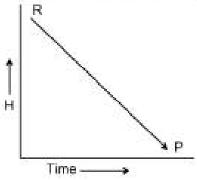
c) $(T_f)_{\text{rev}} = (T_f)_{\text{irrev}}$

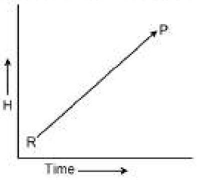
d) $T_f = T_i$ for both reversible and irreversible processes

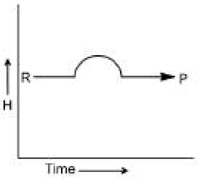
223. Heat of fusion of a molecular solid is :

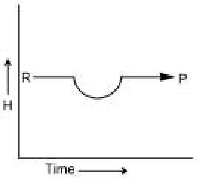
a) Very high b) High c) Low d) None of these

224. Which plot represents for an exothermic reaction?

a)


b)


c)


d)


225. For a spontaneous chemical change the Gibbs energy change is:

a) Positive

b) Negative

c) Zero

d) Depends whether the reaction is exothermic or endothermic

226. An ideal gas undergoing expansion in vacuum shows:

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

227. Select the incorrect statement

a) PV work is usually negligible for solid and liquid

b) For a closed system with $P - V$ work only, an isobaric process that has $q = +$ve must have $\Delta T = +$ve.

c) For a cyclic process $q = 0$

d) Black phosphorus is most stable form of P but $H°_f = 0$ for white phosphorus.

228. Entropy decreases during:

a) Crystallization of sucrose from solution

b) Rusting of iron

c) Melting of ice

d) Vaporization of camphor

229. At 27°C latent heat of I fusion of a compound is 2930 J/mol. Entropy change during fusion is:

a) 9.77 J/mol K b) 10.77 J/mol K c) 9.07 J/mol K d) 0.977 J/mol K

230. The values of ΔH and ΔS for the reaction.

$C_{(\text{graphite})} + CO_2(g) \rightarrow 2CO(g)$

Are 170 kJ and 170 JK^{-1} respectively. this reaction will be spontaneous at:

a) 510 K b) 710 K c) 910 K d) 1110 K

231. The temperature of 5 mL of a strong acid increases by 5°C when 5 mL of strong base is added to it. If 10 mL of each is mixed and complete neutralisation takes place then rise in temperature will be

a) 20°C b) 10°C c) 5°C d) 2°C

232. When an ideal gas is compressed adiabatically and reversibly, the final temperature is:

a) Higher than the initial temperature

b) Lower than the initial temperature
c) The same as the initial temperature
d) Dependent on the rate of compression

233. In a closed insulated container, a liquid is stirred with a paddle to increase its temperature. In this process, which of the following is true
a) $\Delta E = W \neq 0, Q = 0$ b) $\Delta E \neq 0, Q = W = 0$
c) $\Delta E = W = Q = 0$ d) $\Delta E = 0, Q \neq 0, W = 0$

234. If the bond dissociation energies of XY, X_2 and Y_2 (all diatomic molecules) are in the ratio of 1: 1: 05 and ΔH for the formation of XY is -200 kJ mol^{-1}. The bond dissociation energy of X_2 will be
a) 100 kJ mol^{-1} b) 800 kJ mol^{-1} c) 300 kJ mol^{-1} d) 400 kJ mol^{-1}

235. The dissociation energy of CH_4 and C_2H_6 are respectively 360 and 620 kcal/mol. The bond energy of $C-C$ bond is:
a) 260 kcal/mol b) 180 kcal/mol c) 130 kcal/mol d) 80 kcal/mol

236. In a calorimeter, the temperature of the calorimeter increases by 6.12 K, the heat capacity of the system is 1.23 kJ/g deg. What is the molar heat of decomposition for NH_4NO_3?
a) -7.53 kJ/mol b) -398.1 kJ/mol c) -16.1 kJ/mol d) -602 kJ/mol

237. The bond energies of F_2, Cl_2, Br_2 and I_2 are 155.4, 243.6, 193.2 and 151.2 kJ mol^{-1} respectively. The strongest bond is :
a) $F-F$ b) $Cl-Cl$ c) $Br-Br$ d) $I-I$

238. The enthalpy changes of formation of the gaseous oxides of nitrogen (N_2O and NO) are positive because of:
a) The high bond energy of the nitrogen molecule
b) The high electron affinity of oxygen atoms
c) The high electron affinity of nitrogen atoms
d) The tendency of oxygen to form O^{2-}

239. If 900 J/g of heat is exchanged at boiling point of water then increase in entropy
a) 43.4 J/mol b) 87.2 J/mol c) 900 J/mol d) Zero

240. A reaction occurs spontaneously if:
a) $T\Delta S = \Delta H$ and both ΔH and ΔS are positive
b) $T\Delta S > \Delta H$ and both ΔH and ΔS are positive
c) $T\Delta S < \Delta H$ and both ΔH and ΔS are positive
d) $T\Delta S > \Delta H$ and ΔH is positive and ΔS are negative

241. $H_2(g) + Cl(g) = 2HCl(g)$; $\Delta H(298\ K) = 22.06$ kcal. For this reaction, ΔU is equal to:
a) $-22.06 + 2 \times 10^{-3} \times 298 \times 2$ kcal
b) $-22.06 + 2 \times 298$ kcal
c) $-22.06 - 2 \times 298 \times 4$ kcal
d) -22.06 kcal

242. The heat change taking place during the reaction $H_2O(l) \rightarrow H_2O(g)$ is: [Given, ΔH_f of $H_2O(g) = -57$ kcal, $\Delta H_f = H_2O(l) = -68.3$ kcal]
a) + 11.3 kcal b) − 11.3 kcal c) − 115.3 kcal d) + 115.3 kcal

243. ΔH for $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ is 176 kJ mol^{-1} at 1240 K. The ΔU for the change is equal to:
a) 160 kJ b) 165.6 kJ c) 186.3 kJ d) 180.0 kJ

244. When one mole of monoatomic ideal gas at TK 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) $T - \frac{2}{3 \times 0.0821}$

245. $\Delta H°$, (298 K)of methanol is given by the chemical equation
a) $CH_4(g) + \frac{1}{2}O_2(g) \rightarrow CH_3OH(g)$ b) $C(\text{graphite}) + \frac{1}{2}O_2(g) + 2H_2(g) \rightarrow CH_3OH(l)$

c) $C(\text{diamond}) + \frac{1}{2}O_2(g) \rightarrow CH_3OH(l)$ d) $CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$

246. For the reaction, $C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O$; $\Delta U = -1415$ kJ. Then ΔH at 27°C is :

a) −1410 kJ b) −1420 kJ c) + 1420 kJ d) +1410 KJ

247. The heat of combustion of ethanol determined by a bomb calorimeter is − 670.48 kcal mol^{-1} at 25°C. What is ΔU at 25°C for the following reaction?

$$C_2H_5OH(l) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

a) −335.24 kcal b) −669.28 kcal c) −670.48 kcal d) +670.48 kcal

248. Which has the least entropy?

a) Graphite b) Diamond c) $N_2(g)$ d) $N_2O(g)$

249. A carnot engine operates between temperature T and 400 K ($T > 400$ K). If efficiency of engine is 25%, the temperature T is:

a) 400 K b) 500 K c) 533.3 K d) 600 K

250. It is a general principle that if a system has the less energy then it is:

a) More stable b) Less stable c) Unstable d) More unstable

251. For the reaction, $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, which is true?

a) $\Delta H = \Delta U$ b) $\Delta H < \Delta U$ c) $\Delta H > \Delta U$ d) None of these

252. $H_2 + \frac{1}{2}O_2 \rightarrow H_2O$;

$\Delta H = -68.39$ kcal … … . . (i)

$K + aq. \rightarrow KOH(aq) + \frac{1}{2}H_2$;

$\Delta H = -48$ kcal … … . . (ii)

$KOH + aq. \rightarrow KOH(aq)$;

$\Delta H = -14$ kcal … … . . (iii)

The heat of formation (in kcal) of KOH is :

a) − 68.39 + 48 − 14

b) − 68.39 − 48 + 14

c) 68.39 − 48 + 14

d) 68.39 + 48 + 14

253. Which of the following expressions represents the first law of thermodynamics?

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

254. A thermodynamic state function is:

a) One which obeys all the laws of thermodynamics

b) A quantity which is used to measure thermal changes

c) A quantity whose value is independent of the path

d) A quantity which is used to express pressure-volume work

255. When two atoms of hydrogen combine to form a molecule of hydrogen gas, the energy of the molecule is:

a) Greater than that of separate atoms

b) Equal to that of separate atoms

c) Lower than that of separate atoms

d) Sometimes lower and sometimes higher

256. The enthalpies of formation of N_2O and NO are 28 and 90 kJ mol^{-1} respectively. The enthalpy of the reaction, $2N_2O(g) + O_2(g) \longrightarrow 4NO(g)$ is equal to :

a) 8 kJ b) 88 kJ c) − 16 kJ d) 304 kJ

257. Heat of combustion of CH_4, C_2H_4, C_2H_6 are − 890, −1411 and −1560 kJ/mol respectively. Which has the lowest calorific fuel value in kJ/g?

a) CH_4 b) C_2H_4 c) C_2H_6 d) All same

258. Given that $\Delta H_{r298\text{ K}} = -54.07$ kJ mol^{-1} and $\Delta S^{\circ}_{r298\text{ K}} = 10$ J mol^{-1} and $R = 8.314$ JK^{-1}mol^{-1}. The value of $\log_{10} K$ for a reaction, $A \rightleftharpoons B$ is:

a) 5 b) 10 c) 95 d) 100

259. Hess's law is based on

a) Law of conservation of mass b) Law of conservation of energy

c) First law of thermodynamics d) None of the above

260. What is the entropy change (in JK^{-1}mol^{-1}) when one mole of ice is converted into water at 0°C? (The enthalpy change for the conversion of ice to liquid water is 6.0 kJ mol^{-1} at 0°C)

a) 20.13 b) 2.013 c) 2.198 d) 21.98

261. Which is an extensive property of the system?

a) Temperature b) Volume c) Refractive index d) Viscosity

262. For the reaction; $C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$ which one is true?

a) $\Delta H = \Delta U - RT$

b) $\Delta H = \Delta U + RT$

c) $\Delta H = \Delta U + 2RT$

d) $\Delta H = \Delta U - 2RT$

263. The heat atomisation of $PH_3(g)$ is 228 kcal per mol and that of $P_2H_4(g)$is 335 kcal per mol. The energy of $P-P$ bond is

a) 102 kcal/mol b) 31 kcal/mol c) 26 kcal/mol d) 204 kcal/mol

264. If, $H_2(g) + Cl_2(g) \rightarrow 2HCl$; $\Delta H^{\circ} = -44$ kcal

$2Na(s) + 2HCl(g) \rightarrow 2NaCl(s) + H_2(g)$;

$\Delta H = -152$ kcal then,

$Na(s) + 0.5\ Cl_2(g) \rightarrow NaCl(s)$; $\Delta H^{\circ} = ?$

a) 108 kcal

b) 196 kcal

c) -98 kcal

d) 54 kcal

265. From the reaction, $P_{(\text{white})} \longrightarrow P_{(\text{red})}$; $\Delta H = -18.4$ kJ it following that :

a) Red P is readily formed from white P

b) White P is readily formed from red P

c) White P cannot be converted to red P

d) White P can be converted into red P and red P is more stable

266. If $H^+ + OH^- \longrightarrow H_2O + 13.7$ kcal then the heat of neutralization for complete neutralization of one mole of H_2SO_4 by a base will be :

a) $-$ 13.7 kcal b) $-$ 27.4 kcal c) $-$ 6.85 kcal d) $-$ 3.425 kcal

267. Assuming that water vapours are ideal gas. The change in internal energy (ΔU) when 1 mol of water is vaporized at 1 bar pressure and 100°C. (Given molar enthalpy of vaporization at 1 bar and 373 K is 41 kJ mol^{-1} and $R = 8.3$ J mol^{-1} K^{-1}) will be:

a) 41.00 kJ mol^{-1} b) 4.100 kJ mol^{-1} c) 3.7904 kJ mol^{-1} d) 37.904 kJ mol^{-1}

268. Change in entropy for a reaction is given by:

a) $2.303\ nR \log_{10} \frac{V_2}{V_1}$ b) $nR \log_e \frac{V_2}{V_1}$ c) $nR \log_e \frac{P_1}{P_2}$ d) All of these

269. At constant pressure and temperature, the direction of any chemical reaction is one where, the ... decrease.

a) Entropy b) Enthalpy c) Gibbs energy d) None of these

270. Which of the following conditions will always lead to a non spontaneous change?

a) Positive ΔH and positive ΔS b) Negative ΔH and negative ΔS

c) Positive ΔH and negative ΔS d) Negative ΔS and positive ΔS

271. Equal volume of two monoatomic gases, A and B, at same temperature and pressure are mixed. The ratio of specific heats (C_P/C_V) of the mixture will be:

a) 0.83 b) 1.50 c) 3.3 d) 1.67

272. Two atoms of hydrogen combine to form a molecule of hydrogen gas, the energy of the H_2 molecule is :
a) Greater than that of separate atoms
b) Equal to that of separate atoms
c) Lower than that of separate atoms
d) Sometimes lower and sometimes higher

273. The heats of neutralization of four acids A, B, C and D are - 13.7, -9.4, -11.2 and -12.4 kcal respectively when they are neutralized by a common base. The acidic character obeys the order :
a) $A > B > C > D$ b) $A > D > C > B$ c) $D > C > B > A$ d) $D > B > C > A$

274. $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$; $\Delta H = -68$ kcal. The heat change, for the decomposition of 3.6 g of water is :
a) 136 kcal b) 13.6 kcal c) 1.36 kcal d) 68 kcal

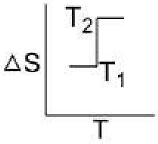
275. When 500 J heat is given to the gas X in an isobaric process its work done comes out as 142.8 J. The gas X is
a) O_2 b) NH_3 c) He d) SO_2

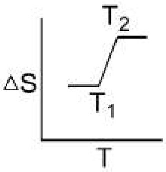
276. Diborane is a potential rocket fuel which undergoes combustion according to the equation
$B_2H_6(g) + 3O_2(g) \rightarrow B_2O_3(g) + 3H_2O(g)$
Calculate the enthalpy change for the combustion of diborane. Given
(i) $2B(s) + \frac{3}{2}O_2(g) \rightarrow B_2O_3(s)$; $\Delta H = -1273$ kJ per mol
(ii) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$; $\Delta H = -286$ kJ per mol
(iii) $H_2O(l) \rightarrow H_2O(g)$; $\Delta H = 44$ kJ per mol
(iv) $2B(s) + 3H_2(g) \rightarrow B_2H_6(g)$; $\Delta H = 36$ kJ per mol
a) +2035 kJ per mol b) −2035 kJ per mol c) +2167 kJ per mol d) −2167 kJ per mol

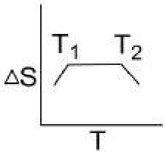
277. To calculate the amount of work done in joules during reversible isothermal expansion of an ideal gas, the volume must be expressed in
a) m^3 only b) dm^3 only c) cm^3 only d) Any of these

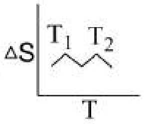
278. Calorific value of carbohydrates is approximately:
a) 4.0 kcal/g b) 16.0 kcal/g c) 20 kcal/g d) 9.0 kcal/g

279. For a given substance T_1 and T_2 are freezing point and melting point of a substance. Which of the graph represents correctly, the variation of ΔS with temperature?

a) 

b) 

c) 

d) 

280. Which is correct for an endothermic reaction?
a) ΔH is positive b) ΔH is negative c) ΔE is negative d) $\Delta H = 0$

281. When a solid melts, there is:
a) An increase in enthalpy
b) No change in enthalpy
c) A decrease in enthalpy
d) A decrease in internal energy

282. Maximum entropy will be in which of the following?
a) Ice b) Liquid water c) Snow d) Water vapour

283. When enthalpy and entropy change for a chemical reaction are -2.5×10^3 cal and 7.4 cal deg^{-1} respectively predict the reaction at 298 K is
a) Spontaneous b) Reversible c) Irreversible d) Non-spontaneous

284. A closed flask contains water in all its three states, solids, liquid and vapour at 0°C. In this situation the average KE of the water molecule will be:

a) Maximum in vapour state
b) Maximum in solid state
c) Greater in the liquid than in vapour state
d) Same in all the three states

285. If $C(s) + O_2(g) \rightarrow CO_2(g);\ \Delta H = r$
and $CO(g) + \frac{1}{2}O_2 \rightarrow CO_2(g);\ \Delta H = s$ then, the heat of formation of CO is
a) r+s b) r-s c) $s - r$ d) rs

286. The value of ΔH° for the reaction $Cu^{+}(g) + I^{-}(g) \rightarrow CuI(g)$ is -446 kJ mol^{-1}. If the ionisation energy of Cu(g) is 745 kJ mol^{-1}, and the electron affinity of (I)g is -295kJ mol^{1}, then the value of ΔH° for the formation of one mole of CuI(g) from Cu(g) and I(g) is :
a) -446 kJ mol^{-1} b) 450 kJ mol^{-1} c) 594 kJ mol^{-1} d) 4 kJ mol^{-1}

287. The entropy of the universe:
a) Increasing and tending towards maximum value
b) Decreasing and tending to be zero
c) Remains constant
d) Decreasing and increasing with a periodic rate

288. The internal energy of a substance
a) Increases with increase in temperature
b) Decreases with increases in temperature
c) Can be calculated by the relation $E = mc^2$
d) Remains unaffected with change in temperature

289. ΔH_f of graphite is 0.23 kJ/mol and ΔH_f of diamond is 1.896 kJ/mole. $\Delta H_{\text{transition}}$ from graphite to diamond is :
a) 1.66 kJ/mole b) 2.1 kJ/mole c) 2.33 kJ/mole d) 1.5 kJ/mole

290. When two moles of hydrogen expands isothermally against a constant pressure of 1 atm, at 25°C from 15 L to 50 L, the work done (in litre atm) will be
a) 17.5 b) 35 c) 51.5 d) 70

291. Which value of heat of formation indicates that the product is the least stable?
a) $-$ 94 kcal b) $-$ 231.6 cal c) $+$ 21.4 kcal d) $+$ 64.8 kcal

292. The heat of combustion for C, H_2 and CH_4 are $-$ 349.0, $-$241.8 and $-$ 906.7 kJ respectively. The heat of formation of CH_4 is :
a) 174.1 kJ b) 274.1 kJ c) 374.1 kJ d) 74.1 kJ

293. Given,
(i)$S + O_2 \rightarrow SO_2,\ \Delta H = -298.2$ kJ
(ii)$SO_2 + \frac{1}{2}O_2 \rightarrow SO_3,\ \Delta H = -98.7$ kJ
(iii)$SO_3 + H_2O \rightarrow H_2SO_4,\ \Delta H = -130.2$ kJ
(iv)$H_2 + \frac{1}{2}O_2 \rightarrow H_2O,\ \Delta H = -287.3$ kJ
Then the enthalpy of formation of H_2SO_4 at 298 K will be
a) -814.4 kJ b) $+320.5$ kJ c) -650.3 kJ d) -933.7 kJ

294. Based on the first law of thermodynamics, which one of the following is correct?
a) For an isochoric process$=\Delta E = -Q$ b) For an adiabatic process$=\Delta E = -W$
c) For an isothermal process$=Q = +W$ d) For a cyclic process$=Q = -W$

