Thermodynamics

SOLUTIONS

Exercise-I (Conceptual Questions)

1. Ans. (2)
It is the branch of science which deals with the energy changes taking place in all physical and chemical process.

2. Ans. (3)
Thermo flask is made up of insulating material, so there is no exchange of heat, so it is an example of isolated system.

3. Ans. (4)
Intensive quantities are those which is independent from the mass of the matter.
Extensive are those which depend on the mass of the matter.
Enthalpy, Volume are Extensive quantity
Temperature, refractive index are intensive.

4. Ans. (4)
Mass, Energy, Enthalpy all are extensive property.

5. Ans. (3)
In adiabatic process, q = 0
We know from 1st Law-
$\Delta U = q + W$
$\Delta U = W$
$q = 0$

6. Ans. (3)
In adiabatic process, work is independent of path
$\Delta U = q + W$ (q=0)
$\Delta U = W$
Internal energy is state function, depends only upon initial and final state.

7. Ans. (1)
In adiabatic process, q = 0
From 1st Law
$\Delta U = q + W$
$\Delta U = W$
We know $\rightarrow \Delta U = nC_V\Delta T$
$W = nC_V\Delta T$
When gas is compressed, work is positive, which means, temperature will increase.
So final temperature will be higher than initial temperature.

8. Ans. (4)
Heat supplied at constant pressure.
$q_p = \Delta H$ (Enthalpy change)
Heat supplied at constant volume
$q_v = \Delta E$ (Internal energy change)
ΔH, ΔE, Enthalpy all are state function, depends only upon initial and final state.

9. Ans. (1)
The quantities which doesn't depend on mass of the matter in intensive,
boiling point, EMF of cell, pH doesn't depend upon mass of matter,

10. Ans. (1)
We Know in reversible process, where P_{ext} is variable
$W = -\int P.dV$

11. Ans. (1)
In Isothermal expansion, $\Delta U = 0$
From 1st Law-
$\Delta U = q + W$
$q = -W$
$W = -q$
$W = -(-100\ Cal)$
$W = 100\ Cal$ ($\because$ 1Cal = 4.184J)
$W = 100 \times 4.184J$
$W = 418.4\ J$

12. Ans. (4)
Temperature is intensive property (doesn't depend on mass of matter)
Volume is extensive property/depends upon mass off matter)

13. Ans. (2)
From 1st Law-
$\Delta U = q + W$
For isothermal process, $\Delta U = 0$, $q = -W$
For adiabatic process, $q = 0$, $\Delta U = W$

14. Ans. (1)

From theory part we have study that in adiabatic expansion, temperature decreases, and in adiabatic compression, temperature increases.

→ In isothermal, temperature is constant throughout the process.

15. Ans. (4)

In reversible process,

→ takes in multisteps.

→ Driving force is infinitesimally greater than apposing force.

→ Work obtained is maximum.

16. Ans. (4)

Both q & w are path function.

Path function → Which depends upon path as well as initial and final state of the system.

17. Ans. (4)

Work done by system (W) = –300 J

Heat is supplied (q) = 100 Cal ($\therefore$ 1 Cal = 4.2J)

= 420 J

From 1st Law-

$\Delta U = q + W$

$\Delta U = 420 - 300$

$\Delta U = 120$ J

18. Ans. (1)

Internal Energy E_1,

q = –450J

W = 600J

From 1st Law

$\Delta E = -450$J

From 1st Law-

$\Delta E = q + W$

$(E_2 - E_1) = -450 + 600$

$E_2 = 150 + E_1$

19. Ans. (1)

Work done by system W = –8J

Heat is supplied q = 40J

From 1st Law-

$\Delta U = q + W$

$\Delta U = 40 - 8$

$\Delta U = 32$J

20. Ans. (3)

Internal Energy change = ?