295. According to Hess's law, the heat of reaction depends upon
a) Initial condition of reactants b) Initial and final conditions of reactants
c) Intermediate path of the reaction d) End conditions of reactants

296. In which case, a spontaneous reaction is possible at any temperature?
a) $\Delta H -$ ve, $\Delta S +$ ve b) $\Delta H -$ ve, $\Delta S -$ ve c) $\Delta H +$ ve, $\Delta S +$ ve d) In none of the cases

297. Select the incorrect statement :

a) Combustion of F_2 is exothermic

b) Combustion of N_2 to N_2O is endothermic

c) A good fuel have higher calorific value

d) Nutrition calorie $= 10^3$ calories or 1 cal or 1 kcal

298. The enthalpy change for the following reaction. $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$ is -57 kJ. Predict the value of the enthalpy change in the following reaction : $Ba(OH)_2 + H_2SO_4(aq) \rightarrow BaSO_4\,(s) + 2H_2O(l)$

a) $-$ 57 kJ　　b) $-$ 76 kJ　　c) $-$ 114 kJ　　d) $-$ 228 kJ

299. At 27°C one mole of an ideal gas is compressed isothermally and reversible from a pressure of 2 atm to 10 atm. The value of ΔE and q are (R=2 cal)

a) 0, -965.84 cal　　b) -965.84 cal, -865.58 cal

c) +865.58 cal, -865.58 cal　　d) +965.84 cal, +865.58 cal

300. Change in entropy is negative for:

a) Bromine $(l) \longrightarrow$ Bromine (g)

b) $C(s) + H_2O(g) \longrightarrow CO(g) + H_2(g)$

c) $N_2(g, 10\text{ atm}) \longrightarrow N_2(g, 1\text{ atm})$

d) Fe (1 mol, 400 K) $\longrightarrow$ Fe(1 mol, 300 K)

301. Hess's law is related to:

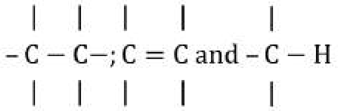
a) Change in heat during a reaction

b) Rates of reaction

c) Equilibrium constant

d) Influence of pressure on volume of a gas

302. Heat of dissociation of benzene of elements is 5335 kJ/mol. The bond enthalpies of



$-\underset{|}{\overset{|}{C}}-\underset{|}{\overset{|}{C}}-;\ \underset{|}{\overset{|}{C}}=\underset{|}{\overset{|}{C}}$ and $-\underset{|}{\overset{|}{C}}-H$

bonds are 347.3, 615 and 416.2 kJ respectively. Resonance energy of benzene is

a) 1.15 kJ　　b) 15.1 kJ　　c) 937.2 kJ　　d) 1511 kJ

303. 2 mole of ideal gas at 27°C temperature is expanded reversibly from 2 litre to 20 litre. Find entropy change ($R = 2$ cal/mol K).

a) 92.1　　b) 0　　c) 4　　d) 9.2

304. Work done by the system on surroundings is:

a) Positive　　b) Negative　　c) Zero　　d) None of these

305. What is ΔE for system that does 500 cal of work on surrounding and 300 cal of heat is absorbed by the system?

a) -200 cal　　b) -300 cal　　c) +200 cal　　d) +300 cal

306. Which fuel provides the highest calorific value?

a) Charcoal　　b) Kerosene　　c) Wood　　d) Dung

307. The value of ΔE for combustion of 16 g of CH_4 is -885389 J at 298 K. The ΔH combustion for CH_4 in J mol^{-1} at this temperature will be

(Given that, $R = 8.314\ \text{JK}^{-1}\ \text{mol}^{-1}$)

a) -55337　　b) -880430　　c) -885389　　d) -890348

308. Human body is an example of

a) Open system　　b) Closed system　　c) Isolated system　　d) None of these

309. A hypothetical reaction $A \rightarrow 2B$, proceeds through following sequence of steps

(i) $A \rightarrow C; \Delta H = q$　　(ii)$C \rightarrow D; \Delta H = v$

(iii)$\frac{1}{2}D \rightarrow B;\ \Delta H = x$

Then the heat of reaction i

a) $q - v + 2x$ b) $q + v - 2x$ c) $q + v + 2x$ d) $q + 2v - 2x$

310. The entropy of crystalline substances at absolute zero by the third law of thermodynamics should be taken as

a) 100 b) 50

c) Zero d) Different for different substance

311. Identify the state quantity among the following:

a) q b) $q - W$ c) $q + W$ d) q/W

312. For the following two reactions,

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

$\Delta H = -890.4$ kJ

(ii)$2HgO(s) \rightarrow 2Hg(l) + O_2(g) - 181.6$ kJ

Which one of the following statements is correct?

a) Both of them are exothermic

b) Both of them are endothermic

c) (i) is exothermic and (ii) is endothermic

d) (i) is endothermic and (ii) is exothermic

313. From the following data, the heat of formation of $Ca(OH)_2(s)$ at 18°C is .....kcal:

$CaO(s) + H_2O(l) = Ca(OH)_2(s);$

$\Delta H = -15.26$ kcal .... (i)

$H_2O(l) = H_2(g) + \frac{1}{2}O_2(g);$

$\Delta H = 68.37$ kcal ... ... (ii)

$Ca(s) + \frac{1}{2}O_2(g) = CaO(s);$

$\Delta H = -151.80$ kcal ... ... (iii)

a) − 98.69

b) − 235.43

c) 194.91

d) 98.69

314. If $\frac{1}{2}X_2O(s) \longrightarrow X(s) + \frac{1}{4}O_2(g);\ \Delta H = 90$ kJ, then heat change during reaction of metal 'X' with one more of O_2 to form oxide to maximum extent is :

a) 360 KJ b) −360 KJ c) −180 KJ d) +180 KJ

315. Decrease in Gibbs energy of a reacting system indicates to:

a) Exothermic reaction b) Equilibrium reaction c) Spontaneous reaction d) Slow reaction

316. $S + \frac{3}{2}O_2 \rightarrow SO_3 + 2x$ kcal;

$SO_2 + \frac{1}{2}O_2 \rightarrow SO_3 + y$ kcal;

The heat of formation of SO_2 is :

a) $y - 2x$ b) $(2x + y)$ c) $(x + y)$ d) $2x/y$

317. The standard molar heat of formation of ethane, CO_2 and water (l) are respectively −21.1 − 94.1 and −68.3 kcal. The standard molar heat of combustion of ethane will be

a) −372 kcal b) 162 kcal c) −240 kcal d) 183.5 kcal

318. Among them intensive property is

a) Mass b) Volume c) Surface tension d) Enthalpy

319. Equal volume of C_2H_2 and H_2 are combusted under identical condition. The ratio of their heat of combustion is :

$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g); \quad \Delta H = -241.8\ kJ$

$C_2H_2(g) + 2\frac{1}{2}O_2(g) \longrightarrow 2CO_2(g) + H_2O(g);$

$\Delta H = -1300\ kJ$

a) 5.37/1
b) 1/5.37
c) 1/1
d) None of these

320. 1 litre – atmosphere is equal to:

a) 101.3 J
b) 24.206 cal
c) 101.3×10^7 erg
d) All of these

321. For the precipitation of AgCl by Ag^+ ions and HCl:

a) $\Delta H = 0$
b) $\Delta G = 0$
c) $\Delta G = -ve$
d) $\Delta H = \Delta G$

322. When a certain amount of ethylene was burnt 6226 kJ heat was evolved. If heat of combustion of ethylene is 1411 kJ, the volume of O_2 (at NTP) that entered into the reaction is:

a) 296.5 mL
b) 296.5 litre
c) 6226 × 22.4 litre
d) 22.4 litre

323. The heat change for the reaction, $C(s) + 2S(s) \longrightarrow CS_2(l)$ is called:

a) Heat of solution of CS_2
b) Heat of fusion of CS_2
c) Heat of formation of CS_2
d) Heat of combustion of carbon

324. Given enthalpy of formation of $CO_2(g)$ and $CaO(s)$ are -94.0 kJ and -152 kJ respectively and the enthalpy of the reaction, $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ is 42 kJ. The enthalpy of formation of $CaCO_3(s)$ is :

a) – 42 kJ
b) – 202 kJ
c) + 202 kJ
d) – 288 kJ

325. The ratio of slopes of $\log P$ *vs* $\log V$ for reversible adiabatic process and reversible isothermal process of an ideal gas is equal to:

a) γ
b) $1 - \gamma$
c) $\gamma - 1$
d) $\frac{1}{\gamma}$

326. In which of the following processes of neutralization magnitude of $\Delta H_{\text{neutralisation}}$ is less than that of $\Delta H_{\text{ionisation}}$ of water?

a) HCl + NaOH
b) H_2SO_4 + NaOH
c) CH_3COOH + NaOH
d) $HClO_4$ + KOH

327. A boiled egg show a/an ...in entropy.

a) Increase
b) Decrease
c) No change
d) None of these

328. Which unit represents largest amount of energy?

a) Calorie
b) Joule
c) Erg
d) Electron volt

329. The heat of combustion of carbon to CO_2 is -393 kJ/mol. The heat released upon formation of 35.2 g of CO_2 from carbon and oxygen gas is

a) +325 kJ
b) −31.5 kJ
c) −315 kJ
d) +31.5 kJ

330. A thermally isolated gaseous system can exchange energy with the surroundings. The mode of transference of energy can be:

a) Heat
b) Work
c) Heat and radiation
d) None of these

331. Warming ammonium chloride with sodium hydroxide in a test tube is an example of:

a) Closed system
b) Isolated system
c) Open system
d) None of these

332. Identify the correct statements regarding entropy

a) At 0°C, the entropy of a perfectly crystalline substance is taken to be zero
b) At absolute zero of temperature, the entropy of all perfectly crystalline substance is positive
c) At absolute zero of temperature, the entropy of all crystalline substance is taken to be negative
d) At absolute zero of temperature, the entropy of a perfectly crystalline substance is taken to be zero

333. The matter has highest entropy in:
a) Solid state b) Liquid state c) Gaseous state d) Equal in all

334. $H_2(g) + \frac{1}{2}O_2(g) = H_2O(l)$; $\Delta H_{298K} = -68.32$ kcal. Heat of vaporization of water at 1 atm and 25°C is 10.52 kcal. The standard heat of formation (in kcal) of 1 mole of water vapour at 25°C is :
a) -78.84 b) 78.84 c) $+57.80$ d) -57.80

335. For vaporization of water at 1 atmospheric pressure, the values of ΔH and ΔS are 40.63 kJ mol^{-1} and 108.8 $JK^{-1}mol^{-1}$, respectively. The temperature when Gibbs energy change (ΔG) for this transformation will be zero, is:
a) 273.4 K b) 393.4 K c) 373.4 K d) 293.4 K

336. For an adiabatic process:
a) $Q = +W$ b) $Q = 0$ c) $\Delta U = q$ d) $P\Delta V = 0$

337. The $\Delta H^{\circ}_{\text{ionisation}}$ for HCN and CH_3COOH are 45.2 and 2.1 kJ mol^{-1}.Which of the following correct?
a) $pKa_{HCN} < pKa_{CH_3COOH}$
b) $pKa_{HCN} > pKa_{CH_3COOH}$
c) $pKa_{HCN} = pKa_{CH_3COOH}$
d) None of the above

338. Which of the following taking place in the blast furnace is endothermic?
a) $CaCO_3 \rightarrow CaO + CO_2$ b) $2C + O_2 \rightarrow 2CO$
c) $C + O_2 \rightarrow CO_2$ d) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

339. Hess's law is used to calculate
a) Enthalpy of reaction b) Entropy of reaction
c) Work done in reaction d) All of these

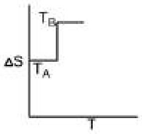
340. For spontaneity of a cell, which is correct?
a) $\Delta G = 0, \Delta E = 0$ b) $\Delta G = -\text{ve}, \Delta E = 0$ c) $\Delta G = +\text{ve}, \Delta E = +\text{ve}$ d) $\Delta G = -\text{ve}, \Delta E = +\text{ve}$

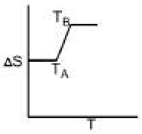
341. If heat of neutralisation is -13.7kcal and H_f^0 $H_2O = -68$ kcal, then enthalpy of OH^- would be :
a) 54.3 kcal b) -54.3 kcal c) 71.3 kcal d) -71.3 kcal

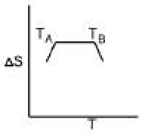
342. Work done in reversible isothermal process is given by:
a) $2.303\, RT \log\frac{V_2}{V_1}$ b) $\frac{nR}{(\gamma - 1)}(T_2 - T_1)$ c) $2.303\, RT \log\frac{V_1}{V_2}$ d) None of these

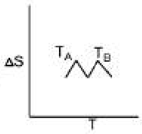
343. Internal energy of an ideal gas depends on:
a) Pressure b) Temperature c) Volume d) None of these

344. For a given substance, melting point T_B and freezing point is T_A, which of the following represents correct variation of ΔS vs T?

a)


b)


c)


d)


345. Bond energies of (H—H), (O═O) and (O—H) are 105, 120 and 220 kcal/mol respectively, then ΔH in the reaction , $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$:
a) – 115 b) – 130 c) – 118 d) – 550

346. The apparatus generally used for measuring heat changes is:
a) Voltameter b) Voltmeter c) Calorimeter d) Coulometer

347. The enthalpy change for the process, $C(s) \longrightarrow C(g)$ is known as enthalpy of :
a) Fusion b) Vaporisation c) Combustion d) Sublimation

348. Standard heat of formation of $CH_4(g), CO_2(g)$ and water 25°C are $-17.9, -94.1$ and -68.3 kcal mol^{-1} respectively. Calculate the heat change (in kcal) in the following reaction at 25°C:

$$CH_4(g) + 2O_2(g) = CO_2(g) + 2H_2O(l)$$

a) -144.5 b) -180.3 c) -248.6 d) -212.8

349. Which is the best definition of heat of neutralization?

a) The heat absorbed when one gram molecule of an acid is neutralized by one gram molecule of a base in dilute solution at a stated temperature

b) The heat set free or absorbed when one gram atom of an acid is neutralized by one gram atom of a base at a stated temperature

c) The heat set free or absorbed when a normal solution containing one gram-equivalent of an acid neutralized by a normal solution containing one gram-equivalent of a base at a stated temperature

d) The heat set free when one gram-equivalent of an acid is neutralized by one gram-equivalent of a base in dilute solution at a stated temperature

350. Thermochemistry is the study of relationship between heat energy and :

a) Chemical energy b) Activation energy c) Frictional energy d) None of these

351. Enthalpy change for the reaction, $4H(g) \rightarrow 2H_2(g)$ is -869.6 kJ

The dissociation energy of H—H bond is :

a) $+217.4$ kJ b) -434.8 kJ c) -869.6 kJ d) $+434.8$ kJ

352. Which of the following is true for an adiabatic process?

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

353. Which of the following is an intensive property?

a) Volume b) Enthalpy c) Surface tension d) Free energy

354. $C_6H_{12}(l) + 9O_2(g) = 6H_2O(l) + 6CO_2(g)$; $\Delta H_{298K} = -936.9$ kcal. Thus :

a) $-936.9 = \Delta U - (2 \times 10^{-3} \times 298 \times 3)$ kcal

b) $+936.9 = \Delta U + (2 \times 10^{-3} \times 298 \times 3)$ kcal

c) $-936.9 = \Delta U - (2 \times 10^{-3} \times 298 \times 2)$ kcal

d) $-936.9 = \Delta U + (2 \times 10^{-3} \times 298 \times 2)$ kcal

355. The work done by a weightless piston in causing an expansion ΔV (at constant temperature), when the opposing pressure, P is variable, is given by:

a) $W = -\int P\Delta V$ b) $W = 0$ c) $W = -P\Delta V$ d) None of these

356. If a gas has 2 atm and 5 atm pressure at 30°C and 27°C respectively. Then it will

a) Cool on expansion b) Warm on expansion

c) No change on expansion d) None of these

357. Two moles of helium gas expanded isothermally and irreversible at 27°C from volume 1 dm^3 to 1 m^3 at constant pressure of 100 k Pa. Calculate the work done.

a) 99900 kJ b) 99900 J c) 34464.65 kJ d) 34464.65 J

358. The efficiency of heat engine is maximum when:

a) Temperature of source > temperature of sink

b) Temperature of sink > temperature of source

c) Temperature difference of source and sink is minimum

d) Temperature difference of source and sink is maximum

359. Which one of the following has $\Delta S°$ greater than zero?

a) $CaO + CO_2(g) \rightleftharpoons CaCO_3(s)$ b) $NaCl(aq) \rightleftharpoons NaCl(s)$

c) $NaNO_3(s) \rightleftharpoons Na^+(aq) + NO_3^-(aq)$ d) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

360. A spontaneous change is one in which the system suffers

a) A lowering of entropy b) No energy change

c) An increase in internal energy d) A lowering of free energy

361. Which of the following are not state functions?

(I) $q + w$ (II) q

(III) w (IV) $H - TS$

a) (II), (III) and (IV) b) (I), (II) and (III) c) (II) and (III) d) (I) and (IV)

362. An isolated system is that system in which:

a) There is no exchange of energy with the surroundings
b) There is exchange of mass and energy with the surroundings
c) There is no exchange of mass and energy with the surroundings
d) There is exchange of mass with the surroundings

363. Thermodynamics is concerned with:
a) Total energy of a system
b) Energy changes in a system
c) Rate of chemical change
d) Mass changes in nuclear reactions

364. Which of the reactions defines $\Delta H^{\circ}f$?
a) $C_{(\text{diamond})} + O_2(g) \rightarrow CO_2(g)$
b) $\frac{1}{2}H_2(g) + \frac{1}{2}F_2(g) \rightarrow HF(g)$
c) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
d) $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$

365. A process is taking place at constant temperature and pressure. Then
a) $\Delta H = \Delta E$ b) $\Delta H = T\Delta S$ c) $\Delta H = 0$ d) $\Delta S = 0$

366. An isothermal process is associated with:
a) Constant entropy
b) Constant temperature
c) Constant enthalpy
d) Large change in heat content

367. $C(s) + O_2(g) \rightarrow CO_2(g);\ \Delta H = -94$ kcal
$2CO(g) + O_2 \rightarrow 2CO_2(g);\ \Delta H = -135.2$ kcal
The heat of formation of CO(g) is
a) -26.4 kcal   b) 41.2 kcal   c) 26.4 kcal   d) 229.2 kcal

368. $C_{\text{graphite}} + O_2(g) \rightarrow CO_2(g);$
$\Delta H = -94.05 \text{ kcal mol}^{-1}$
$C_{\text{diamond}} + O_2(g) \rightarrow CO_2(g);$
$\Delta H = -94.05 \text{ kcal mol}^{-1}$ therfore :
a) $C_{\text{diamond}} \rightarrow C_{\text{graphite}}$; $\Delta H^{\circ}_{298K} = +450 \text{ cal mol}^{-1}$
b) $C_{\text{graphite}} \rightarrow C_{\text{diamond}}$; $\Delta H^{\circ}_{298K} = -450 \text{ cal mol}^{-1}$
c) Diamond is harder than graphite
d) Graphite is the stabler allotrope

369. Enthalpy change for a reaction does not depend upon
a) The physical states of reactants and products
b) Use of different reactants for the same products
c) The nature of intermediate reaction steps
d) The differences in initial and final temperature of involved substances