In reversible isothermal expansion

In isothermal, temperature remains constant so internal energy change is zero

21. Ans. (1)

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

At constant temperature $\Delta E = 0$

At constant pressure, $W = P\Delta V$

22. Ans. (3)

Heat of reaction at constant pressure $(q_p) = \Delta H$

Heat of reaction at constant volume $(q_v) = \Delta E/\Delta U$

$\Delta H = \Delta U + \Delta ngRT$

$\Delta H - \Delta U = (-3)RT$ $(\Delta ng = 12 - 15 = -3)$

$\Delta H - \Delta U = (-3)(8.314)(298K)$

$= -7.43$ KJ

23. Ans. (2)

$A_{(g)} + 3B_{(g)} \longrightarrow 3C_{(g)} + 3D_{(g)}$

$\Delta E = 17$ kCal, T = 300K (273+27=300K)

$R = 2$Cal K^{-1} mol^{-1}

We know-

$\Delta H = \Delta E + \Delta ngRT$

$\Delta H = 17 + ((2) \times 2 \times 300) \times 10^{-3}$

$= 17 + 1.2$

$\Delta H = 18.2$ kCal

24. Ans. (2)

In reaction

$\Delta ng = 2 - (3) = -1$

$\Delta H = \Delta E + \Delta ngRT$

$\Delta H = \Delta E + (-1)RT$

$\Delta H = \Delta E - RT$

$\Delta H < \Delta E$

25. Ans. (4)

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

$\Delta ng = \frac{1}{2}$

$\Delta H = \Delta E + \Delta ngRT$

$\Delta H = \Delta E + (0.5)RT$

$\Delta H > \Delta E$

26. Ans. (2)

$2CO_{(g)} + 1O_{2(g)} \longrightarrow 2CO_2(g)$

$\Delta ng = 2-3 = -1$

$\Delta H = \Delta E + \Delta ngRT$

$\Delta H = \Delta E - RT$

$\Delta H < \Delta E$

27. Ans. (2)

Combustion of isobutane

$$C_4H_{10} + \frac{13}{2}O_2 \longrightarrow 4CO_2 + 5H_2O$$

$$\Delta ng = \frac{3}{2} = 1.5$$

$\Delta H = \Delta E + \Delta ngRT$

$\Delta H = \Delta E + 1.5RT$

$\Delta H > \Delta E$

28. Ans. (4)

$\Delta H = nC_p\Delta T$ Isothermal, $T \rightarrow$ constant

$\Delta H = (1)C_p(0)$ $\Delta T = 0$

$\Delta H = 0$

29. Ans. (2)

$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

$T = 977°C$

$T = 977+273$

$= 1250K$

$\Delta H = 174$ KJ/mol

$\Delta H = \Delta E + \Delta ngRT$

$174 \text{ KJ/mol} = \Delta E + [(1) \times 8.314 \times 1250] \times 10^{-3}$

$\Delta E = 174 - 10.39$

$\Delta E = 163.6$ KJ

30. Ans. (1)

Heat of reaction at constant $P(q_p) = \Delta H = ?$

Heat of reaction at constant

$V(q_v) = \Delta E = 67.71$kCal

We know-

$\Delta H = \Delta E + \Delta ngRT$

$\Delta H = (67.71 \times 10^3) + (1-1.5)RT$

$(T = 273 + 17 = 290 K)$

$(R = 2)$

$\Delta H = (-67.7) + ((-0.5) \times 290) \times (2)$

$= -67.7 - 0.29$

$\Delta H = -68$ Kcal

31. Ans. (1)

Enthalpy of vaporisation (ΔH_{vap}) and internal energy ΔE then-

$\Delta H = \Delta E + \Delta ngRT$

For vaporisation of water

$H_2O(l) \longrightarrow H_2O(g)$

$\Delta ng = 1-0$

$= 1$

$\Delta H_{vap} = \Delta E + \Delta ngRT$ $T = 100°C$

$= 273 + 100$

$T = 373$ K

$40.63 = \Delta E + RT$

$\Delta E = 40.63 - (8.314 \times 373) \times 10^{-3}$

$\Rightarrow 40.63 - 3.10$

$\Delta E = 37.53$ KJ mol^{-1}

32. Ans. (1)

$S(s) + O_2(g) \longrightarrow SO_2(g)$

$\Delta H = \Delta E + \Delta ngRT$

$\Delta ng = 1-1 = 0$

$\Delta H = \Delta E$

33. Ans. (2)

$$Co(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)$$

$\Delta H = \Delta E + \Delta ngRT$

$$\Delta ng = 1 - \left(1 + \frac{1}{2}\right)$$

$$= \frac{-1}{2}$$

$$\Delta H = \Delta E - \frac{RT}{2}$$

$\Delta H < \Delta E$

34. Ans. (3)

Combustion of sucrose

$C_{12}H_{22}O_{11(s)} + 12O_{2(g)} \longrightarrow 12CO_{2(g)} + 11H_2O_{(l)}$

$\Delta ng = 12-12=0$

$\Delta H = \Delta E$

35. Ans. (4)

$\Delta H \neq \Delta E$

$\Delta ng \neq 0$

(i) $\Delta ng = 0$

(ii) $\Delta ng = 0$

(iii) $\Delta ng = 0$

(iv) $\Delta ng = 2 - 4 = -2$

36. Ans. (2)

Combustion of ethanol

$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$

$\Delta ng = -1$

$\Delta H = \Delta E + \Delta ngRT$ $\quad T = 27°C$

$T = 273+27=300K$

$\Delta H = -670.48 + (-1) \times 8.314 \times 300$

$\Delta H = -671.08$ Kcal

37. Ans. (2)

$\Delta H - \Delta E = \Delta ngRT$

$\underset{(g)}{CH_4} + 2\underset{(g)}{O_2} \longrightarrow \underset{(g)}{CO_2} + 2\underset{(l)}{H_2O}$

$\Delta ng = -2$

$\Delta H - \Delta E = (-2)\ (2) \times 298$ Cal

38. Ans. (2)

$\Delta H = \Delta E + \Delta ngRT$

When $\Delta ng < 0$, then $\Delta H < \Delta E$

$\Delta ng = 2-3 = -1$

39. Ans. (3)

$2x_{(s)} + 2y_{(s)} \longrightarrow 2C_{(l)} + D_{(g)}$

$q_p = -28$kCal mol^{-1} $\quad q_v = ?$

$q_p = q_v + \Delta ngRT$

$-28 = q_v + (1) \times 2 \times 300$

$q_v = -28-0.6 = -28.6$ kCal mol^{-1}

40. Ans. (1)

For reversible expansion

n = 1 mol

$V_1 = 10L$, $\quad V_2 = 20L$

For isothermal reversible expansion

$1\ J = 10^7$erg

$$W = -2.303\ nRT \log_{10}\left(\frac{V_2}{V_1}\right)$$

$$W = -2.303 \times 1 \times 8.314 \times 298 \times 10^7 \log \frac{2}{1}$$

41. Ans. (1)

n = 2 mol

Vaccum, $\quad$ (in vaccum work done is zero)

W = 0

42. Ans. (1)

n = 1 mol , V = 3 dm^3 , constant P = 1 atm

1 dm^3 = 1L

$V_1 = 3L$, $\quad V_2 = 13$ L

Work in irreversible expansion-

$W = -P_{ext}(\Delta V)$

$= -1(13-3) = -10$L atm

43. Ans. (2)

Gaseous moles are increasing, so entropy increases.

(ii) gaseous moles are increasing

(i), (iii), (iv) gaseous moles are same.

44. Ans. (3)

In reversible adiabatic process, entropy change of system as well as surrounding is zero.

$\Delta S_T = 0$

45. Ans. (4)

Entropy means-

disorderness, randomness

46. Ans. (3)

Solid is converted into gas, so Δs will be positive distance between molecules is increasing.

47. Ans. (4)

(i) $Br_2(l) \longrightarrow Br_2(g)$

Liquid to gas, entropy increases.

(ii) $C(s) + H_2O(g) \longrightarrow Co(g) + H_2(g)$

Moles of gas molecules increases, so entropy increase

(iii) $N_2(g)$ (10 atm) $\longrightarrow N_2$(g, 1atm)

Pressure is decreasing, so entropy increases.

(iv) Fe(at 400K) $\longrightarrow$ Fe (300 K)

Temperature is decreasing, so entropy will decrease.

48. Ans. (3)

(i) Liquid to solid, ΔS = –ve

(ii) 3 mole gas to 2 mole gas, ΔS = –ve

(iii) Liquid to gas, ΔS = +ve

(iv) 4 mole gas to 2 mol gas, ΔS = –ve

49. Ans. (1)

When egg is boiled, denaturation of protein occurs so entropy increases.

50. Ans. (1)

$H_2(g) + Cl_2(g) \longrightarrow 2HCl$

$$\Delta S^0_{reaction} = \sum S^0_{Product} - \sum S^0_{reactant}$$

$$\Delta S^0 = \left[2 \times S^0_{HCl}\right] - \left[S^0_{H_2} + S^0_{Cl_2}\right]$$

$$\Delta S^0 = [2 \times 0.19] - [0.13 + 0.22]$$

$\Delta S^0_{reaction} = 0.03$ KJ K^{-1}mol^{-1}

$= 30$ J K^{-1} mol^{-1}

51. Ans. (2)

Graphite has more entropy than diamond, because diamond has highly ordered structure, so it has less entropy.