370. Which of the following is correct option for free expansion of an ideal gas under adiabatic condition?
a) $q = 0, \Delta T < 0, w \neq 0$ b) $q = 0, \Delta T \neq 0, w = 0$ c) $q \neq 0, \Delta T = 0, w = 0$ d) $q = 0, \Delta T = 0, w = 0$

371. For a reaction at 25°C enthalpy change (ΔH) and entropy change (ΔS) are $-11.7 \times 10^3 \text{ J mol}^{-1}$ and $-105 \text{ J mol}^{-1}\text{K}^{-1}$ respectively. The reaction is:
a) Spontaneous   b) Non-spontaneous   c) Instantaneous   d) None of these

372. Which of the following is a path function?
a) Internal energy   b) Enthalpy   c) Work   d) Entropy

373. The work done by a system is 8 J, when 40 J heat is supplied to it. The change in internal energy of the system during the process is:

a) 32 J b) 40 J c) 36 J d) 44 J

374. Heat of reaction at constant volume is equal to :

a) $\Sigma U_P - \Sigma U_R$ b) $\Sigma U_R - \Sigma U_P$ c) $\Sigma H_P - \Sigma H_R$ d) $\Sigma H_R - \Sigma H_P$

375. Boiling point of a liquid is 50 K at 1 atm and $\Delta H_{\text{vap.}} = 460.6$ cal mol^{-1}. What will be its b. p. at 10 atm?

a) 150 K b) 75 K c) 100 K d) 200 K

376. The change in the enthalpy during the reaction, $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$, is called :

a) Heat of reaction b) Heat of neutralization c) Heat of formation d) Heat of liquefaction

377. The heat of combustion of rhombic and monoclinic sulphur are 70.96 and 71.03 kcal. The heat of transition of $S_{R\rightarrow M}$ is:

a) 70.96 kcal b) 71.03 kcal c) -70 cal d) $+70$ cal

378. Hess law is applicable for the determination of heat of

a) Reaction b) Formation c) Transition d) All of these

379. A heat engine absorbs heat Q_1 at temperature T_1 and heat Q_2 at temperature T_2, work done by the engine is $(Q_1 + Q_2)$ this data

a) Violates 1st law of thermodynamics

b) Violates 1st law of thermodynamics if a_1 is -ve

c) Violates 1st law of thermodynamics if a_2 is $-$ve

d) Does not violates 1st law of thermodynamics

380. In which of the following condition a chemical reaction can not occur?

a) ΔH and ΔS increase and $T\Delta S > \Delta H$ b) ΔH and ΔS decrease and $\Delta H > T\Delta S$

c) ΔH increase and ΔS decreases d) ΔH decreases and ΔS increases

381. $C_{\text{diamond}} + \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g});\ \Delta H = -395$ kJ …….(i)

$C_{\text{graphite}} + \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g});\ \Delta H = -393.5$ kJ …….(ii)

The ΔH, when diamond is formed from graphite, is :

a) -1.5 kJ b) $+1.5$ kJ c) $+3.0$ kJ d) -3.0 kJ

382. Entropy change of fusion at constant pressure is given by:

a) $\Delta S_{(f)} = \dfrac{\Delta H_f}{T}$ b) $\Delta S_{(f)} = \dfrac{\Delta G_f}{T}$ c) $\Delta S_{(f)} = \dfrac{\Delta H_f}{\Delta T}$ d) None of these

383. At 27°C, one mole of an ideal gas is compressed isothermally and reversibly from a pressure of 2 atm to 10 atm. The values of ΔE and q are $(R = 2)$

a) $0, -965.84$ cal b) -965.84 cal, -865.58 cal

c) $+865.58$ cal, -865.58 cal d) -865.58 cal, -865.58 cal

384. When hydrogen and oxygen burn to form water in an oxyhydrogen torch, the entropy change is:

a) Negative

b) Positive

c) Zero

d) May be positive or negative

385. The temperature coefficient of e. m. f. of a cell can be given by:

a) $\left(\dfrac{\partial E}{\partial T}\right)_P = \dfrac{\Delta S}{nF}$ b) $\left(\dfrac{\partial E}{\partial T}\right)_P = \left(\dfrac{E_2 - E_1}{T_2 - T_1}\right)_P$ c) $\left(\dfrac{\partial E}{\partial T}\right)_P = \left[\dfrac{\Delta H}{nF} + E\right]\dfrac{1}{T}$ d) All of these

386. The internal energy of one mole of a gas is:

a) $\dfrac{3}{2}RT$ b) $\dfrac{KT}{2}$ c) $\dfrac{RT}{2}$ d) $\dfrac{3KT}{2}$

387. The resultant heat change in a reaction is the same whether it takes place in one or several stages. This statement is called

a) Lavoisier and Laplace law b) Hess's law

c) Joule's law d) Le-Chatelier's principle

388. ΔH for transition of carbon in the diamond form to carbon in the graphite form, is – 453.5 cal. This suggests that :

a) Graphite is chemically different from diamond
b) Graphite is as stable as diamond
c) Graphite is more stable than diamond
d) Diamond is more stable than graphite

389. The enthalpy of fusion of ice per mole is

a) 18 kJ b) 8 kJ c) 80 kJ d) 6 kJ

390. For the reversible vaporisation of water at 100°C and 1 atmospheric pressure, ΔG is equal to:

a) ΔH b) ΔS c) Zero d) $\Delta H/T$

391. Molar heat of vaporisation of a liquid is 6 $\mathrm{kJmol^{-1}}$. If the entropy change is 16 $\mathrm{J\ mol^{-1}K^{-1}}$, the boiling point of the liquid is

a) 273 K b) 375°C c) 375 K d) 102°C

392. The enthalpy of fusion of water is 1.435 kcal/mol. The molar entropy change for the melting of ice at 0°C is:

a) 5.260 cal/(mol K) b) 0.526 cal/(mol K) c) 10.52cal/(mol K) d) 21.04 cal/(mol K)

393. Which of the following is an endothermic reaction?

a) $N_2(g) + 3H_2(g) - 92kJ \rightarrow 2NH_3(g)$
b) $N_2(g) + O_2(g) + 180.8\ kJ \rightarrow 2NO(g)$
c) $H_2(g) + Cl_2(g) \rightarrow 2HCl(g) + 184.6\ kJ$
d) $C(graphite) + 2H_2(g) \rightarrow CH_4(g) + 74.8\ kJ$

394. The ΔG in the process of melting of ice at -15°C is:

a) $\Delta G = -ve$ b) $\Delta G = +ve$ c) $\Delta G = 0$ d) All of these

395. A container has hydrogen and oxygen mixture in ratio of 1 : 1 by weight, then

a) Internal energy of the mixture decreases
b) Internal energy of the mixture increases
c) Entropy of the mixture increases
d) Entropy of the mixture decreases

396. Which one is not a spontaneous process?

a) Dissolution of $CuSO_4$ in water
b) Water flowing down the hills
c) Flow of current from low potential to high potential
d) None of the above

397. 1 mole of gas occupying 3 L volume is expanded against a constant external pressure of 1 atm to a volume of 15 litre. The work done by the system is:

a) 1.215×10^3 J b) 12.15×10^3 J c) 121.5×10^3 J d) None of these

398. The heat evolved during the combustion of 112 litre of water gas (mixture of equal volume of H_2 and CO) is :

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g);\ \ \Delta H = -241.8\ kJ$$

$$CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g);\ \ \Delta H = -283\ kJ$$

a) 241.8 kJ b) 283 kJ c) $-$ 1312 kJ d) 1586 kJ

399. Work done in reversible adiabatic process is given by:

a) $2.303\ RT \log\frac{V_2}{V_1}$ b) $\frac{nR}{(\gamma-1)}(T_2 - T_1)$ c) $2.303\ RT \log\frac{V_1}{V_2}$ d) None of these

400. The H—H bond energy is 430 kJ $\mathrm{mol^{-1}}$ and Cl—Cl bond energy is 240 kJ $\mathrm{mol^{-1}}$. ΔH for HCl is -90 kJ. The H—Cl bond energy is about:

a) 425 kJ $\mathrm{mol^{-1}}$ b) 213 kJ $\mathrm{mol^{-1}}$ c) 360 kJ $\mathrm{mol^{-1}}$ d) 180 kJ $\mathrm{mol^{-1}}$

401. Given,

$NH_3(g) + 3Cl_2(g) \rightleftharpoons NCl_3(g) + 3HCl(g);\ -\Delta H_1$
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g);\ \ -\Delta H_2$
$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g);\ \ \Delta H_3$
The heat of formation of $NCl_3(g)$ in the terms of $\Delta H_1, \Delta H_2$ and ΔH_3 is:

a) $\Delta H_f = -\Delta H_1 + \frac{\Delta H_2}{2} - \frac{3}{2}\Delta H_3$

b) $\Delta H_f = \Delta H_1 + \frac{\Delta H_2}{2} - \frac{3}{2}\Delta H_3$

c) $\Delta H_f = \Delta H_1 - \frac{\Delta H_2}{2} - \frac{3}{2}\Delta H_3$

d) None of the above

402. 50 mL of water takes 5 min to evaporate from a vessel on a heater connected to an electric source which delivers 400 W. The enthalpy of vaporisation of water is

a) 40.3 kJ per mol b) 43.2 kJ per mol c) 16.7 kJ per mol d) 180.4 kJ per mol

403. The work done by a system in an expansion against a constant external pressure is:

a) $\Delta P \cdot \Delta V$ b) $-P \cdot \Delta V$ c) Q d) $V \cdot \Delta P$

404. If, combustion of 4 g of CH_4 liberates 2.5 kcal of heat, the heat of combustion of CH_4 is :

a) -2 kcal mol^{-} b) -10 kcal mol^{-} c) 2.5 kcal mol^{-1} d) -5 kcal mol^{-1}

405. A gas expands isothermally against a constant external pressure of 1 atm from a volume of 10 dm^3 to a volume of 20 dm^3. It absorbs 800 J of thermal energy from its surroundings. The ΔU is

a) -312 J b) $+123$ J c) -213 J d) $+231$ J

406. Enthalpy of vaporisation for water is 186.5 kJ mol^{-1}. The entropy change during vaorisation is ...kJ K^{-1}mol^{-1}.

a) 0.5 b) 1.0 c) 1.5 d) 2.0

407. ΔC_p for : $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$ is :

a) $C_pNH_3 + C_pN_2 + C_pH_2$

b) $C_pNH_3 - C_pN_2 - C_pH_2$

c) $2C_pNH_3 - C_pN_2 - 3C_pH_2$

d) $2C_p\,NH_3 + C_pN_2 - 3C_pH_2$

408. Which statement(s) is/are correct?

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

b) $\left(\frac{\partial H}{\partial T}\right)_P > \left(\frac{\partial U}{\partial T}\right)_V$

c) $\left(\frac{\partial U}{\partial V}\right)_T$ for ideal gas is zero

d) All of the above

409. For a reversible process at $T = 300$ K, the volume of 2 mole of ideal gas is increased from 1 litre to 10 litre, the ΔH for isothermal change is:

a) 11.47 kJ b) 4.98 kJ c) 0 d) 2.49 kJ

410. If $C(s) + 2H_2(g) \rightarrow CH_4(g);\ \Delta H = -X_1$ kcal
$C(g) + 4H(g) \rightarrow CH_4(g);\ \Delta H = -X_2$ kcal
$CH_4(g) \rightarrow CH_3(g) + H(g);\ \Delta H = +Y$ kcal
The average bond energy of C – H bond in kcal mol^{-1} is :

a) $\frac{X_1}{4}$

b) Y

c) $\frac{X_2}{4}$

d) X_1

411. The enthalpy of formation of $H_2O(l)$ is given by :

a) $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l);\ \Delta H = -X_1$ kJ

b) $H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l);\quad \Delta H = -X_2$ kJ

c) $CO_2(g) + H_2(g) \longrightarrow CO(g) + H_2O(l);\ \Delta H = -X_3$ kJ

d) $C_2H_2(g) + \frac{5}{2}O_2(g) \longrightarrow 2CO_2(g) + H_2O(l);\ \Delta H = -X_4$ kJ

412. A gas can expand from 100 mL to 250 mL under a constant pressure of 2 atm. The work done by gas is

a) 30.38 J
b) 25 J
c) 5 kg J
d) 16 J

413. A system is changed from state A to state B by one path and from B to A by another path. If E_1 and E_2 are the corresponding changes in internal energy, then;

a) $U_1 + U_2 = +\text{ve}$
b) $U_1 + U_2 = -\text{ve}$
c) $U_1 + U_2 = 0$
d) None of these

414. Three moles of an ideal gas expanded spontaneously into vacuum. The work done will be:

a) Infinite
b) 3 joules
c) 9 joules
d) Zero

415. Which of the following is always negative for exothermic reaction?

a) ΔH
b) ΔS
c) ΔG
d) None of these

416. In a flask colourless N_2O_4 is in equilibrium with brown coloured NO_2. At equilibrium when the flask is heated at 100°C, the brown colour deepens and on cooling it becomes less coloured. The change in enthalpy, ΔH for this system is :

a) Negative
b) Positive
c) Zero
d) Undefined

417. If gas at constant temperature and pressure expands then its

a) Internal energy increases and then decreases
b) Internal energy increases
c) Internal energy remains the same
d) Internal energy decreases

418. If the enthalpy change for the reaction

$$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g),$$

$\Delta H = -25$ kcal. Bond energy of C—H is 20 kcal mol^{-1} greater than the bond energy of C—Cl and bond energies of H—H and H—Cl are same in magnitude, then for the reaction : $\frac{1}{2}H_2(g) + \frac{1}{2}Cl_2(g) \rightarrow HCl(g);\quad \Delta H = ?$

a) − 22.5 kcal/mol
b) − 20.5 kcal/mol
c) − 32.5 kcal/mol
d) − 12.5 kcal/mol

419. Internal energy is an example of

a) Path function
b) State function
c) Both (a) and (b)
d) None of these

420. In a closed container, a liquid is stirred with a paddle to increase the temperature. Which of the following is true?

a) $\Delta E = W \neq 0, q = 0$
b) $\Delta E = W = q \neq 0$
c) $\Delta E = 0, W = q \neq 0$
d) $W = 0, \Delta E = q \neq 0$

421. In a reaction, ΔH and ΔS both are positive. In which of the following cases, the reaction would not be spontaneous?

a) $\Delta H > T\Delta S$
b) $\Delta S = \frac{\Delta H}{T}$
c) $\Delta H = T\Delta S$
d) All of these

422. The bond dissociation energies of gaseous H_2, Cl_2 and HClare 104, 58 and 103 kcal respectively. The enthalpy of formation of HClgas would be

a) −44 kcal
b) 44 kcal
c) −22 kcal
d) 22 kcal

423. The internal energy change when a system goes from state A to B is 40 kJ/mol. If the system goes from A to B by a reversible path and returns to state A by an irreversible path, what would be the net change in internal energy?

a) 40 kJ
b) >40 kJ
c) <40 kJ
d) Zero

424. The enthalpy of vaporisation of a substance is 840 J/mol and its boiling point is −173°C. Its entropy of vaporisation is

a) 4.8 J/mol/K
b) 12 J/mol/K
c) 210 J/mol/K
d) 49 J/mol/K

425. The ΔH_f^o of O_3, CO_2, NH_3 and HI are 142.2 −393.3, −46.2 and +25.9 kJ per mol respectively. The order of their increasing stabilities will be

a) O_3, CO_2, NH_3, HI
b) CO_2, NH_3, HI, O_3
c) O_3, HI, NH_3, CO_2
d) NH_3, HI, CO_2, O_3

426. For the reaction, C (graphic)

$+\frac{1}{2}O_2(g) \rightarrow CO(g)$ at 298 K and 1 atm, $\Delta H = -26.4$ kcal. What is ΔE, if the molar volume of graphite is 0.0053 L?

$(R = 0.002 \text{ kcal mol}^{-1}\text{ K}^{-1})$

a) -26.7 kcal b) +26.7 kcal c) -52.4 kcal d) +52.4 kcal

427. For the reaction of one mole of zinc dust with one mole of sulphuric acid in a bomb calorimeter, ΔU and W correspond to

a) $\Delta U < 0, W = 0$ b) $\Delta U = 0, W < 0$ c) $\Delta U > 0, W = 0$ d) $\Delta U = 0, W > 0$

428. The quantity of heat measured for a reaction in a bomb calorimeter is equal to

a) ΔG b) ΔH c) $p\Delta V$ d) ΔE

429. For the reactions:

$C + O_2 \rightarrow CO_2;\ \Delta H = -393$ kJ

$2Zn + O_2 \rightarrow 2ZnO;\ \Delta H = -412$ kJ

which one is correct?

a) Carbon can reduce ZnO to Zn

b) Oxidation of carbon is not feasible

c) Oxidation of Zn is not feasible

d) Zn liberates more heat than carbon during oxidation

430. The following two reactions are known

$Fe_2O_3(s) + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)};$

$\Delta H = -26.8$ kJ

$FeO_{(s)} + CO_{(g)} \rightarrow Fe_{(s)} + CO_{2(g)};$

$\Delta H = -16.5$ kJ

The value of ΔH for the following reaction

$Fe_2O_{3(s)} + CO_{(g)} \rightarrow 2FeO_{(s)} + CO_{2(g)}$ is :

a) + 10.3 kJ

b) − 43.3 kJ

c) −10.3 kJ

d) + 6.2 kJ

431. The process in which pressure remains constant throughout a change is:

a) Adiabatic b) Isochoric c) Isobaric d) Isothermal

432. Entropy of vaporisation of water at 100°C, if molar heat of vaporisation is 9710 cal mol^{-1} will be:

a) 20 cal $\text{mol}^{-1}\text{K}^{-1}$ b) 26.0 cal $\text{mol}^{-1}\text{K}^{-1}$ c) 24 cal $\text{mol}^{-1}\text{K}^{-1}$ d) 28.0 cal $\text{mol}^{-1}\text{K}^{-1}$

433. Which is an extensive property?

a) Temperature b) Chemical potential c) Gibb's free energy d) Molar volume

434. $\Delta E°$of combustion of isobutylene is $-x$ kJ mol^{-1}. The value of $\Delta H°$ is

a) $= \Delta E°$ b) $> \Delta E°$ c) $= 0$ d) $< \Delta E°$

435. Which of the following reaction defines ΔH_f^o?

a) $C(\text{diamond}) + O_2(g) \rightarrow CO_2(g)$ b) $\frac{1}{2}H_2(g) + \frac{1}{2}F(g) \rightarrow HF(g)$

c) $N_2(l) + 3H_2(g) \rightarrow 2NH_3(g)$ d) $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$

436. The heat change for the reaction, $H_2O(l) \longrightarrow H_2O(g)$ is called as :

a) Heat of vaporisation b) Heat of solution c) Heat of fusion d) Heat of formation

437. The law of conservation of energy states that :

a) The internal energy of a system is constant

b) The heat content of a system is constant

c) Energy is neither created nor destroyed

d) There is an equivalence between energy and mass

438. Heat of neutralization of HF is :

a) 1. kJ b) > 57.32 kJ c) < 57.32 kJ d) None of these

439. From the following bond energies :

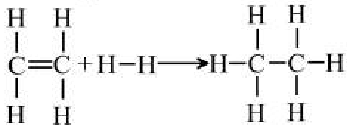
H—H bond energy : 431.37 kJ mol^{-1}

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

C—C bond energy : 336.49 kJ mol^{-1}

C—H bond energy : 410.50 kJ mol^{-1}

Enthalpy for the reaction,



$$\underset{\mathrm{H}}{\overset{\mathrm{H}}{\mathrm{C}}}=\underset{\mathrm{H}}{\overset{\mathrm{H}}{\mathrm{C}}}+\mathrm{H-H}\longrightarrow \mathrm{H-}\underset{\mathrm{H}}{\overset{\mathrm{H}}{\mathrm{C}}}\mathrm{-}\underset{\mathrm{H}}{\overset{\mathrm{H}}{\mathrm{C}}}\mathrm{-H}$$

will be :

a) 553.0 kJ mol^{-1} b) 1523.6 kJ mol^{-1} c) $-$ 243.6 kJ mol^{-1} d) $-$ 120.0 kJ mol^{-1}

440. The enthalpies of formation of organic substances can be conveniently determined from :

a) Heats of combustion data

b) Boiling point

c) Melting point

d) Heats of neutralisation

441. The free energy change for a reversible reaction at equilibrium is

a) Large, positive b) Small, negative c) Small, positive d) 0

442. In an irreversible process, the value of $\Delta S_{\text{system}} + \Delta S_{\text{surr}}$ is:

a) +ve b) $-$ve c) $= 0$ d) All of these

443. A container has hydrogen and oxygen mixture in ratio of 4 : 1 by weight, then:

a) Entropy of these gases increase

b) Internal energy increases

c) Internal energy of the gas decreases

d) Entropy of the gases decrease

444. Equal volume of 1M HCl and 1 M H_2SO_4 are neutralised by dilute NaOH solution and x and y kcal of heat are liberated respectively. Which of the following is true?

a) $x = y$ b) $x = 0.5\,y$ c) $x = 0.4\,y$ d) None of these

445. Identify the intensive quantity from the following

a) Enthalpy and temperature b) Volume and temperature

c) Enthalpy and volume d) Temperature and refractive index

446. A mixture of hydrogen and chlorine on exposure to ultra violet sunlight reacts with explosion. The step involved in the initiation of the reaction is :

a) $H_2 \longrightarrow H^{\bullet} + H^{\bullet}$ b) $Cl^{\bullet} + Cl^{\bullet} \longrightarrow Cl_2$ c) $H_2 + Cl_2 \longrightarrow 2HCl$ d) $Cl_2 \longrightarrow Cl^{\bullet} + Cl^{\bullet}$

447. The amount of heat measured for a reaction in a bomb calorimeter is

a) ΔG b) ΔH c) ΔE d) $p.\Delta V$

448. The mathematical form of the first law of thermodynamics when heat (q) is supplied and W is work done by the system (+ve) is:

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

449. Which one of the following bonds has the highest average bond energy (kcal/mol)?

a) $S = O$ b) $C \equiv C$ c) $C \equiv N$ d) $N \equiv N$

450. If gas, at constant temperature and pressure expands then its

a) Entropy increases and then decreases b) Internal energy increases

c) Internal energy remains the same d) Internal energy decreases

451. For a reaction, $\Delta H = 9.08 \text{ kJ mol}^{-1}$and $\Delta S = 35.7 \text{ JK}^{-1}\text{mol}^{-1}$. Which of the following statement is correct for the reaction?

a) Reversible and isothermal
b) Reversible and exothermic
c) Spontaneous and endothermic
d) Spontaneous and exothermic

452. The correct thermochemical equation is :

a) $C + O_2 \longrightarrow CO_2;\ \Delta H = -94$ kcal
b) $C + O_2 \longrightarrow CO_2;\ \Delta H = +94$ kcal
c) $C(s) + O_2(g) \longrightarrow CO_2(g);\ \Delta H = -94$ kcal
d) $C(s) + O_2(g) \longrightarrow CO_2(g);\ \Delta H = +94$ kcal

453. In which of the following reactions, standard reaction entropy change ($\Delta S°$) is positive and standard Gibbs's energy change ($\Delta G°$) decreases sharply with increasing temperature?

a) $Mg(s) + \frac{1}{2}O_2(g) \rightarrow MgO(s)$
b) $\frac{1}{2}C(\text{graphite}) + \frac{1}{2}O_2(g) \rightarrow \frac{1}{2}CO_2(g)$
c) $C(\text{graphite}) + \frac{1}{2}O_2(g) \rightarrow CO(g)$
d) $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$

454. The heat of formation of methane $C(s) + 2H_2(g) \rightarrow CH_4(g)$ at constant pressure is 18500 cal at 25°C. The heat of reaction at constant volume would be:

a) 19096 cal
b) 18798 cal
c) 18202 cal
d) 17904 cal

455. Minimum work is obtained when 1 kg of … gas expanded under 500 kPa to 200 kPa pressure at 0°C.

a) Chlorine
b) Oxygen
c) Nitrogen
d) Methane

456. The temperature of the system decreases in an

a) Adiabatic compression
b) Isothermal compression
c) Isothermal expansion
d) Adiabatic expansion

457. The bond of energies of H—H, Br—Br and H—Br are 433, 192 and 364 kJ mol^{-1}respecively. The $\Delta H°$for the reaction ; $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$ is :

a) − 261 kJ
b) + 103 kJ
c) + 261 kJ
d) − 103 kJ

458. The absolute enthalpy of neutralisation of the reaction $MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$ will be

a) Greater than $-57.33 \text{ kJ mol}^{-1}$
b) 57.33 kJmol^{-1}
c) Less than $-57.33 \text{ kJ mol}^{-1}$
d) $-57.33 \text{ kJ mol}^{-1}$

459. We believe in the laws of thermodynamics because they are

a) Theoretical
b) Derived based on mathematical analysis
c) Empirical and nobody disproved
d) Mere statements

460. The enthalpy of dissolution of $BaCl_2(s)$ and $BaCl_2 \cdot 2H_2O(s)$ are -20.6 and 8.8 kJ per mol respectively. The enthalpy of hydration for,

$BaCl_2(s) + 2H_2O \rightarrow BaCl_2 \cdot 2H_2O(s)$ is

a) 29.4 kJ
b) −29.4 kJ
c) −11.8 kJ
d) 38.2 kJ

461. The enthalpies of the elements in their standard states are assumed to be

a) Zero at 298 K
b) Unit at 298 K
c) Zero at all temperature
d) Zero at 273 K

462. The change in entropy, ΔS is positive for an endothermic reaction. If enthalpy charge ΔH occurs at the same temperature T, then the reaction is feasible:

a) At all temperatures
b) When $\Delta H > T\Delta S$
c) When $\Delta H < T\Delta S$
d) Not feasible at all

463. If the enthalpy change for the transition of liquid water to steam is 30 kJ mol^{-1} at 27°C, the entropy change for the process would be:

a) 100 $\mathrm{J\ mol^{-1}\ K^{-1}}$ b) 10 $\mathrm{J\ mol^{-1}\ K^{-1}}$ c) 1.0 $\mathrm{J\ mol^{-1}\ K^{-1}}$ d) 0.1 $\mathrm{J\ mol^{-1}\ K^{-1}}$

464. For the reaction, $C_3H_8(g) + 5O_2 \longrightarrow 3CO_2(g) + 4H_2O(l)$ at constant temperature, $\Delta H - \Delta U$ is :

a) $+ RT$ b) $- 3RT$ c) $+ 3RT$ d) $- RT$

465. The total amount of energy in the universe is fixed, but:

a) Matter is increasing
b) Gravitation is decreasing
c) Disorder is increasing
d) Lightening is increasing

466. Which compound will absorb the maximum amount of heat when dissolved in the same amount of water? (Integral heats of solution at 25°C in kcal/mol of each solute are given in brackets):

a) $HCl(\Delta H = -17.74)$
b) $HNO_3(\Delta H = -7.85)$
c) $NH_4NO_3(\Delta H = +16.08)$
d) $NaCl(\Delta H = +1.02)$

THERMODYNAMICS

: ANSWER KEY :

1)	c	2)	b	3)	b	4)	a	161)	d	162)	a	163)	c	164)	b
5)	c	6)	d	7)	c	8)	a	165)	b	166)	c	167)	c	168)	d
9)	a	10)	b	11)	a	12)	d	169)	c	170)	a	171)	a	172)	d
13)	a	14)	c	15)	b	16)	a	173)	b	174)	d	175)	d	176)	c
17)	b	18)	d	19)	a	20)	d	177)	d	178)	d	179)	b	180)	a
21)	b	22)	a	23)	a	24)	d	181)	b	182)	a	183)	a	184)	a
25)	c	26)	c	27)	b	28)	a	185)	b	186)	c	187)	d	188)	d
29)	d	30)	a	31)	a	32)	b	189)	c	190)	b	191)	a	192)	c
33)	a	34)	c	35)	a	36)	a	193)	c	194)	a	195)	c	196)	a
37)	c	38)	c	39)	c	40)	a	197)	a	198)	c	199)	b	200)	d
41)	c	42)	c	43)	a	44)	d	201)	c	202)	b	203)	b	204)	a
45)	a	46)	b	47)	c	48)	d	205)	d	206)	b	207)	b	208)	d
49)	b	50)	c	51)	a	52)	d	209)	a	210)	d	211)	d	212)	b
53)	a	54)	b	55)	d	56)	c	213)	b	214)	d	215)	b	216)	b
57)	c	58)	b	59)	a	60)	a	217)	d	218)	a	219)	d	220)	c
61)	b	62)	a	63)	c	64)	a	221)	c	222)	a	223)	c	224)	a
65)	b	66)	c	67)	d	68)	a	225)	b	226)	d	227)	b	228)	a
69)	d	70)	b	71)	c	72)	a	229)	a	230)	d	231)	c	232)	b
73)	d	74)	a	75)	c	76)	a	233)	a	234)	b	235)	d	236)	d
77)	d	78)	a	79)	d	80)	b	237)	b	238)	a	239)	a	240)	b
81)	a	82)	a	83)	a	84)	d	241)	d	242)	a	243)	b	244)	a
85)	a	86)	a	87)	b	88)	c	245)	b	246)	b	247)	c	248)	b
89)	b	90)	d	91)	d	92)	a	249)	c	250)	a	251)	b	252)	b
93)	c	94)	c	95)	d	96)	a	253)	c	254)	c	255)	c	256)	d
97)	a	98)	a	99)	b	100)	c	257)	b	258)	b	259)	b	260)	d
101)	a	102)	d	103)	b	104)	d	261)	b	262)	a	263)	b	264)	c
105)	b	106)	c	107)	d	108)	a	265)	d	266)	b	267)	d	268)	d
109)	c	110)	b	111)	d	112)	c	269)	c	270)	c	271)	d	272)	c
113)	b	114)	a	115)	a	116)	c	273)	b	274)	b	275)	a	276)	b
117)	b	118)	a	119)	c	120)	a	277)	d	278)	a	279)	a	280)	a
121)	a	122)	a	123)	a	124)	c	281)	a	282)	d	283)	a	284)	d
125)	a	126)	c	127)	c	128)	a	285)	b	286)	d	287)	a	288)	a
129)	b	130)	a	131)	a	132)	a	289)	a	290)	b	291)	d	292)	d
133)	d	134)	b	135)	c	136)	b	293)	a	294)	d	295)	b	296)	a
137)	b	138)	b	139)	b	140)	c	297)	a	298)	c	299)	a	300)	d
141)	c	142)	a	143)	b	144)	c	301)	a	302)	c	303)	d	304)	b
145)	b	146)	b	147)	a	148)	b	305)	a	306)	b	307)	b	308)	a
149)	c	150)	a	151)	d	152)	a	309)	c	310)	c	311)	c	312)	c
153)	a	154)	a	155)	b	156)	d	313)	b	314)	b	315)	c	316)	a
157)	d	158)	d	159)	d	160)	c	317)	a	318)	c	319)	a	320)	d

321)	c	322)	b	323)	c	324)	d	397)	a	398)	c	399)	b	400)	a
325)	a	326)	c	327)	a	328)	a	401)	a	402)	b	403)	b	404)	b
329)	c	330)	b	331)	c	332)	d	405)	c	406)	a	407)	c	408)	d
333)	c	334)	d	335)	c	336)	b	409)	b	410)	c	411)	b	412)	a
337)	b	338)	a	339)	a	340)	d	413)	c	414)	d	415)	a	416)	b
341)	b	342)	a	343)	b	344)	a	417)	c	418)	a	419)	b	420)	a
345)	d	346)	c	347)	d	348)	d	421)	d	422)	c	423)	d	424)	a
349)	d	350)	a	351)	d	352)	c	425)	c	426)	a	427)	a	428)	d
353)	c	354)	a	355)	a	356)	b	429)	a	430)	d	431)	c	432)	b
357)	b	358)	d	359)	c	360)	d	433)	c	434)	d	435)	c	436)	a
361)	c	362)	c	363)	b	364)	b	437)	c	438)	b	439)	d	440)	a
365)	a	366)	b	367)	a	368)	d	441)	d	442)	a	443)	a	444)	b
369)	c	370)	a	371)	b	372)	c	445)	d	446)	d	447)	c	448)	b
373)	a	374)	a	375)	c	376)	b	449)	d	450)	c	451)	c	452)	c
377)	d	378)	d	379)	d	380)	c	453)	c	454)	a	455)	a	456)	d
381)	b	382)	a	383)	a	384)	a	457)	d	458)	c	459)	d	460)	b
385)	d	386)	a	387)	b	388)	c	461)	a	462)	c	463)	a	464)	b
389)	d	390)	c	391)	c	392)	a	465)	c	466)	c				
393)	b	394)	b	395)	c	396)	c								

THERMODYNAMICS

: HINTS AND SOLUTIONS :

1 **(c)**

$H_2O(l) \rightarrow H_2O(v) \quad \therefore \Delta n = 1$

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

$40.66 \times 10^3 = \Delta U^\circ + 1 \times 8.314 \times 373$

$\therefore \Delta U^\circ = 37559 \text{ J} = 37.56 \text{ kJ}$

2 **(b)**

$\left(\frac{\partial Q}{\partial T}\right)_P = C_p$ and $\left(\frac{\partial Q}{\partial T}\right)_V = C_v$ and $C_p - C_v = R$.

4 **(a)**

$C(s) + \frac{1}{2}O_2(g) \rightarrow CO_2(g); \Delta H_1 = -26.4$

$C(s) + \frac{1}{2}O_2(g) \rightarrow CO_2(g); \Delta H_2 = -96.0$

We have to find for ΔH, for reaction

$CO + \frac{1}{2}O_2 \rightarrow CO_2; \Delta H_1 = ?$

Eqs.(ii) − (i), we get

$\Delta H = -94 - (-26.4) = -67.6 \text{ kcal}$

5 **(c)**

If $\Delta G = -\text{ve}$, process is spontaneous

6 **(d)**

$\Delta U - W = q$

If cyclic process, then $q = -W$.

7 **(c)**

e.g., U, H, S, etc.

8 **(a)**

$\Delta G = 0$ for an equilibrium state.

9 **(a)**

Find ΔH for, $H_2 + S + 2O_2 \rightarrow H_2SO_4$

11 **(a)**

$q = +10 \text{ kJ}, W = -4 \text{ kJ}$

$\because \Delta E = q + W$

$= 10 - 4 = 6 \text{ kJ}$

So, energy increases by 6 kJ

12 **(d)**

Heat of neutralisation is $-13.7 \text{ kcal eq}^{-1}$.

13 **(a)**

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

$\Delta G = 0$, at equilibrium

$\therefore \Delta H = T\Delta S$

or $30.5 = T \times 0.066$

$T = 462.12 \text{ K}$

14 **(c)**

An adiabatic process is one in which exchange of heat is not taking place in between system and surroundings. This can be made by putting insulation at the boundries of system

15 **(b)**

$$\begin{array}{l} CH_3COOH + OH^- \longrightarrow CH_3COO^- + H_2O + q_1 \\ \qquad \underset{-}{H^+} + \underset{-}{OH^-} \longrightarrow \underset{-}{H_2O} + \underset{-}{q_2} \\ \hline CH_3COOH \rightarrow CH_3COO^- + H^+ + (q_1 - q_2) \end{array}$$

17 **(b)**

$$\Delta G^\circ = -2.303\, RT \log K \quad \left(K \text{ for } H_2O = \frac{10^{-14}}{55.6}\right)$$

$$= -2.303 \times 8.314 \times 298 \times \log \frac{10^{-14}}{55.6}$$

$= -89.84 \text{ kJ}$

18 **(d)**

The spontaneity of reaction cannot be decided by simply looking the chemical change. We need ΔG value for it.

19 **(a)**

ΔH for a reaction is equal but opposite to backward reaction.

20 **(d)**

Follow definition of ionisation energy.

21 **(b)**

Internal energy of an ideal gas is function of temperature and thus $\left(\frac{\partial E}{\partial P}\right)_T = 0$

22 **(a)**

For exothermic reactions $H_p < H_R$

23 **(a)**

Heat of combustion is defined as the amount of the energy released during complete combustion of 1 mole of a substance in oxygen or air.

24 **(d)**

50 Meq. of KOH and 50 Meq. Of H_2SO_4 will produce maximum heat.

25 **(c)**

The first law of thermodynamics can be expressed as :

$$\Delta E = q + W$$
$$q = \Delta E - W$$

26 **(c)**

$H = U + PV$

$\therefore H_2 - H_1 = U_2 - U_1 + (P_2V_2 - P_1V_1)$

$\therefore \Delta H = 30 + (4 \times 5 - 2 \times 3)$

$= 44$ L atm

27 **(b)**

The reactions in which products has lesser energy than reactants, then energy is released in the reaction and such reactions are known as exothermic reactions *e.g.,*

$$N_2 + 3H_2 \rightarrow 2NH_3 + 92 \text{ kJ}$$

In this equation energy is released, so, it is an example of exothermic reaction.

28 **(a)**

$\Delta H_{\text{mixing}} = 0$ for ideal solutions.

29 **(d)**

In (a), (b), (c) randomness increase.

30 **(a)**

$\Delta H = \Delta E + \Delta nRT$

Since, $\Delta n = 2 - 4 = -2$

Therefore, $\Delta H = \Delta E - 2RT$

31 **(a)**

$$\frac{T}{T_{\text{final}}} = \left(\frac{V_2}{V_1}\right)^{\gamma - 1}$$

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

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

32 **(b)**

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Molecular weight of $CH_4 = 12 + 4 = 16$

$\therefore$ On the combustion of 2.0 g of methane = 25.0 kcal

$\therefore$ On the combustion of 16.0 g methane $= \frac{25 \times 16}{2} =$ 200 kcal

33 **(a)**

$$\Delta S = \frac{q}{T} = \frac{2.303\, nRT}{T} \log \frac{V_2}{V_1}$$

$= 2.303 \times 1 \times 8.314 \log 10$

$= 19.15$ JK^{-1}mol^{-1}

34 **(c)**

Heat of formation of $H_2O(l)$

i.e., $H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l)$

Is also heat of combustion of $H_2(g)$.

35 **(a)**

$$\Delta S_{\text{reaction}} = \sum S_{\text{product}} - \sum S_{\text{reactant}}$$

$= 2 \times S_{H_2O} - [2 \times S_{H_2} + S_{O_2}]$

$= 2 \times 68 - [2 \times 126.6 + 201.20]$

$= -318.4$ JK^{-1}mol^{-1}

36 **(a)**

Solve using Hess's law

37 **(c)**

Decomposition of $MgCO_3$ occurs only on heating.

38 **(c)**

Maximum work is obtained under reversible conditions (either isothermal or adiabatic).

39 **(c)**

$$C_6H_6(g) + \frac{15}{2}O_2(g) \rightarrow 6CO_2(g) + 3H_2O(g)$$

$$\Delta n = 6 + 3 - 1 - \frac{15}{2} = +0.5$$

40 **(a)**

$P \times 1 = RT$

Also internal energy, $U = \frac{3}{2}RT$

$\therefore U = \frac{3}{2}P$

or $P = \frac{2}{3}U$

41 **(c)**

Electrical energy is used to bring in electrolysis.