52. Ans. (3)

When two gases are mixed, gaseous molecules increases, so randomness will increase, entropy increases.

53. Ans. (3)

ΔH_{vap} = ?

B.pt. (T) = 79.5°C $\Delta S = 109.8 JK^{-1}mol^{-1}$

= 273 + 79.5

T = 352.5K

We know-

$$\Delta S_{vap} = \frac{\Delta H_{vap}}{T}$$

$109.8 \times 352.5 = \Delta H_{vap}$

$\Delta H_{vap} = 38.70$ KJ mol^{-1}

54. Ans. (1)

$\Delta H = 900 J/g$, T = 100°C

T = 373 K

For 1 g water heat = 900J

For 18 g/1mol water heat = 900×18 J

= 16200 J

We know

$$\Delta S = \frac{\Delta H}{T} = \frac{16200J}{373K} = 43.4 JK^{-1}mol^{-1}$$

55. Ans. (3)

n = 5 mol, $V_1 = 8dm^3$ T = 27°C

$V_2 = 80\ dm^3$ T = 300 K

ΔS for isothermal-

$$\Delta S = nR\ln\left(\frac{V_2}{V_1}\right)$$

$$\Delta S = 5 \times 8.314 \times \ln\left(\frac{80dm^3}{8dm^3}\right)$$

$\Delta S = 2.303 \times 5 \times 8.314 \times 1$

$\Delta S = 95.73\ JK^{-1}$

56. Ans. (2)

For spontaneous irreversible process

$\Delta S_T > 0$, So entropy of system & surrounding increases.

(refer through notes)

57. Ans. (2)

According to 2nd Law of thermodynamics, during an natural or spontaneous thermodynamic process, total entropy or entropy of universe increases.

58. Ans. (1)

$H_2(g) + Br_2(g) \rightarrow 2HBr(g)$ $\Delta S^0 = 20.1 JK^{-1}$

$S^0_{H_2} = 130.6\ J\ mol^{-1}K^{-1}$, $S^0_{HBr} = 198.5\ J\ mol^{-1}K^{-1}$

$$\Delta S^0 = \left[2S^0_{HBr}\right] - \left[S^0_{H_2} + S^0_{Br_2}\right]$$

$20.1 = [2 \times 198.5] - [130.6 + S^0_{Br_2}]$

$S^0_{Br_2} = 246.3\ JK^{-1}\ mol^{-1}$

59. Ans. (4)

(i) Solid → Liquid (entropy increases)
ΔS = +ve

(ii) Expansion of gas (S↑ = ΔS = +ve)

(iii) Crystal dissolve (in dissolution S↑)

(iv) Polymerisation (In polymerisation, mole decreases, S↓ ΔS = –ve)

60. Ans. (2)

$$\Delta H^0_{formation} = 0$$

$\Delta S^0 \neq 0$

Free energy of formation for element in standard state = 0

61. Ans. (4)

Entropy of an adiabatic reversible process is constant,

Whereas $\Delta S = 0$

62. Ans. (3)

In adiabatic reversible process, $\Delta U \neq 0$

$\Delta T \neq 0$

but entropy remain constant, $\Delta S = 0$

63. Ans. (2)

$\Delta H = -11.7 \times 10^3\ J\ mol^{-1}$

$\Delta S = -105\ J\ mol^{-1}\ K^{-1}$ T = 25°C

= 25 + 273

= 298 K

We know-

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

$\Delta G = -11.7 \times 10^3 - (298)(-105)$

$-11700 + 31290$

$\Delta G = 19590\ J\ mol^{-1}K^{-1}$

$\Delta G > 0$

64. Ans. (2)

$\Delta H > 0, \quad \Delta S > 0$

For spontaneously $\quad \Delta G < 0$

We know-

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

$\downarrow \quad \downarrow$

+ve +ve

$T\Delta S > \Delta H$

65. Ans. (1)

For equilibrium ($\Delta G = 0$)