42 **(c)**

The second law of thermodynamics has been defined as – the entropy of universe is always increasing in the course of every spontaneous process.

43 **(a)**

Rest all are endothermic process. Hydration is always exothermic.

44 **(d)**

For monoatomic gas, $\gamma_2 = \frac{C_p}{C_v} = 1.67$

For diatomic gas, $\gamma_2 = \frac{C_p}{C_v} = 1.40$

$\therefore \gamma_1 : \gamma_2 = \frac{1.67}{1.40} = 1.19 : 1$

45 **(a)**

$C + O_2 \rightarrow CO_2;$ $\quad \Delta H = -393.5$ kJ

$CO + \frac{1}{2}O_2 \rightarrow CO_2;$ $\quad \Delta H = -283.5$ kJ

On subtracting, $C + \frac{1}{2}O_2 \rightarrow CO;\ \Delta H - 110.0$ kJ

46 **(b)**

It is based on the fact that neither heat can be produced nor destroyed, i.e., Ist law of thermodynamics.

47 **(c)**

$\Delta H = H_P - H_R$; Measurements at constant pressure give ΔH values.

48 **(d)**

These all are limitations of third law.

49 **(b)**

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

$$\Delta n = -1/2$$

$$\therefore \quad -43 = \Delta U + (-1/2) \times 8.314 \times 298 \times 10^{-3}$$

$$\therefore \quad \Delta U = -241.76 \text{ kJ}$$

50 **(c)**

Molar heat capacity $= \left(\frac{\delta Q}{\delta T}\right)_P$

If $\delta T = 0$ at constant temperature.

Molar heat capacity $= \infty$

51 **(a)**

It is heat of neutralization of water.

52 **(d)**

Vapour phase has more heat enthalpy and the evaporation being spontaneous and thus, $\Delta G =$ $-$ve. Also randomness or disorder is more in vapour phase.

53 **(a)**

$W = -P_{\text{ext.}}(V_2 - V_1)$

$\because P_{\text{ext.}} = 0$

$\therefore W = 0$

54 **(b)**

$\Delta G = -RT \ln K_p = -2.303\, RT \log_{10} K_p$

55 **(d)**

Negative specific heat refers that in order to rise the temperature, certain quantity of heat is to be withdrawn from the body.

56 **(c)**

$$\begin{array}{ll} \text{Na}^+(s) + \text{OH}^-(s) + aq. \rightarrow \text{Na}^+(aq.) + \text{OH}^-(aq.); & \\ & \Delta H = -470.7 \text{ kJ} \\ \text{OH}^-(s) + aq. \rightarrow \text{OH}^-(aq.); & \Delta H = -228.8 \text{ kJ} \\ \quad - \quad\quad - \quad\quad\quad - \quad\quad\quad\quad + & \\ \hline \therefore \ \text{Na}^+(s) + aq. \rightarrow \text{Na}^+(aq.); & \Delta H = -241.9 \text{ kJ} \end{array}$$

57 **(c)**

$W = -P\Delta V = -1 \times 10^5\,(1 \times 10^{-2} - 1 \times 10^{-3})$

$= -900$ J

58 **(b)**

O_3 has more energy than O_2

59 **(a)**

$q = \Delta U - W = \Delta U$ ($\because W = 0$, at constant volume)

60 **(a)**

$q = \Delta U - W$: all have same units.

61 **(b)**

Energy change at constant pressure is called enthalpy change or change in heat content, i.e., ΔH

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

if $\Delta n = 0, \Delta H = \Delta U$, i.e., heat change at constant volume.

62 **(a)**

Vibration energy has both components kinetic and potential energy.

64 **(a)**

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

$\Delta G = 145.6 - 273 \times 0.116$

$= 113.93$ kJ/mol

65 **(b)**

In BF_3 $p\pi - p\pi$ interaction leads to back bonding due to vacant p–orbitals of boron and completely filled p-orbitals of F.

66 **(c)**

$\text{CaO} + \text{H}_2\text{O} \longrightarrow \text{Ca(OH)}_2$; $\Delta H = -$ve; the solution of lime heats up.

67 **(d)**

$T_2 = 150 + 273 = 423$ K,

$T_1 = 25 + 273 = 298$ K

$Q = 500$ K

$$\frac{W}{Q} = \frac{T_2 - T_1}{T_2}$$

$$\therefore W = 500\left(\frac{423 - 298}{423}\right) = 147.7 \text{ J}$$

68 **(a)**

The order of randomness, Gas > Liquid > Solid.

69 **(d)**

$$\Delta H = H_{C_3H_8} - H_{CH_4} - H_{C_2H_4}$$

$$= -24.8 - (-17.9) - 12.5$$

$$= \ 19.4 \text{ kcal/mol}$$

70 **(b)**

Due to randomness of particles is reduced since, entropy decreases

71 **(c)**

1 cal $= 4.185$ J $= 4.18 \times 10^7$ erg $= \frac{4.18}{1.602} \times 10^{19}$eV.

72 **(a)**

The change ΔS will be more and the q or W will be more because $\Delta S = \frac{q}{T}$ and work depends on q.

73 **(d)**

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

For (a), (b), (c) $\Delta n = 0$ and for (d)$\Delta n = -2$

74 **(a)**

Helmholtz free energy change ΔA represents network done.

75 **(c)**

$$\Delta H^\circ = \frac{1}{2} \times \Delta_{\text{diss.}} H^\circ_{Cl_2} + \Delta_{\text{eg}} H^\circ_{Cl} + \Delta_{\text{hyd}} H^\circ_{Cl}$$

$$= \frac{1}{2} \times 240 - 349 - 381$$
$$= -610 \text{ kJ mol}^{-1}$$

76 **(a)**

The energy can neither be created nor destroyed.

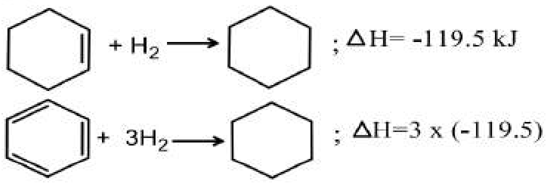
77 **(d)**

Heat changes during any chemical reaction are referred as heat of reaction are referred as heat of reaction for that change.

78 **(a)**

Amount of sugar needed $= \frac{2870 \times 342}{1349} = 727.6$ g

79 **(d)**



$$= -358.5 \text{ kJ mol}^{-1}$$

The obs. $\Delta H = \Delta H +$ Energy needed to disturb resonance

$$= -358.5 + 150.4$$
$$= -208.1 \text{ kJ mol}^{-1}$$

80 **(b)**

$\because \Delta G^\circ = -RT \ln K$

Also $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$\therefore -RT \ln K = \Delta H^\circ - T\Delta S^\circ$

or $\ln K = \frac{T\Delta S^\circ - \Delta H^\circ}{RT}$

81 **(a)**

Work done due to change in volume against constant pressure,

$$W = -p(V_2 - V_1)$$
$$= -1 \times 10^5 \text{ Nm}^{-2}(1 \times 10^{-2} - 1 \times 10^{-3})\text{m}^3$$
$$= -900 \text{ Nm} = -900 \text{ J} \quad (1 \text{ Nm}=1 \text{ J})$$

82 **(a)**

$C + O_2 \rightarrow CO_2;\ \Delta H = x$ ...(i)

$CO + \frac{1}{2}O_2 \rightarrow CO_2;\ \Delta H = y$ ...(ii)

Eqs. (i) –(ii)

$C + \frac{1}{2}O_2 \rightarrow CO, \Delta H = x - y$

83 **(a)**

$\Delta H = H_{\text{product}} - H_{\text{reactant}};\quad H_R > H_P$. Thus, –ve.

84 **(d)**

These all are facts.

85 **(a)**

$Fe + 2HCl \rightarrow FeCl_2 + H_2$

mole of Fe $= \frac{112}{56} = 2, \therefore$ mole of H_2 formed $= 2$.

now, work done $= P(V_2 - V_1); V_2 = V_{H_2}$ and

$V_1 = 0$ (for solid and liquid)

$$= P \cdot V_{H_2} = P \cdot \frac{nRT}{P} = nRT$$
$$= 2 \times 2 \times 300 = 1200 \text{ cal} = 1.2 \text{ kcal}$$

86 **(a)**

No doubt, heat evolved in first process is twice to second but volume absorbing this heat is also twice in comparison to first.

87 **(b)**

At constant volume $P\Delta V = 0, \therefore q = \Delta U$.

88 **(c)**

$C + O_2 \rightarrow CO_2$; $\Delta H^\circ f$ = ? if reaction is made at 25°C and 1 atm.

$\Delta H^\circ f = H^\circ_{CO_2} - H^\circ_C - H^\circ_{O_2} = H^\circ_{CO_2} - 0 - 0$

$\Delta H^\circ f = H^\circ_{CO_2}$ (H°_C *and* $H^\circ_{O_2}$ are assumed arbitrarily zero)

90 **(d)**

A decrease in Gibbs energy results for useful work done by the system, *i.e.*, work of expansion $(-W_{\text{expansion}})$ or $-\Delta G = W_{\text{exp}}$.

91 **(d)**

Strong acid (HNO_3) and strong base (LiOH).

92 **(a)**

This is derived formula.

93 **(c)**

$C + 2H_2 \rightarrow CH_4;\ \Delta H = ?$

Find ΔH by eqs. (i) + 2 × (ii) – (iii)

94 **(c)**

Due to positive ΔH, HI is endothermic compound and unstable.

96 **(a)**

$\Delta S = \frac{\Delta H_v}{T}$

$\Delta H = 40.8$ kJ

$T = 373$ K

$\therefore \Delta S = \frac{40.8 \times 10^3}{373} = 109.38 \text{ JK}^{-1}\text{mol}^{-1}$

97 **(a)**

$$\Delta H = -2[3 \times e_{N-H}] + e_{N\equiv N} + 3 \times e_{H-H}$$
$$= -2 \times 3 \times 391 + 945 + 436 \times 3$$
$$= -93 \text{ kJ}$$

98 **(a)**

The heat is provided by solvent and thus, feels cooler.

99 **(b)**

$\Delta G = \Delta G^\circ + 2.303RT \log Q$

if $Q = 1,\ \Delta G = \Delta G^\circ$

100 **(c)**

Bond energy is the amount of heat required (or evolved) to break (or form) 1 mole bond. Also,

$$H_2 \rightarrow 2H\,;\ \Delta H = 104.3 \text{ kcal}$$

$\therefore$ Heat of formatin for H atom $\frac{1}{2}$ $H_2 \rightarrow H$

i.e., equal to $\frac{1}{2} \times$ 104.3 kcal

101 **(a)**

Sublimation requires energy, *i.e.*, $\Delta H = +$ve, also, $\Delta S = +$ve.

103 **(b)**

Work done$(W) = -p_{\text{ext}}(V_2 - V_1)$

$= -3 \times (6 - 4) = -6$ L atm

$= -6 \times 101.32$ J($\therefore$ 1 L atm=101.32 J)

$= -607.92 \approx -608$ J

104 **(d)**

These are the expressions to explain work.

105 **(b)**

$$\Delta S_{\text{vap}} = \frac{\Delta H_{\text{vap}}}{T} = \frac{37.3 \text{ kJ mol}^{-1}}{373 \text{ K}}$$

$$= 0.1 \text{ kJ mol}^{-1}\text{K}^{-1}$$

$$= 100 \text{ J mol}^{-1}\text{K}^{-1}$$

106 **(c)**

Heat cannot be itself pass from colder to hotter body.

107 **(d)**

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

$$\Delta G = 150 - \frac{100 \times 300}{10^{-3}} = 120 \text{ kJ}$$

108 **(a)**

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O;\quad \Delta H = -x$$

$$CH_3OH + \frac{3}{2}O_2 \rightarrow CO_2 + 2H_2O;\quad \Delta H = -y$$

$$-\qquad -\qquad -\qquad -\qquad +$$

$$CH_4 + \tfrac{1}{2}O_2 \rightarrow CH_3OH;\qquad \Delta H = y - x$$

$$\because\ y - x = -\text{ve}$$

$$\therefore\quad x > y$$

109 **(c)**

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

For $\Delta G = 0$ at equilibrium

$$T\Delta S = \Delta H$$

$$T = \frac{\Delta H}{\Delta S} = \frac{30 \times 10^3}{105} = 285.7 \text{ K}$$

110 **(b)**

The gives rise to cooling.

111 **(d)**

In every process randomness increases.

112 **(c)**

$\Delta H - \Delta U = \Delta nRT = 1 \times 2 \times 373 = 746$ cal.

113 **(b)**

In first case it was equal to ΔU.It II case it is ΔH. The two values are related by

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

114 **(a)**

$$-W = +2.303\, nRT \log \frac{V_2}{V_1}$$

$$-W = 2.303 \times \frac{16}{32} \times 300 \times 8.314 \log \frac{25}{5}$$

$$-W = 2.01 \times 10^3 \text{J}$$

115 **(a)**

0.2 mole of HNO_3 are neutralized by 0.2 mole of NaOH to give heat = 57 × 0.2 = 11.4 kJ

116 **(c)**

Heat of formation is given when compound is formed from its components.

117 **(b)**

The statement is definition of Hess's law.

118 **(a)**

When a gas undergoes adiabatic expansion, $dq = 0$, it gets cooled due to loss of kinetic energy

119 **(c)**

$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3; \Delta H° = -852$ kJ

$\because$ 2 mol Al (=54 g) evolved heat = -852 kJ

$\therefore$ 2.7 gAl will evolve heat = $-\frac{852 \times 2.7}{54}$

$= -42.6$ kJ

120 **(a)**

Decomposition of $CaCO_3$ is made on heating.

121 **(a)**

$\Delta G = \Delta H - T\Delta S = -2500 - 298 \times 7.4 = -4705.2$ kcal and thus, spontaneous.

122 **(a)**

$$q = \Delta U + W$$

$$50 = \Delta U + 30$$

$$\therefore \Delta U = 20 \text{ cal}$$

123 **(a)**

The compressor has to run for longer time releasing more heat to the surroundings

124 **(c)**

$S^{\circ}_V - S^{\circ}_L = +$ve. Conversion of liquid to vapour occurs process.

125 **(a)**

$$-\Delta G° = 2.303\, RT \log K_p$$

126 **(c)**

As graphite is thermodynamically more stable than diamond hence, more heat is required to convert graphite to gaseous carbon

127 **(c)**

Closed systems cannot exchange mass with surrounding. Only exchange of energy is possible.

128 **(a)**

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

at equilibrium, $\Delta G = 0, \therefore T = \frac{\Delta H}{\Delta S} = \frac{30\times10^3}{75} =$ 400 K

129 **(b)**

ΔH/mol of FeS$= \frac{3.77\times56}{2.1} = 100.5$

130 **(a)**

For isochoric process, $\Delta V = 0$ so, $q_v = \Delta E$ *ie*, heat given to a system under constant volume is used up in increasing ΔE

131 **(a)**

According to Hess's law total heat changes during a chemical reaction are independent of path of reaction.

Given, $I_2(s) \rightarrow I_2(g), \Delta H_1 = 57.3$ kJ/mol ...(i)

$I_2(s) \rightarrow I_2(l), \Delta H_2 = +15.5$ kJ/mol ...(ii)

Required equation $I_2(l) \rightarrow I_2(g), \Delta H_1 =?$

subtract

Eq. (ii) from Eq. (i)

$\therefore \quad I_2(l) \rightarrow I_2(g), \quad \Delta H = 57.3 + (-15.5)$

$= +41.8$ kJ/mol

132 **(a)**

$\Delta G = -2.303\, RT \log K$

$-4.606 = -2.303 \times 0.002 \times 500 \log K$

$\log K = 2, K = 100$

133 **(d)**

$\Delta S^o = 2S^o_{HCl} - (S^o_{H_2} + S^o_{Cl_2})$

$= 2 \times 186.7 - (130.6 + 223.0) = 19.8\ JK^{-1}mol^{-1}$

134 **(b)**

The energy required to break a bond.

135 **(c)**

25 Meq. Of acid and 25 Meq. of base (the maximum value of Meq.) react.

136 **(b)**

Eq.(b) shows largest phase change *ie*,gas → solid

137 **(b)**

$$\frac{1}{2}H_2 + \frac{1}{2}Cl_2 \rightarrow 2HCl$$

$$\Delta H_f = -e_{H-Cl} + \left[\frac{1}{2}e_{H-H} + \frac{1}{2}e_{Cl-Cl}\right]$$

$$= -431 + \left[\frac{1}{2} \times 434 + \frac{1}{2} \times 240\right]$$

$$= -93 \text{ kJ mol}^{-1}$$

139 **(b)**

$A_2 + E_{given} \rightarrow 2A + E_{left}$

$\therefore E_{left}$ per molecule

$= E_{given} - E_{used}$ for dissociation

$= 4.4 \times 10^{-19} - 4.0 \times 10^{-19}$

$= 4 \times 10^{-20}$ J

The kinetic energy per atom $= \frac{4 \times 10^{-20}}{2} = 2 \times 10^{-20}$ J

140 **(c)**

Macroscopic properties which determine the state of a system are referred as state functions. The change in the state properties depends only upon the initial and final state of the system. All thermodynamic functions are state functions except work and heat.

141 **(c)**

Temperature is a measure of intensity of energy, whereas heat is a measure of quantity of energy.

142 **(a)**

For exothermic reactions, K_{eq} varies inversely with T while in case of endothermic reactions, K_{eq} varies directly with T

143 **(b)**

$\Delta G = \Delta H - T\Delta S, T = 25 + 273 = 298$ K

$= -11.7 \times 10^3 - 298 \times (-105) = 19590$ J

$= 19.59$ kJ

144 **(c)**

Standard heat of formation of methane is represented by C(graphite)+2H(g) = $CH_4(g)$ because the elements taken are in their standard state

145 **(b)**

ΔH per 1.6 g $= \frac{72\times1.6}{180} = 0.64$ kcal

146 **(b)**

$(dS)_{V,E>0}(dG)_{T,P<0}$

147 **(a)**

According to second law of thermo chemistry the law states that the total heat change (ΔH) accompanying a chemical reaction is the same whether the reaction takes place in one or more steps.

It means that heat of a reaction depends only on the initial reactants and final products and not on intermediate products that may be formed.

Now, $\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$

Enthalpy change in a reaction is always constant and independent of the path followed.

148 **(b)**

Every system having some quantity of matter, is associated with a definite amount of energy. This energy is known as internal energy. It is sum of many type of energies, such as translation energy, rotational energy, vibrational energy, electronic energy and bonding energy of the molecule.

$$E = E_{\text{trans}} + E_{\text{rot}} + E_{\text{vib}} + E_{\text{bonding}} + E_{\text{electronic}}$$

149 **(c)**

For an isochoric process V = constant ; thus $\Delta V = 0$ from 1st law $q = \Delta U - W$

or $q_v = \Delta U$.

150 **(a)**

For a spontaneous process $\Delta G = -$ve or < 0.

151 **(d)**

It is a physical change.

152 **(a)**

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$; $\Delta H = -188$ kJ/mol ...(i)

$H_2 + O_2 \rightarrow H_2O_2$; $\Delta H = -286$ kJ/mol ...(ii)

By Eq.(i) − (ii)

$2H_2O_2 \rightarrow 2H_2O + O_2$;

$\Delta H = 2(-188) - 2(-286) = +196$ kJ

153 **(a)**

$C + O_2(g) \longrightarrow CO_2(g)$

$\Delta n = 0$

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

154 **(a)**

$\Delta H = +$ve and $\Delta S = +$ve; the disorder increases with increase in moles.