$\Delta H = T\Delta S$

$30.5 \text{ KJ mol}^{-1} = T \times 0.066 \text{ KJ mol}^{-1}$

$T = 462.12 \text{ K}$

66. Ans. (4)

$H_2O(l) \rightleftharpoons H_2O(g)$ at T = 100°C P = 1 atm

373K

Liquid to gas, $\Delta S = +ve, \Delta H \neq 0$

$\Delta H = \Delta U + \Delta ngRT$

$\Delta U = \Delta H - \Delta ngRT$

$\Delta U \neq 0$

But we know at equilibrium, $\Delta G = \Delta H - T\Delta S$

$\Delta G = 0$

$\Delta H = T\Delta S$

67. Ans. (3)

$\Delta H = 30.56 \text{ KJ mol}^{-1} \quad \Delta S = 66 \text{ JK}^{-1} \text{ mol}^{-1}$

$\Delta G = 0$

$\Delta H = T\Delta S$

$30.56 \times 10^3 = T \times 66$

$T = \frac{30.56 \times 10^3}{66} = 463K$

68. Ans. (2)

$\frac{1}{2}A_2(g) + \frac{3}{2}B_2(g) \rightarrow AB_3(g) \qquad \Delta H = -20KJ$

$S^0_{A_2} = 60, \quad S^0_{B_2} = 40, \quad S^0_{AB_3} = 50$

For equilibrium –

$\Delta G = 0$

$\Delta H^0 = T\Delta S^0$

$\Delta S^0 = \left[\sum S^0_{product}\right] - \left[\sum S^0_{reactant}\right]$

$\Delta S^0 = \left[S^0_{AB_3}\right] - \left[\frac{1}{2}S^0_{A_2} + \frac{3}{2}S^0_{B_2}\right]$

$\Delta S^0 = [50] - \left[\frac{60}{2} + 40 \times \frac{3}{2}\right]$

$\Delta S^0 = [50] - [90] = -40 \text{ JK}^{-1} \text{ mol}^{-1}$

$\Delta H = T\Delta S$

$-20 \times 10^3 = T \times (-40)$

$T = \frac{20000}{40} = 500 \text{ K}$

69. Ans. (3)

Precipitation of AgCl by Ag^+ ions and HCl

It is a spontaneous process, as soon as the Ag^+ ions added in HCl, precipitate formed.

70. Ans. (3)

Ice melting at 1 atm, 283K

Ice force at 273K or 0°C, at 283 K, it starts melting itself.

This process is spontaneous.

71. Ans. (4)

$A + B \longrightarrow C + D + q$

(Exothermic reaction)

$\Delta S = +ve, \quad \Delta H = -ve$

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

$\downarrow \quad \downarrow$

$-ve \quad +ve$

Which means at any temperature

$\Delta G < 0$

So reaction is spontaneous.

72. Ans. (1)

We know

$\Delta G = \Delta G^0 + RT\ln Q$

At equilibrium

$\Delta G = 0, \ Q = K_{eq}$

$\Delta G^0 = -RT\ln K_{eq}$

Standard free energy change ΔG^0 is related so K_{eq}.

73. Ans. (2)

Vant Hoff equation

$\Delta G^0 = -RT\log_e K_p$

$-\Delta G^0 = RT\ln K_p$

74. Ans. (2)

$\Delta G^0 > 0$

We know

$\Delta G^0 = -RT\ln K_p$

For $\Delta G^0 > 0$

$\ln K_p < 0$

$K_p < 1$

75. Ans. (2)

$K_p = 10$

$\Delta G^0 = -RT\ln K_p$ $\quad R = 8\ JK^{-1}\ mol^{-1}$

$\Delta G^0 = -8 \times 300 \times \ln(10)$ $\quad T = 300K$

$\Delta G^0 = -5.527\ KJ\ mol^{-1}$

76. Ans. (4)

Evaporation of liquid

$H_2O(l) \longrightarrow H_2O(g)$

(Endothermic reaction)

ΔS = +ve (liquid to gas conversion)

Free energy change (ΔG) decreases

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

ΔG decreases

77. Ans. (1)

$CO_2(s) \longrightarrow CO_2(g)$

Endothermic process (ΔH = +ve)

Entropy will increase due to solid to gas conversion.