$PCl_5 \rightleftharpoons PCl_3 + Cl_2; \Delta H = +$ve (dissociation)

155 **(b)**

$q = \Delta U - W; -W$ is work done by the system

$\therefore \Delta U = q + W = 50 + 10 = 60$

($+W$ is work done on the system)

157 **(d)**

$\Delta G_{\text{system}} = -$ve, the system is spontaneous ;

$\Delta G_{\text{system}} = 0$, the system has attained equilibrium

$\Delta G = +$ve, the system is non-spontaneous.

158 **(d)**

$$2B + \frac{3}{2}O_2 \longrightarrow B_2O_3; \qquad \Delta_f H^{\circ}(B_2O_3)$$

$$\therefore \ \Delta_c H^{\circ}(B) = \frac{1}{2}\Delta_f H^{\circ}(B_2O_3)$$

The heat is given out.

159 **(d)**

At isothermal condition T = constant.

160 **(c)**

Heat changes at constant pressure are referred as ΔH. Heat changes at constant volume are referred as ΔU.

161 **(d)**

(i) $2C(s) + H_2(g) \rightarrow H{-}C \equiv C{-}H(g) \quad \Delta H = 225 \text{ kJ mol}^{-1}$

(ii) $2C(s) + 2C(g) \quad \Delta H = 1410 \text{ kJ mol}^{-1}$

(iii) $H_2(g) + 2H(g) \quad \Delta H = 330 \text{ kJ mol}^{-1}$

From equation (i) :

$$225 = [2 \times \Delta H_{C(s)\rightarrow C(g)} + 1 \times e_{H-H}] - [2 \times e_{C-H} + 1 \times e_{C\equiv C}]$$

$$225 = [1410 + 1 \times 330] - [2 \times 350 + 1 \times e_{C\equiv C}]$$

$$225 = [1410 + 330] - [700 + e_{C\equiv C}]$$

$$225 = 1740 - 700 - e_{C\equiv C}$$

$$e_{C\equiv C} = 1040 - 225 = 815 \text{ kJ mol}^{-1}$$

$$BE_{C\equiv C} = 815 \text{ kJ mol}^{-1}$$

162 **(a)**

Heat capacity of water per gram $= \frac{75}{18} = 4.17$ J

$Q = mst = 100 \times 4.17 \times t = 1000$

$$t = \frac{1000}{100 \times 4.17} = 2.4 \text{ K}$$

163 **(c)**

When gas is compressed its entropy decreases so, ΔS is negative

164 **(b)**

Bond energy of C − H bond $= \frac{-166}{4} = -41.5$ kJ/mol

165 **(b)**

$$-\frac{dQ}{dW} = \frac{dQ}{dQ - dE}; [\because dE = dW + dQ: -dW = dQ - dE]$$

$$= \frac{nC_p dT}{nC_p dT - nC_v dT} = \frac{C_p}{(C_p - C_v)}$$

$= \frac{7}{2}$ for diatomic gas

166 **(c)**

In an adiabatic process, no exchange of heat takes place between the system and surroundings, *i.e.,* $dQ = 0$. Such a condition exists when the system is thermally isolated.

167 **(c)**

$\Delta H = H_P - H_R$

$H_{H_2O} - H_{H_2} - H_{O_2} = -ve$

168 **(d)**

In the adiabatic process no heat enters or leaves the system *i. e.*, $q = 0$.

169 **(c)**

An isolated system neither shows exchange of heat nor matter with surroundings.

170 **(a)**

The exact value of internal energy is not known as it includes all type of energies of molecules constituting the given mass of matter such as translational, vibrational and rotational. The kinetic and potential energy of the nuclei and electron with in the individual molecules and the manner in which the molecule are linked together, are

$E = E_{\text{translational}} + E_{\text{rotational}} + E_{\text{vibrational}}$

Thus, we can say that internal energy is partly potential ad partly kinetic

171 **(a)**

Heat of formation of NH_3 is given by

$\frac{1}{2}N_2(g) + \frac{3}{2}H_2(g) \longrightarrow NH_3(g);\ \Delta H_f$

172 **(d)**

By (ii)-(i) C(graphite) → C (diamond);

$\Delta H = -393.4 - (-395.3) = +1.9$

173 **(b)**

CH_4 required $= \frac{445.15 \times 16}{890.3} = 8$ g

174 **(d)**

Calorific value : Fat > Carbohydrate > Protein

175 **(d)**

$\frac{1}{2}N_2 + O_2 \rightarrow NO_2;\ \Delta H = 8$ kcal …… (i)

$N_2 + 2O_2 \rightarrow N_2O_4;\ \Delta H = 2$ kcal ……… (ii)

By eq. (i) × 2 – (ii) ,

$2NO_2 \rightarrow N_2O_4;\ \Delta H = -14.0$ kcal

176 **(c)**

Calculate it for the equation

$2Fe(s) + \frac{3}{2}O_2(g) \rightarrow Fe_2O_3(s)$

Eqs. 3(ii) – (i), we have

$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

$\Delta H^\circ = 3(-94050) - (-93657)$

$= -18493$ cal $= 188.493$ kcal

$\because$ 2 moles Fe_2O_3 has $\Delta H^\circ = -188.493$ kcal

$\therefore$ 1 mole Fe_2O_3 has$\Delta H^\circ = -\frac{188.493}{2}$

$= -94.25$ kcal/mol

177 **(d)**

At STP, 16 g O_2 or $\frac{1}{2}$mole O_2 will occupy 11.2 litre.

Thus, if volume is doubled, it means

$(V_2 - V_1) = 22.4 - 11.2 = 11.2$ litre

Now, $W = P \times (V_2 - V_1) = 1 \times 11.2$ litre atm

$= \frac{1 \times 11.2 \times 2}{0.0821}$

$= 272.84$ kcal

178 **(d)**

For isothermal process $\Delta U = 0$.

179 **(b)**

Gibbs energy change ΔG is given by:

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

Also, $G = H - TS$

180 **(a)**

For insulated container $q = 0$.

181 **(b)**

$\Delta_{\text{solution}} = \Delta H_i + \Delta H_h$

or $1 = 180 + \Delta H_h$

$\Delta H_h = -179$ kcal mol^{-1}

The total $\Delta H_h = \Delta H_{h_{Na^+}} + \Delta H_{h_{Cl^-}} = \frac{61}{11} + \frac{5a}{11} = a$

Where a is total heat of hydration (ΔH_h).

Thus, $\Delta H_{h_{Na^+}} = -\frac{6 \times 179}{11} = -97.63$ kcal mol^{-1}

182 **(a)**

Since, process is exothermic, heat is evolved, due to this temperature of water increases

183 **(a)**

$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$

This equation can be obtained by subtraction of

$$\left[CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)\right]$$

from $[C(s) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)]$

Hence, $\Delta H_f = (CO) = [-393.3 - (-282.8)]$ kJ

$= 110.5$ kJ/mol

184 **(a)**

Heat of neutralisation is also defined as the heat of formation of H_2O from H^+ and OH^- ions.

185 **(b)**

$q = \Delta U - W$, if $q = 0$ for adiabatic process, than $-\Delta U = -W$ or $\Delta U = W$, *i. e.*, work done on the system or work of compression brings in an increase in temperature

186 **(c)**

$PV^r =$ constant

${}_{P\cdot\gamma}V^{\gamma-1}dV + V^\gamma \cdot dP = 0$

$\therefore \frac{dP}{P} = \frac{{}_{\gamma .}V^{r-1}dV}{V^\gamma} = -\gamma\left(\frac{dV}{V}\right)$

187 **(d)**

Strong acid (HCl) and strong base (NaOH).

188 **(d)**

The randomness in gaseous state is more than liquid state.

189 **(c)**

The definition of third law of thermodynamics.

190 **(b)**

$\Delta S = +$ve for irreversible process.

191 **(a)**

It is the definition of heat of formation.

192 **(c)**

This is definition of third law of thermodynamics.

193 **(c)**

Follow Le-chatelier principle.

194 **(a)**

$\Delta S = S_P - S_R$

$= (2 \times 0.19) - 0.13 - 0.22$

$= 0.03 \text{ kJ K}^{-1}\text{mol}^{-1}$

$= 30 \text{ JK}^{-1}\text{mol}^{-1}$

196 **(a)**

$W_{\text{exp}} = -P \times \Delta V$

$= -1 \times (13 - 3) = -10 \text{ atm dm}^3$

197 **(a)**

Heat evolved during combustion of 3.2 g

$\text{CH}_4 = \dfrac{880 \times 3.2}{16} = -176 \text{ kJ}$

198 **(c)**

It is defined of heat of solution.

199 **(b)**

Formation of CO_2 from CO is an exothermic reaction. Heat is evolved from the system *ie*, energy is lowered thus exothermic reactions occur spontaneously on account of decrease in enthalpy of system. Thus, $\Delta E > \Delta H$

200 **(d)**

$\text{dQ} = \text{nS } \Delta\text{T}, \qquad \therefore \text{S} = \frac{\text{dQ}}{\text{dT}}$ (for 1 mole).

201 **(c)**

$\Delta H = nC_p\Delta T$

The process is isothermal therefore,

$\Delta G = 0$

$\therefore \Delta H = 0$

202 **(b)**

The system returns to its original state, *i.e.*, cyclic process.

203 **(b)**

$\Delta G = \Delta H - T\Delta S$; at equilibrium,

$\Delta G = 0, \qquad \therefore \Delta H = T\Delta S$

or $\Delta H = 273 \times (60.01 - 38.20)$

$= 5954.13 \text{ J mol}^{-1}$

204 **(a)**

$EN_{\text{F}} \sim \text{EN}_{\text{Cl}} = 0.2028\sqrt{\Delta}$

and $\quad \Delta = \left[e_{\text{F-Cl}} - (e_{\text{F-F}} \times e_{\text{Cl-Cl}})^{1/2}\right]$

$\therefore EN_{\text{F}} \sim \text{EN}_{\text{Cl}} = 0.2028\left[e_{\text{F-Cl}} - (e_{\text{F-F}} \times e_{\text{Cl-Cl}})^{1/2}\right]^{1/2}$

Or $1 = 0.2028\left[e_{\text{F-Cl}} - (38 \times 58)^{1/2}\right]^{1/2}$

$\therefore \; e_{\text{F-Cl}} = 71.26 \text{ kcal mol}^{-1}$.

206 **(b)**

0.2 mole will neutralize 0.2 mole of HNO_3 heat evolved $= 51 \times 0.2 = 11.4$ kJ

207 **(b)**

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

208 **(d)**

Δn depends on stoichiometry of reaction.

209 **(a)**

$e_{A-A} = a$ Also, $\frac{1}{2}A_2 + \frac{1}{2}B_2 \to AB;$

$e_{A-B} = a \; \Delta H = -100 \text{ kJ mol}^{-1}$

$e_{B-B} = 0.5a$

$\therefore \quad \Delta H = -[e_{A-B}] + \dfrac{1}{2}[e_{A-A} + e_{B-B}]$

$= a + \dfrac{1}{2}[a + 0.5a]$

$-100 = -0.25\,a$

$\therefore \quad a = 400 \text{ kJ mol}^{-1}$

210 **(d)**

The properties of the system whose value is independent of the amount of substance present in the system are called intensive properties *e.g.,* viscosity, surface tension, temperature, pressure etc.

211 **(d)**

When a real gas is forced through a porous plug into a region of low pressure, it is found that due to expansion, the gas on the side of low pressure gets cooled

212 **(b)**

The room got heated because heat is lost to surroundings.

213 **(b)**

$T_b = \dfrac{\Delta H}{\Delta S} = \dfrac{30 \times 10^3}{75} = 400 \text{ K}$

214 **(d)**

Heat of combustion is always exothermic; Few combustion reactions such as F_2 to F_2O, N_2 to N_2O and NO are endothermic but these reactions do not give heat of combustion because the substance should be completely oxidized. In F_2O, F_2 is reduced and N_2O and NO are not completely oxidized state of N_2. However, three reactions are exceptions but these do not represent heat of combustion. These are,

$N_2 + O_2 \longrightarrow N_2O;\quad \Delta H = +\text{ve}$

$N_2 + O_2 \longrightarrow NO;\quad \Delta H = +\text{ve}$

and $F_2 + (1/2)O_2 \longrightarrow F_2O;\quad \Delta H = +\text{ve}$

215 **(b)**

For an isothermal process $\Delta T = 0$ and $\Delta E = 0$ and $q \neq 0$.

216 **(b)**

Given: (i)$H_2 + \frac{1}{2}O_2 \rightarrow H_2O;\ \Delta H = -241$ kJ

(ii)$C_6H_{10} + \frac{17}{2}O_2 \rightarrow 6CO_2 + 5H_2O;\ \Delta H = -3800$ kJ

(iii)$C_6H_{12} + 9O_2 \rightarrow 6CO_2 + 6H_2O;\ \Delta H = -3920$ kJ for the reaction

$C_6H_{10} + H_2 \rightarrow C_6H_{12}$

[It is infact Eq.(i)+Eq.(ii)−Eq.(iii)]

Thus, $\Delta H = -241 - 3800 - (-3920) = -121$ kJ

217 **(d)**

In isothermal reversible process, ideal gas has constant volume and so, $\Delta E = 0$ and $\Delta H = \Delta E = 0$

218 **(a)**

$\Delta H = -2 \times e_{H-Cl} + e_{H-H} + e_{Cl-Cl}$

$\therefore\quad n182 = -2 \times a + 430 + 242$

$\therefore\quad a = 245 \text{ kJ mol}^{-1}$

219 **(d)**

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

$\Delta n = +1/2$

Thus, $\Delta H > \Delta U$

220 **(c)**

Cylinder contains 11.2 kg or 193.10 mole butane.

($\because$ molecular mass of butane =58)

$\because$ Energy released by 1 mole of butane = −2658

$\therefore$ Energy released by 193.10 mole of butane

$= -2658 \times 193.10$

$= 5.13 \times 10^5$ kJ

$\therefore \dfrac{5.13 \times 10^5}{20000} = 25.66$ or 26 days

221 **(c)**

Heat of formation of H_2O = − heat of decomposition of water.

222 **(a)**

$T_{f_{irreversible}} > T_{f_{reversible}}$ it is an adiabatic expansion and W(rev) is maximum.

223 **(c)**

Molecular solids are covalent compounds having low m.p.

224 **(a)**

$\Delta H = H_P - H_R$

Thus, ΔH is negative because $H_P < H_R$.

225 **(b)**

$\Delta G = -$ve for a spontaneous change.

226 **(d)**

Ideal gas does not show intermolecular forces of attractions.

227 **(b)**

Rest all are correct .

228 **(a)**

During solidification disorder decreases.

229 **(a)**

$$\Delta S = \frac{\Delta H_f}{T} = \frac{2930}{300} = 9.77 \text{ J mol}^{-1}\text{K}^{-1}$$

230 **(d)**

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

The reaction will be spontaneous

If $T\Delta S > \Delta H \quad (i.e., \Delta G = -\text{ve})$

$$T > \frac{\Delta H}{\Delta S} = \frac{170}{170 \times 10^{-3}} = 1000 \text{ K}$$

231 **(c)**

θ is independent of initial amount as long as relative amount is constant

232 **(b)**

$q = \Delta U - W$, if adiabatic process $q = 0$, then $-\Delta U = -W$, *i.e.*, a decrease in free energy brings in work done by the system $(-W)$.

233 **(a)**

As the system is closed and insulated, no heat enter or leave the system, *ie*, $q = 0$

$\therefore \Delta E = q + W = W$

234 **(b)**

$XY \rightarrow X(g) + Y(g);\ \Delta H = +a$ kJ/mol ...(i)

$X_2 \rightarrow 2X;\ \Delta H = +a$ kJ/mol ...(ii)

$Y_2 \rightarrow 2Y;\ \Delta H = +0.5a$ kJ/mol ...(iii)

$\frac{1}{2} \times$ (ii) $+ \frac{1}{2} \times$ (iii) − (i) gives

$$\frac{1}{2}X_2 + \frac{1}{2}Y_2 \rightarrow XY$$

$$\Delta H = \left(+\frac{a}{2} + \frac{0.5}{2}a - a\right) \text{kJ/mol}$$

$$\therefore\ -200 = +\frac{a}{2} + \frac{0.5a}{2} - a$$

or $a = 800$

235 **(d)**

$CH_4 \rightarrow C + 4H;\ \Delta H = 360$ kcal/mol

$e_{C-H} = 90$ kcal

$C_2H_6 \rightarrow 2C + 6H;\ \Delta H = 620$ kcal/mol

$\therefore\quad 620 = e_{C-C} + 6e_{C-H}$

$\therefore\quad e_{C-C} = 620 - 540 = 80$ kcal/mol

236 **(d)**

Molecular weight of $NH_4NO_3 = 80$

$\therefore$ Molar heat of decomposition

$H = ms\Delta t = 80 + 1.23 \times 6.12$

$= 602$ kJ/mol

237 **(b)**

Greater is bond energy more is stability to bond.

238 **(a)**

Due to high bond energy of $N \equiv N$, more heat is absorbed to break up N_2 molecule.

239 **(a)**

$$\Delta S_{vap} = \frac{(900 \times 18)}{373} = 43.4 \text{ JK}^{-1}\text{mol}^{-1}$$

240 **(b)**

For spontaneous reaction $\Delta G = -$ve.

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

$\Delta H = +$ve, $\Delta S = +$ve and $T\Delta S > \Delta H$

241 **(d)**

$\Delta n = 0$

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

242 **(a)**

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O(g);\ \Delta H = -57.0$ kcal ..... (i)

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O(l);\ \Delta H = -68.3$ kcal ... ... (ii)

By eq. (i) and (ii),

$H_2O(l) \rightarrow H_2O(g);\quad \Delta H = +11.3$ kcal

243 **(b)**

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

$\therefore\ \Delta U = 176 - 1 \times 8.314 \times 1240 \times 10^{-3}$

$= 165.6\ kJ$

244 **(a)**

$TV^{\gamma-1}$ =constant

$$\frac{T}{T_{\text{final}}} = \left(\frac{V_2}{V_1}\right)^{\gamma-1}$$

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

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

246 **(b)**

$\Delta n = -2$

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

$= -1415 + (-2) \times 0.0083 \times 300$

$= -1420$ kJ

247 **(c)**

Experimental determination of heats of reaction by bomb calorimeter represents its value at constant volume, i.e., ΔU.

248 **(b)**

Graphite possesses sp^2-hybridisation and has flat layer structure whereas diamond possesses sp^3-hybridisation and has rigid tetrahedral nature.

249 **(c)**

$$n_{\text{efficiency}} = \frac{T_2 - T_1}{T_2}$$

$$\text{or } 0.25 = \frac{T - 400}{T}$$

$\therefore T = 533.3$ K

250 **(a)**

Lower is energy level of a system, more is its stability.

251 **(b)**

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

Since, $\Delta n = -2$

Thus, $\Delta H < \Delta U$

252 **(b)**

$K + \frac{1}{2}O_2 + \frac{1}{2}H_2 \rightarrow KOH;\ \Delta H = ?$

Find ΔH by Eq. [(i) + (ii)] – (iii).

254 **(c)**

The fact for a quantity referred as state function.

255 **(c)**

Bond formation is always exothermic.

256 **(d)**

$N_2 + \frac{1}{2}O_2 \longrightarrow N_2O;\quad \Delta H = 28$ kJ

$\frac{1}{2}N_2 + \frac{1}{2}O_2 \rightarrow NO;\ \Delta H = 90$ kJ

By eq. $[4 \times (ii)] - [2 \times (i)]$,

$2N_2O + O_2 \rightarrow 4NO;\ \Delta H = 304$ kJ

257 **(b)**

Calorific value = Heat of combustion per g of fuel, i.e., for C_2H_4, it is $\frac{-1411}{28}$, the lowest value.

258 **(b)**

$\Delta G° = \Delta H° - T\Delta S°$

$= -54.07 - 298 \times 10 \times 10^{-3}$

$= -57.05$ kJ

Also, $\Delta G° = 2.303\ RT \log_{10} K$

$$\log_{10} K = \frac{-57.05 \times 10^3}{2.303 \times 8.314 \times 298}$$

259 **(b)**

Hess's law is based upon law of conservation of energy *i.e.*, first law of thermodynamics.

260 **(d)**

$$\Delta S_f = \frac{\Delta H_f}{T} = \frac{6 \times 10^3}{273} = 21.98 \text{ J}$$

261 **(b)**

Properties which are mass independent are intensive properties and others which are mass dependent are extensive properties.

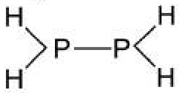
262 **(a)**

$\Delta n = 2 - 3 = -1 \quad \therefore\ \Delta H = \Delta U - RT$

263 **(b)**

Step1. P – H bond energy from bond dissociation energy of $PH_3(g)$ containing 3 such P – H bonds $= \frac{228}{3} = 76$ kcal/mol

Step 2. The structure of P_2H_4 is



ie, it contains 4 P – H bonds and one P – P bond, so P – P bond energy can be calculated by $4 \times$ P – H + P – P = bond dissociation energy P_2H_4

$\therefore$ P – P bond energy = 335–4(76)

= 31 kcal per mol

264 **(c)**

By Eq. (i) + (ii),

$2Na + Cl_2 \rightarrow 2NaCl;\ \ \Delta H = -196$ kcal

265 **(d)**

Energy of red P is lesser than white P and thus, red P is more stable.

266 **(b)**

$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$

$\Delta H = 2 \times (-13.7)$ kcal

267 **(d)**

$H_2O(l) \rightleftharpoons H_2O\ (v);\ \ \Delta H = 41.0$ kJ mol^{-1}

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

$41.0 = \Delta U + 1 \times 8.314 \times 373 \times 10^{-3}$

$\therefore\ \ \Delta U = 37.89$ kJ mol^{-1}

268 **(d)**

These are derived formulae.

269 **(c)**

Spontaneous process shows a decrease in ΔG.

271 **(d)**

For monoatomic gases $C_P : C_V$ is 1.67.

272 **(c)**

Bond formation is always exothermic.

273 **(b)**

Lower is heat of neutralisation, more is dissociation energy, weaker is acid.

274 **(b)**

Heat change for 3.6 g $H_2O = \frac{68\ \ 3.6}{18} = 13.6$ kcal

275 **(a)**

$$\because\ \Delta T = \frac{W}{nR} \therefore Q = nC_p(\Delta T) = nC_p\frac{W}{nR} = \frac{C_pW}{R}$$

$$C_p = \frac{QR}{W} = \frac{500 \times 2}{142.8} = 7$$

$C_p = 7$indicates that the gas is diatomic. Thus, it should be O_2

276 **(b)**

For the equation

$B_2H_6(g) + 3O_2(g) \rightarrow B_2O_3(g) + 3H_2O(g)$

Eqs.(i) +3(ii)+3(iii)–(iv)

$\Delta H = -1273 + 3(-286) + 3(44) - 36$

$= -1273 - 858 + 132 - 36$

$= -2035$ kJ/mol

277 **(d)**

As we know that,

Work done$(W) = 2.303\ nRT \log\frac{V_2}{V_1}$.

Hence, $V_1$and V_2 are in ratio in the relation. So, unit may be expressed in any one of m^3, dm^3 or cm^3.

278 **(a)**

Fuel value per g of substance produced on burning.

279 **(a)**

T_1 and T_2 are same for a substance.

280 **(a)**

For an endothermic reactions ΔH is positive because in endothermic reaction heat is always absorbed.

281 **(a)**

Ice takes up heat to melt and thus, enthalpy change is +ve.

283 **(a)**

When $\Delta H = -$ve, $\Delta S = +$ve and $\Delta G = -$ve than reaction is spontaneous

284 **(d)**

KE = (3/2)RT

285 **(b)**

$C(s) + O_2(g) \rightarrow CO_2(g); \Delta H = r$

...(i)

$CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g);\ \Delta H = s$

...(ii)

$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g);\ \Delta H = ?$

Subtract Eq. (ii) from Eq. (i)

$C(s) + O_2(g) \rightarrow CO_2(g); \Delta H = r$

$CO(g) + \frac{1}{2}O_2 \rightarrow CO_2(g);\ \Delta H = s$

$-\qquad -\qquad -\qquad -$

$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g);\ \Delta H = r - s$

286 **(d)**

$Cu(g) \rightarrow Cu^+(g) + e,\qquad \Delta H = 745$ kJ mol^{-1}

$I(g) + e \rightarrow I^-(g);\qquad \Delta H = -295$ kJ mol^{-1}

Adding $Cu^+(g) + I^-(g) \rightarrow CuI(g);\ \Delta H^\circ = -446$ kJ

$Cu(g) + I(g) \rightarrow CuI(g);\qquad \Delta H^\circ = 4$ kJ mol$^-$

287 **(a)**

Entropy of universe is tending towards maximum.

289 **(a)**

$\Delta H_{(\text{reaction})} = \Delta H_{f(\text{diamond})} - \Delta H_{f(\text{graphite})}$
$= 1.896 -$
$0.23 = 1.666$ kJ/ mol

290 **(b)**

p=1 atm

$$\Delta V = (50 - 15) = 35 \text{ L}$$

$$\therefore \quad W = -p.\Delta V = -1 \times 35$$

$$= -35 \text{ Latm}$$

Hence, work done by the system on the surroundings is equal to 35 L-atm.

291 **(d)**

The product possesses maximum energy and thus, least stable.

292 **(d)**

By eq. [(i) + 2 × (ii)] – (iii),

$$\text{C} + 2\text{H}_2 \longrightarrow \text{CH}_4;\ \Delta H = 74.1\ kJ$$

293 **(a)**

For the equation,

$\text{H}_2 + \text{S} + 2\text{O}_2 \rightarrow \text{H}_2\text{SO}_4$

Eqs. (i) +(ii)+(iii)+(iv)

$\Delta H = -287.3 + (298.2) + (-98.7) + (-130.2)$

$= 814.4$ kJ

294 **(d)**

(a) For isochoric process, $\Delta V = 0$

$$W = p\Delta V = 0$$

$$\therefore \quad \Delta E = Q$$

(b) For adiabatic process, $Q = 0$

$$\Delta E = W$$

(c) For isothermal process, $\Delta T = 0$

$$\text{and} \quad \Delta E = 0$$

$$Q = -W$$

(d) For cyclic process, state functions like

$$\Delta E = 0$$

$$Q = -W$$

296 **(a)**

$\Delta G = \Delta H - T\Delta S = -\text{ve} - \text{ve} = -\text{ve}$

297 **(a)**

$$\text{F}_2 + \frac{1}{2}\text{O}_2 \longrightarrow \text{F}_2\text{O};\quad \Delta H = +\text{ve}.$$

298 **(c)**

Two equivalent of each are used.

299 **(a)**

Isothermally (at constant temperature) and reversible work.

$$W = 2.303\ nRT \log \frac{p_2}{p_1}$$

$$= 2.303 \times 1 \times 2 \times 300 \log \frac{10}{2}$$

$$= 2.303 \times 600 \times \log 5 = 965.84$$

At constant temperature, $\Delta E = 0$

$$\Delta E = q + W, q = -W = -965.84 \text{ cal}$$

300 **(d)**

The gaseous phase have more entropy and thus, ΔS is +ve in (a) and (b). Also decrease in pressure increases disorder and thus, ΔS is +ve in (c). In (d) the disorder decreases in liquid state due to decrease in temperature. Thus, $\Delta S = -$ve.

301 **(a)**

Hess's law states that the total change in heat enthalpy during the complete course of reaction is same, whether the change is brought in one step or in several steps by one method or other method.

302 **(c)**

First we calculate the expected bond dissociation energy of benzene molecules as

$3 \times \text{C} - \text{C} + 3 \times \text{C} = \text{C} + 6 \times \text{C} - \text{H}$

∴ Calculated value = 3(347.3) + 3(615) + 6(412.2)

= 4397.8

Resonance energy = Experimental value – calculated value

= 5335 – 4397.8

= 937.2 kJ per mol

303 **(d)**

$$\Delta S = 2.303nR \log \frac{V_2}{V_1}$$

$$= 2.303 \times 2 \times 2 \log \frac{20}{2} = 9.2$$

304 **(b)**

$$\left.\begin{array}{c}\text{Work done by the system or}\\ \text{work of expansion is negative.}\\ \text{Work done on the system or}\\ \text{work of compression positive.}\end{array}\right]\text{The modern concept.}$$

305 **(a)**

From first law of thermodynamic.

$\Delta E = q + W$ Given, $q = +300$ cal

(∵Heat is absorbed)

$W = -500$ cal

(∵Work is done on surroundings)

$\therefore \quad \Delta E = q + W = 300 + (-500)$

$= -200$ cal

306 **(b)**

An experimental fact.

307 **(b)**

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

$\Delta n_g = 1 - 3 = -2$

We know that,

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

$\therefore \quad \Delta H = (-885389) - (-2) \times 8.314 \times$ 298

$= -885389 + 4955.1440$

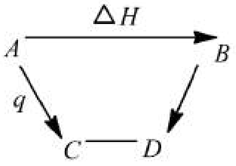
$= -880433.86 \text{ J mol}^{-1}$

308 **(a)**

Human body is an example of open system as it can exchange both mass and energy with the surroundings.

309 **(c)**

According to Hess's law, the total heat changes occurring during a chemical reaction are independent of path.



$\Delta H = q + V + 2x$

311 **(c)**

$\because q_{abs} = \Delta U + (-W)$

$\therefore \Delta U = q + W$; ΔU is state function.

312 **(c)**

For exothermic reaction, $\Delta H = (-)$ for endothermic reaction, $\Delta H = (+)$.

313 **(b)**

Find ΔH for,

$Ca + O_2 + H_2 \rightarrow Ca(OH)_2$

314 **(b)**

For maximum extent of reaction,

$4X(s) + O_2(g) \longrightarrow 2X_2O(s); \quad \Delta H = a$

Also, $X(s) + \frac{1}{4}O_2(g) \longrightarrow \frac{1}{2}X_2O(s); \quad \Delta H = -90$ kJ

$\therefore \quad a = -90 \times 4 = -360$ kJ

315 **(c)**

For spontaneous process $\Delta G = -$ve

316 **(a)**

$$\begin{array}{l}
S + \frac{3}{2}O_2 \rightarrow SO_3;\ \Delta H = -2x \\
SO_2 + \frac{1}{2}O_2 \rightarrow CO_2 + SO_3;\ \Delta H = -y \\
- \quad - \qquad\qquad - \qquad + \\
\hline
S + O_2 \rightarrow SO_2; \qquad \Delta H = (y - 2x)
\end{array}$$

317 **(a)**

Given, $2C + 3H_2 \rightarrow C_2H_6$; $\Delta H = -21.1$

$C + O_2 \rightarrow CO_2$; $\Delta H = -94.1$

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$; $\Delta H = -68.3$

Eqs .2(ii) + 3(iii)−(i)

$C_2H_6 + \frac{3}{2}O_2 \rightarrow 2CO_2 + 3H_2O$

$\Delta x = 2(-94.1) + 3(-68.3) - (-21.1)$

$= -372$ kcal

318 **(c)**

Surface tension is an intensive property because it does not depend upon the quantity of matter present in the system

319 **(a)**

$$\frac{1300}{241.8} = \frac{5.37}{1}$$

320 **(d)**

$PV = 1 \times 1$ lit − atm

$= 10^{-3}\text{m}^3 \times 0.76 \times 13.6 \times 9.8 \times 10^3 \text{Nm}^{-2}$

$= 101.3$ J

321 **(c)**

$Ag^+ + Cl^- \longrightarrow AgCl$ is a spontaneous reaction.

322 **(b)**

$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$

Thus, V_{O_2} used $= \frac{6226 \times 3 \times 22.4}{1411}$

$= 296.5$ litre

323 **(c)**

CS_2 is formed from its initial components carbon and hydrogen.

324 **(d)**

$C + O_2 \rightarrow CO_2(g); \qquad \Delta H = -94$ kJ ..... (i)

$Ca + \frac{1}{2}O_2 \rightarrow CaO(s); \qquad \Delta H = -152$ kJ .... (ii)

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$; ΔH

$= 42$ kJ .... (iii)

By eq. [(i) + (ii)]- (iii),

$Ca + C + \frac{3}{2}O_2 \rightarrow CaCO_3$; $\Delta H = -288$ kJ.

325 **(a)**

$PV^\gamma =$ constant for adiabatic expansion and

$PV =$ constant for isothermal expansion

$\therefore \log P = -\gamma \log V$ slope $= -\gamma$
$\log P = -\log V$ slope $= -1$

326 **(c)**
A part of heat is used in dissociation of CH_3COOH, a weak acid.

327 **(a)**
No doubt solidification shows a decrease in entropy but in egg proteins structure are disordered in solid state due to denaturation.

328 **(a)**

$$1 \text{ cal} = 4.18 \text{ J} = 4.18 \times 10^7 \text{ erg}$$
$$= \frac{4.18}{1.602} \times 10^{19} \text{ eV}$$

329 **(c)**
$C + O_2 \rightarrow CO_2$; $\Delta H = -393.5$ kJ/mol
$\because$ 44 g of CO_2 formed by which heat released $= -393.5$ kJ
$\therefore$ 1 g of CO_2 formed by which heat released $= \frac{-393.5}{44}$
$\therefore$ 35.2 g (given) of CO_2 formed by which heat released

$$= -\frac{393.5}{44} \times 35.2 = -314.8 \text{kJ}$$

330 **(b)**
Only work can be done by a thermally isolated system between it and surroundings.

331 **(c)**
An open system is one which involves exchange of mass and energy.

333 **(c)**
Gaseous molecules have more random motion.

334 **(d)**

$$H_2 + \frac{1}{2}O_2 \rightarrow H_2O(l); \quad \Delta H = -68.32 \text{ kcal}$$
$$H_2O(l) \rightarrow H_2O(g); \quad \Delta H = 10.52 \text{ kcal}$$
$$\therefore H_2 + \frac{1}{2}O_2 \rightarrow H_2O(g); \ \Delta H = -57.80 \text{ kcal}$$

335 **(c)**
$\Delta G = \Delta H - T\Delta S$
$\Delta G = 0, \therefore \Delta H = T\Delta S$

$$T = \frac{\Delta H}{\Delta S} = \frac{40.63 \times 10^3}{108.8} = 373.4 \text{ K}$$

336 **(b)**
No exchange of heat in between system and surroundings under adiabatic conditions.

337 **(b)**
More is heat of ionization of acid more is stability of acid or lesser is dissociation or $K_{aCH_3COOH} > K_{aHCN}$. Thus, $pK_{a_{HCN}} > pK_{a_{CH_3COOH}}$

338 **(a)**
Bond breaking process or decomposition processes are endothermic process.

339 **(a)**
Hess's law states that enthalpy changes during and process are independent of path. So, this law is used in calculating enthalpy.

340 **(d)**
$\Delta G = \Delta H - T\Delta S$: $\Delta G = \Delta E + p\Delta V - T\Delta S$
For spontaneity $\Delta G = -$ve

341 **(b)**
$H^+ + OH^- \longrightarrow H_2O$; $\Delta H = -13.7$ kcal
Also, $\Delta H = H_f^\circ H_2O - [H^\circ_{H^+} + H^\circ_{OH^-}]$
Or $-13.7 = -68.0 - [0 + H^\circ_{OH^-}]$ [$\therefore$ $H^\circ_{H^+} = 0$]
$\therefore$ $H^\circ_{OH^-} = -54.3$ kcal

342 **(a)**
This is the derived formula for W_{rev} is isothermal change.

343 **(b)**
Internal energy depends upon the temperature of gas and not on P and V.

344 **(a)**
T_A and T_B are same for a liquid.

345 **(d)**

$$\Delta H = -2 \times [2 \times e_{O-H}] + 2 \times e_{H-H} + e_{O-O}$$
$$= -4 \times 220 + 2 \times 105 + 120 = -550 \text{ kJ}$$

346 **(c)**
Heat measurements are made in calorimeter usually made of copper.

347 **(d)**
The process involves conversion of 1 mole of C(s) to C(g), $i.e.$, sublimation.

348 **(d)**
$C + 2H_2 \rightarrow CH_4$; $\Delta H = -17.9$ kcal ……. (i)
$C + O_2 \rightarrow CO_2$; $\Delta H = -94.1$ kcal ……. (ii)
$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$; $\Delta H = -68.3$ kcal …… (iii)
Eqs. [(ii) + 2 × (iii)] – (i),
$CH_4 + 2O_2 \rightarrow 2H_2O$

350 **(a)**
The branch deals with interconversion of heat and chemical energy.

351 **(d)**
$2H_2 \rightarrow 4H$; $\Delta H = -869.6$ kJ
$\therefore e_{H-H} = \frac{969.6}{2} = +434.8$ kJ

353 **(c)**

The properties, which do not depend on the amount of substance, are called intensive property. *e.g.*, surface tension, viscosity etc.

354 **(a)**

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

$\Delta n = -3$

355 **(a)**

$W_{\text{rev}} = -\int PdV\ or -\int P\Delta V$; note that opposing pressure is not constant throughout .

356 **(b)**

Joule-Thomson coefficient

$$\mu = \frac{dT}{dP} = \frac{27-30}{5-2} = -1$$

For all negative values of μ,the gas warms on expansion

357 **(b)**

$W = -p\Delta V$

Given, $p + 100\text{kPa} = 10^5$ Pa,

$V_1 = 1\text{dm}^3 = 10^{-3}\text{m}^3, V_2 = 1\text{m}^3$

$W = 10^5 \times (1 - 10^{-3})$ J

$\therefore \quad W = 99900$ J

358 **(d)**

The efficiency of engine is given as,

$\eta = \frac{T_2 - T_1}{T_2}$; η is more when $T_2 - T_1$ is maximum.

360 **(d)**

A spontaneous change is accompanied by lowering of free energy

361 **(c)**

E and $G(H - TS)$ are state functions. Also, $\Delta E = q + w$ is state function. But q and w are path dependent and not state functions.

362 **(c)**

An isolated system neither shows exchange of heat nor matter with surroundings.

364 **(b)**

No doubt (a) and (b) both represent heat of formations but standard heat of formation $(\Delta H^\circ f)$ for CO_2 will be from $C_{(\text{graphite})} + O_2 \rightarrow CO_2$ as $C_{(\text{graphite})}$ is most stable form of carbon.

365 **(a)**

For an isothermal process, $\Delta E = 0$

As the process is taking place at constant T and p hence, from equation,

$$\Delta H = \Delta E + \Delta p.V$$

We have, $\Delta H = 0 + 0 \times V = 0$

Hence, for the process, $\Delta H = \Delta E = 0$

366 **(b)**

At isothermal condition T = constant.