78. Ans. (4)

All CO, CO_2, ice are exceptions of 3rd Law of thermodynamics. (As per theory)

79. Ans. (1)

$\Delta G = -26$ kCal, $\quad \Delta S = -60\ Cal\ K^{-1}$

$\Delta H = ?$ $\quad T = 27°C$

$= 300\ K$

We know-

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

$-26 \times 10^3 = \Delta H - (300)\ (-60)$

$\Delta H = -26000 - 18000$

$\Delta H = -44$ kCal

80. Ans. (4)

For spontaneously

$\Delta G < 0$

(i) $\Delta G = \Delta H - T\Delta S$

$\downarrow \quad \downarrow$

$-ve \quad +ve$

$\Delta G < 0$ (always spontaneous)

(ii) $\Delta G = \Delta H - T\Delta S$

$\downarrow \quad \downarrow$

$-ve \quad -ve$

Can be spontaneous when $\Delta H > T\Delta S$

(iii) $\Delta G = \Delta H - T\Delta S$

$\downarrow \quad \downarrow$

$+ve \quad +ve$

Spontaneous when $T\Delta S > \Delta H$

(iv) $\Delta G = \Delta H - T\Delta S$

$\downarrow \quad \downarrow$

$+ve \quad -ve$

$\Delta G > 0$ (always non spontaneous)

81. Ans. (1)

Formation of water

$H_2(g) + O_2(g) \longrightarrow H_2O(l)$

$\Delta H < 0$

$H_P - H_R < 0$

$H_P < H_R$

Chemical energy of H_2 & O_2 is more than that of water.

82. Ans. (1)

In Exothermic Reaction

$\Delta H < 0$

$H_P < H_R$

83. Ans. (2)

Thermochemical equation tells the physical state doesn't tell about the spontaneously.

84. Ans. (3)

$C(s) + O_2(g) \longrightarrow CO_2(g)$ $\quad \Delta H = -94$ Kcal

exothermic reaction $\quad$ It tells about phase.

85. Ans. (1)

High bond energy of nitrogen molecule.

(As per theory)

86. Ans. (3)

Diamond → Graphite

$\Delta H = -453.5$ Cal

Diamond releases heat to make graphite so energy of graphite is Low. So graphite is more stable than diamond.

87. Ans. (4)

Max. Energy, least stability

64.8 kCal

88. Ans. (1)

Explosive compound NCl_3 so, formation energy is +ve reaction is endothermic.

89. Ans. (2)

$C(s) + \frac{1}{2}O_2(g) \longrightarrow CO(g) \quad \Delta H = -26.4$ kCal

CO is exothermic compound, ΔH = –ve

90. Ans. (4)

(i) ΔH = +ve (Endo)

(ii) $\Delta H = \Delta E + \Delta ngRT$

↓ ↓

+ve +ve

ΔH = +ve (Endo)

(iii) 180.4 KJ energy is given (Endo)

(iv) ΔE = –ve

$\Delta ng = -1$

$\Delta H = \Delta E + \Delta ngRT$ ΔH =–ve

↓ ↓ (Exo)

–ve –ve

91. Ans. (2)

$3O_2 \rightarrow 2O_3 \quad \Delta H$ = +ve

Energy of O_3 > Energy of O_2

Stability of O_3 < Stability of O_2

92. Ans. (4)

P(White) → P(Red) ΔH = –ve

Energy of P(Red) < Energy of P(White)

Stability of P(Red) > Stability of P(White)

93. Ans. (2)

Kirchoff's equation tells about temperature. (As per theory)

94. Ans. (1)

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

$$O = \frac{\Delta H_2 - \Delta H_1}{100K}$$

$\Delta H_2 = \Delta H_1 = -3.57$ KJ

95. Ans. (2)

From 2nd reaction

HCl(*l*) → HCl(g) by giving some heat so some heat is used to convert HCl(l) into HCl(g) due to which heat evolved in reaction is less.

X < y

96. Ans. (4)

$(\Delta H_f)_{Compound}$ can be +ve or –ve

ΔH_f (element) = 0

97. Ans. (4)

$H_2(s) + I_2(g) \longrightarrow 2HI, \quad ...(1) \quad \Delta H = 12.4$ kCal

$\Delta H_f(HI) \Rightarrow$

$$\frac{1}{2}H_2 + \frac{1}{2}I_2 \longrightarrow HI \qquad ...(2)$$

So reaction (2) $= \frac{1}{2} \Rightarrow \frac{12.4}{2} = 6.2$kCal

98. Ans. (2)

For heat of formation, 1 mole compound is formed and reactants are taken as reference state.