367 **(a)**

The heat of formation of CO is calculated by using Hess's law. According to it, the total heat changes occurring during a chemical reaction are independent of path.

$2CO(g) + O_2(g) \rightarrow 2CO_2(g); \quad \Delta H = -135.2$ kcal

(I)$CO_2(g) \rightarrow CO(g) + \frac{1}{2}O_2(g); \quad \Delta H = \frac{135.2}{2}$ kcal

(II) $C(s) + O_2(g) \rightarrow CO_2(g); \quad \Delta H = -94$ kcal

Required equation

$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g); \quad \Delta H = ?$

Add Eqs. (I) and (II)

$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g); \quad \Delta H = -26.4$ kcal

368 **(d)**

Graphite possesses lesser energy than diamond.

370 **(a)**

Under adiabatic conditions work is done on the cost of internal energy of system.

371 **(b)**

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

$\therefore \Delta G = -11700 - 298 \times (-105)$

$= +19590\ \text{J} = +19.59\ \text{kJ}$

Thus, reaction is non-spontaneous.

372 **(c)**

Internal energy, enthalpy and entropy are state functions but work and heat are path functions.

373 **(a)**

$q = \Delta U - W; -W$ is work done by the system

$\Delta U = 40 - 8 = 32\ \text{J} \quad (\because -W = 8)$

374 **(a)**

At constant volume, heat of reaction is ΔU.

375 **(c)**

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

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

$\therefore\ T_2 = 100$ K

376 **(b)**

Heat of neutralization is defined as the energy released during neutralization of 1 eq. of an acid by 1 eq. of base.

377 **(d)**

$S_R + O_2 \rightarrow SO_2; \quad \Delta H = -70.96 \ldots\ldots$ (i)

$S_M + O_2 \rightarrow SO_2; \quad \Delta H = -71.03 \ldots\ldots$ (ii)

By eq. (i) – (ii) ,

$S_R \rightarrow S_M$; $\Delta H = 0.07$ kcal or 70 cal

379 **(d)**

It does not violate the first law of thermodynamics but violates the II law of thermodynamics

380 **(c)**

If $\Delta H = +$ ve and $\Delta S = -$ ve then the reaction is spontaneous

381 **(b)**

By eq. (ii) – (i),

$$C_G \longrightarrow C_D; \qquad \Delta H = +1.5 \text{ kJ}$$

382 **(a)**

This is derived formula.

383 **(a)**

$$W = 2.303\, nRT \log \frac{p_2}{p_1}$$

$$= 2.303 \times 1 \times 2 \times 300 \log \frac{10}{2} = 965.84$$

At constant temperature, $\Delta E = 0$

$\Delta E = q + W$;

$q = -W = -965.84$ cal

384 **(a)**

$H_2(g) + (1/2)O_2(g) \longrightarrow H_2O(g)$

$\Delta n = -1/2$ and thus, ΔS decreases or – ve

385 **(d)**

$\Delta G = \Delta H + T\left[\frac{\partial \Delta G}{\partial T}\right]_P$ (Gibbs-Helmholtz equation)

Also, $\Delta G = \Delta H - T\Delta S$; and $-\Delta G = nF \cdot E$

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

or $-T\Delta S = T\left[\frac{\partial \Delta G}{\partial T}\right]_P = T\left(\frac{-nF\partial E}{\partial T}\right)_P$

or $\left[\frac{\partial E}{\partial T}\right]_P = \frac{\partial S}{nF}$; similarly derive for other values .

386 **(a)**

Internal energy of 1 mole of gas $= \frac{3}{2}RT$

388 **(c)**

$C_D \longrightarrow C_G$; $\Delta H = -453.5$ cal,

i.e., Energy of C_G is less and thus, more stable.

390 **(c)**

At equilibrium $\Delta G = 0$.

391 **(c)**

$\Delta S = 16 \text{ J mol}^{-1}\text{K}^{-1}$

$$T_{\text{b.p}} = \frac{\Delta H_{\text{vapour}}}{\Delta S_{\text{vapour}}} = \frac{6 \times 1000}{16}$$

$= 375$K

392 **(a)**

$$\Delta S = \frac{\Delta H}{T} = \frac{1.435 \times 10^3}{273} = 5.260 \text{ cal mole}^{-1}\text{K}^{-1}$$

393 **(b)**

Endothermic reactions are those in which heat energy is absorbed.

394 **(b)**

The melting of ice at −15°C is not an spontaneous process.

395 **(c)**

Mixing of gases increases the entropy

396 **(c)**

Current flows from higher potential to lower one.

397 **(a)**

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

$= 1 \times (15 - 3) = 12$ litre atm

$$= \frac{12 \times 1.987 \times 4.184}{0.0821}$$

$= 1.215 \times 10^3$J

398 **(c)**

ΔH for combustion of 56 litre $H_2 = \frac{-24.1 \times 56}{22.4}$

ΔH for combustion of 56 litre CO $= \frac{-263 \times 56}{22.4}$

$\therefore$ Total $\Delta H = -1312$ kJ

399 **(b)**

This is the derived formula for W_{rev} in adiabatic process.

400 **(a)**

$\frac{1}{2} H_2 + \frac{1}{2} Cl_2 \rightarrow HCl$; $\qquad \Delta H = -90$kJ

$$\therefore \qquad \Delta H = \frac{1}{2}\, e_{\text{H−H}} + \frac{1}{2}\, e_{\text{Cl−Cl}}$$

or $\quad -90 = \frac{1}{2} \times 430 + \frac{1}{2} \times 240 - e_{\text{H−Cl}}$

$\therefore \quad e_{\text{H−Cl}} = 425 \text{ kJ mol}^{-1}$

401 **(a)**

Find $\frac{1}{2}N_2 + \frac{3}{2}Cl_2 \longrightarrow NCl_3$;

Multiply Eqs. (ii) by 1/2, (iii) 3/2 and subtract from Eq.(i); we get

$$\Delta H_f = -\Delta H_1 - \left[-\frac{\Delta H_2}{2} + \frac{3}{2}\,\Delta H_3\right]$$

$$= -\Delta H_1 + \frac{\Delta H_2}{2} - \frac{3}{2}\,\Delta H_3.$$

402 **(b)**

Work done in 1 s = 400 J

Hence, work in 5 min (300 s)

= 400 × 300 = 120 kJ

$\Delta H^{\circ}_{\text{vap.}} = \frac{120 \text{ kJ}}{50/18} = 43.2$ kJ/mol

403 **(b)**

$W_{\text{irr.}}$ For a process at constant pressure $= -P \cdot \Delta V$; note that work is irreversible if expansion is made at constant pressure.

404 **(b)**

$$\Delta H = \frac{2.5\ \times 16}{4} = \ -10 \text{ kcl mol}^{-1}$$

405 **(c)**

$$W = \int_{1/1}^{1/2} pdV = -p(V_2 - V_1)$$

$W = -1(20 - 10) = -10 \text{ dm}^3 \text{ atm}$

$$= -10 \text{ dm}^3 \times \frac{8.314 \text{ JK}^{-1}\text{mol}^{-1}}{0.0821 \text{ dm}^3\text{K}^{-1}\text{mol}^{-1}} = -1013 \text{ J}$$

From, 1st law of thermodynamics

$\Delta U = q + W$

$= 800 \text{ J} + (-1013 \text{ J}) = -213 \text{ J}$

406 **(a)**

$$\Delta S = \frac{\Delta H_v}{T}$$

$\Delta H = 186.5$ kJ

$T = 373$ K

$$\therefore \Delta S = \frac{186.5}{373} = 0.5 \text{ kJ K}^{-1}\text{mol}^{-1}$$

407 **(c)**

$\Delta C_p = \ \Sigma C_p$ product $- \Sigma C_p$ reactant note C_p is for 1 mole.

408 **(d)**

$\left(\frac{\partial H}{\partial T}\right)_P = C_P; \left(\frac{\partial U}{\partial T}\right)_V = C_V.$

Also, $(\delta U/\delta V)T$ for ideal gas $= 0$.

409 **(b)**

$T = 300 \text{ K}, \Delta V = 10 - 1 = 9$ litre

$\because \Delta H = \Delta U + \Delta PV = \Delta U + 2 \times RT \quad (\because PV = nRT)$

$= 0 + 2 \times 8.314 \times 300 = 4.98$ kJ

$(\because \Delta E = 0$ for isothermal)

410 **(c)**

Average bond energy

$$= \frac{\text{Heat of dissociation of CH}_4}{4}$$

411 **(b)**

Follow definition of heat of formation.

412 **(a)**

$V_1 = 100$ mL

$V_2 = 250$ mL

Pressure $p = 2$ atm or $2 \times 1.01 \times 10^5 \text{ Nm}^{-2}$

Work done by the gas

$W = p\Delta V$ or $p(V_2 - V_1)$

Put the value in given formula

$W = 2 \times 1.01 \times 10^5 (0.250 \times 10^{-3} - 0.100 \times 10^{-3})$

$= 2 \times 1.01 \times 10^5 \times 0.15 \times 10^{-3}$

$= 30.30$ J

413 **(c)**

$\Delta U = 0$ for a cyclic process.

414 **(d)**

$W = P \times \Delta V$

$= 0 \times \Delta V = 0$

416 **(b)**

On heating $N_2O_4 = 2NO_2$, if NO_2 is formed more, it means $\Delta H = $ +ve because endothermic reactions are favoured with increase in temperature.

417 **(c)**

We know that internal energy of a gas depends upon its pressure and temperature. Thus, if a gas expands at constant temperature and pressure, than its internal energy remains same

418 **(a)**

$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g); \ \Delta H = \ -25$ kcal

Given, $e_{C-H} = 20 + e_{C-Cl} = 20 + a \ (e_{C-Cl} = a)$

and $e_{H-H} = e_{HCl} = b$

Now, ΔH reaction $= -[e_{C-Cl} + e_{H-Cl}] + \ [e_{C-H} + e_{Cl-Cl}]$

or $- 25 = \ -[a + b] + [20 + a + e_{Cl-Cl}]$

$\therefore \ e_{Cl-Cl} = \ -25 - 20 + b = \ -45 + b$

Now for, $H_2(g) + Cl_2(g) \rightarrow 2HCl(g); \ \ \Delta H_1 = ?$

$\therefore \ \Delta H_1 = \ -2\ [e_{H-Cl}] + [e_{H-H} + \ e_{Cl-Cl}]$

$= \ -2\ [b] + [b + (-45 + b)]$

$\Delta H1 = \ -45 \text{ kcal mol}^{-1}$

$\therefore \ \Delta H$ formation for HCl $= -22.5 \text{ kcal mol}^{-1}$

420 **(a)**

$\Delta E = q + W$

q=0

($\because$ Temperature is to be increase, no heat should enter or leave the system)

$\Delta E = q + W = 0 + W$ or $\Delta E = W$

$\therefore \quad W \neq 0, q = 0$

421 **(d)**

$\Delta G = $ +ve in each case.

422 **(c)**

Required equation is $\frac{1}{2}H_2 + \frac{1}{2}Cl_2 \rightarrow HCl$

$$\Delta H = \sum BE_{(\text{products})} - \sum BE_{(\text{reactants})}$$

$$= BE(HCl) - \left[\frac{1}{2}BE(H_2) + \frac{1}{2}BE(Cl_2)\right]$$

$$= -103 - \left[\frac{1}{2}(-104) + \frac{1}{2}(-58)\right]$$

$= -103 - (-52 - 29) = -22$ kcal

423 **(d)**

In a cyclic process, $\Delta E = 0$.

424 **(a)**

$\Delta S_{\text{vap}} = \frac{\Delta H_{\text{vap}}}{T_b} = \frac{840}{173} = 4.8$ J/mol/K

425 **(c)**

Energy absorbed $\propto \frac{1}{\text{stability of compound}}$

Energy released $\propto$ stability of compound

Thus, the order of stability is

$142.2 > 25.9 > -46.2 > -393.2$

ie, $O_3 > HI > NH_3 > CO_2$

426 **(a)**

$C\ (\text{graphite}) + \frac{1}{2}O_2(g) \rightarrow CO(g)$

$$\Delta n_g = 1 - \frac{1}{2} = \frac{1}{2}$$

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

$$\Delta E = \Delta E - \Delta n_g RT$$

$$= -26.4 - \frac{1}{2} \times 0.002 \times 298$$

$$= -26.7 \text{ kcal}$$

427 **(a)**

Bomb calorimeter is commonly used to find the heat of combustion of organic substance which consists of a sealed combustion chamber called a bomb, if a process is run in a sealed container then no expansion or compression is allowed, so $W = 0$ and $\Delta U = q$. Thus it has $\Delta U < 0, W = 0$

428 **(d)**

Heat of combustion is usually measured by placing a known mass of a compound in a steel container called a constant volume bomb calorimeter.

$$\Delta E = C \times \Delta t \times \frac{M}{m}$$

Where, C=heat capacity of calorimeter, $\Delta t = (t_2 - t_1)$ m=mass of substance taken and M=molar mass of substance

429 **(a)**

ΔH/mole for carbon is more and thus carbon can reduce ZnO to Zn.

430 **(d)**

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2; \quad \Delta H = -26.8 \text{ kJ}$$

$$2FeO + 2CO \rightarrow 2Fe + 2C; \Delta H = -33.0 \text{ kJ}$$

$$- \quad - \quad - \quad - \quad\quad +$$

$$\overline{Fe_2O_3 + CO \rightarrow 2FeO + CO_2;\ \Delta H = +6.2 \text{ kJ}}$$

431 **(c)**

An isobaric process is one in which changes are made at constant pressure.

432 **(b)**

$$\Delta S_{\text{vap}} = \frac{\Delta H_{\text{vap}}}{T} = \frac{9710}{373} = 26.032 \text{ cal K}^{-1}\text{mol}^{-1}$$

433 **(c)**

The property of the system whose value depends upon the amount of substance present in the system is called extensive property.

Gibb's free energy is an extensive property.

434 **(d)**

$$(CH_3)_2C = CH_2(g) + 6O_2(g) \rightarrow 4CO_2(g) + 4H_2O(l)$$

$\Delta ng = 4 - 7 = -3$(*ie*, negative)

We know that $\Delta H = \Delta E + \Delta n_g RT$

$= \Delta E - (\Delta m)RT \quad (\because\ \Delta n_g = -\text{ve})$

$\therefore\ \Delta H < \Delta E$

436 **(a)**

The process involves conversion of 1 mole of H_2O from liquid to vapour state and thus, heat changes are called heat of vaporisation.

437 **(c)**

This is definition.

438 **(b)**

Due to extensive solvation of F^- ion on account of smaller size, observed value of heat of neutralization of HF appears more.

439 **(d)**

$$\Delta H_{\text{reaction}} = -[l_{C-C} + 6 \times l_{C-H}] + [l_{C=C} + 4 \times l_{C-H} + l_{H-H}]$$

$$= [336.49 + 6 \times 410.50] + [606.10 + 4 \times 410.50 + 431.37]$$

$$= 120.02 \text{ kJ mol}^{-1}$$

440 **(a)**

This is significance of Hess's law, e.g., Heat of formation cannot be determined experimentally for $6C + 3H_2 \longrightarrow C_6H_6$; $\Delta H =$? Because the reaction does not occur. However, if heat of combustion for C, H_2 and C_6H_6 are determined experimentally as A, B, C, respectively, then we can calculate, $\Delta H = 6A + 3B - C$.

442 **(a)**

$\Delta S_{\text{total}} = \Delta S_{\text{sys.}} + \Delta S_{\text{surr.}} = +\text{ve}.$

443 **(a)**

On mixing gases entropy increases due to increase in disorderness.

444 **(b)**

$1\ M\ H_2SO_4 = 2$ eq. H_2SO_4

$1\ M$ HCl $= 1$ eq. HCl

Thus, for equal volume of two acids to be neutralized separately with NaOH, heat evolved will be twice in case of H_2SO_4 to that of HCl.

446 **(d)**

It is Cl_2 molecule which dissociates to give free radicals on exposure to light.

447 **(c)**

Bomb calorimeter measures q_v which is equal to ΔE.

448 **(b)**

If work done by the system is positive, then $q = \Delta U + W$. However, new terminology has revealed that work done by the system is negative and work done on the system is positive. Thus, according to this, $q = \Delta E - W$.

449 **(d)**

B. E. of $S = O$, $C \equiv C$, $C \equiv N$ and $N \equiv N$ are 523, 839, 891 and 941 kJ mol^{-}1 respectively.

450 **(c)**

Internal energy of a gas depends upon its pressure and temperature. Thus, if a gas expands at constant temperature and pressure, then its internal energy remains same.

451 **(c)**

ΔH and ΔS both are +ve for spontaneous change and $\Delta H = +\text{ve}$ for endothermic reaction

452 **(c)**

It provides information about physical states of reactants and products as well as about thermal changes. (d) is wrong because combustion is exothermic.

453 **(c)**

$$C_{(s)} + \frac{1}{2}O_{2(g)} \rightarrow CO_{(g)}; \Delta n + \frac{1}{2};$$

Also the moles of gases increase and therefore entropy change (ΔS) is positive. An increase in temperature will cause more change in $T\Delta S$. Also it is a combustion reaction and thus $\Delta H = -\text{ve}$

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

$= -\text{ve} - (+\text{ve}) = -\text{ve}$

454 **(a)**

$\Delta H = 18500 = \Delta U + \Delta nRT$

or $\quad 18500 = \Delta U + (-1) \times 2 \times 298$

or $\quad \Delta U = 19096$ cal

455 **(a)**

Work obtained is isothermal

$$W = -2.303\, nRT \log \frac{p_1}{p_2}$$

Given, $p_1 = 500$ kPa, $p_2 = 200$ kPa

$$(W) = -2.303\, nRT = \log \frac{500}{200}$$

$$= -2.303\, nRT \times 0.3980\, nRT$$

$\therefore$ Minimum work will be obtained when number of moles (n) is minimum.

Moles of substances, mass=1 kg=1000 g

2. Moles of $Cl_2 = 1000/71$
3. Moles of $O_2 = 1000/32$
4. Moles of $N_2 = 1000/28$
5. Moles of $CH_4 = 1000/16$

$\because$ Moles of Cl_2 are least.

Minimum work is obtained when 1 kg of chlorine gas expands.

456 **(d)**

In adiabatic expansion work is done by the system so it will lose heat to the surrounding. Thus, temperature of the system decreases.

457 **(d)**

$$\Delta H^{\circ} = -2 \times e_{H-Br} + e_{H-H} + e_{Br-Br}$$
$$= -2 \times 364 + 433 + 192$$
$$= -103 \text{ kJ}$$

458 **(c)**

Heat of neutralisation will be less than −57.33 kJ/mole because some amount of this energy will be required for the dissociation of weak base (MgO)

459 **(d)**

Only the laws of thermodynamics are mere statements, so they, are believed. Many mathematical expression are derived on the basis of these laws.

462 **(c)**

Given $\Delta H = +\text{ve}; \Delta S = +\text{ve}$

Thus, ΔG is −ve only when $\Delta H < T\Delta S$.

463 **(a)**

$$\Delta S = \frac{\Delta H}{T} = \frac{30 \times 10^3}{300} = 100 \text{ J mol}^{-1}\text{K}^{-1}$$

464 **(b)**

$\Delta H = \Delta U + \Delta nRT \quad \because \Delta n = 3 - 5 - 1 = -3$

$\therefore \quad \Delta H - \Delta U = -3RT$

465 **(c)**

Entropy (a measure of disorder) of universe is increasing towards maximum. This is second law of thermodynamics.

466 **(c)**

Move +ve is ΔH_s more is heat of solution.