99. Ans. (2)

$\Delta H_f(NH_3) = -46.0$ KJ/mol

$$\frac{1}{2}N_2 + \frac{3}{2}H_2 \longrightarrow NH_3 \qquad ...(1)$$

Given

$2NH_3 \rightarrow N_2 + 3H_2 \qquad ...(2)$

$\Rightarrow$ Eq. (2) = – Eq(1) × 2

$\Delta H = (-46) \times 2$

$= 92$ KJ mol^{-1}

100. Ans. (4)

$(\Delta H_g)\ CO_2 = -94$KJ $\quad \Delta H_f(CaO)) = -152$ KJ

$C + O_2 + CO_2 \qquad ... (1)$

$$Ca + \frac{1}{2}O_2 \longrightarrow CaO \qquad ... (2)$$

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g) \Rightarrow \Delta H = 42$ KJ ...(3)

$$\Delta H_f(CaCO_3) \rightarrow Ca + C + \frac{3}{2}O_2 \rightarrow CaCO_3$$

(1) + (2) – (3)

$\Delta H = -94 - 152 = 42$

$\Delta H \Rightarrow -288$ KJ

101. Ans. (4)

$\Delta H_f^0\ (CH_4) = -17.9$

$\Delta H_g^0\ (C_2H_4) = 12.5$

$\Delta H_g^0\ (C_3H_8) = -24.8$

ΔH for $CH_4 + C_2H_4 \rightarrow C_3H_8$ is

$$\Delta H = \Sigma(\Delta H_f)_{C_3H_8} - \Sigma(\Delta H_f)_{reactant}$$

$= (-24.8) - (-17.9 + 12.5)$

$\Delta H = -19.4$ kCal

102. Ans. (1)

$\Delta H_f(C_2H_6) = -21.1$

$\Delta H_f(CO_2) = -94.1$

$\Delta H_f(H_2O) = -68.3$ kCal

$\Delta H_C(C_2H_6)$

$C_2H_6 + \frac{7}{2}O_2 \longrightarrow 2CO_2 + 3H_2O$

$\Delta H_c(C_2H_6)$

$= \left[2\Delta H_f(Co_2) + 3 \times \Delta H_f(H_2O)\right] - \left[\Delta H_f(C_2H_6)\right]$

$= [2 \times (-94.1) + 3 \times (-68.3)] - [-21.1]$

$= -372$ K Cal

103. Ans. (2)

$\Delta H_f^0 = (CO_2) = -393.5$

$\Delta H_f^0 (CO) = -110.5$

$\Delta H_f^0 (H_2O) = -241.8$

$C + O_2 \rightarrow CO_2$ -393.5 ... (1)

$C + \frac{1}{2}O_2 \rightarrow CO$ -110.5 ... (2)

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$ -241.8 ... (3)

$CO_2 + H_2 \rightarrow CO + H_2O$ ...(4) $\Delta H = ?$

Eq. (4) = (2) + (3)–(1)

= 41.2 KJ/mol

104. Ans. (2)

$\Delta H_C(C) = -393.5$

$C + O_2 \rightarrow CO_2$

$\Delta H_C = -393.5$ ... (1)

$\Delta H_C(CO) = -283$

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

$\Delta H_C = -283$ ... (2)

$\Delta H_f(CO) \Rightarrow C + \frac{1}{2}O_2 \rightarrow CO$ ... (3)

Eq. (3) = – Eq(2) + Eq(1)

+283 –393.5

= –110.5 KJ

105. Ans. (3)

$\Delta H_C(C) = -393.3$ $\Delta H_C(S) = -293.72$

$\Delta H_C(CS_2) = -1108.76$

$\Delta H_f(CS_2)$

$C + 2S \longrightarrow CS_2$

$\Delta H_f(CS_2) = (\Delta H_C)_{reactant} - (\Delta H_C)_{product}$

$= [(-393.3) + (-293.72) - (-1108.76)]$

$= -980.74 + 1108.76$

$= 128.02$ KJ mol^{-1}

106. Ans. (2)

$C + 2H_2 \rightarrow CH_4$

$\Delta H = (\Delta H_C)_{reactant} - (\Delta H_C)_{product}$

$= [(-94) + (2) \times (-68.4)] - [-212.4]$

$= -230.8 + 212.4$

$\Delta H = -18.4$ kCal.

107. Ans. (3)

For reference state of element, standard heat of formation is zero.

$\Delta H_f(C_{graphite}) = 0$

108. Ans. (4)

$\Delta H_f(NO_2) = 8.0$, $\Delta H_f(N_2O_4) = 2$ kCal mol^{-1}

Dimerization of NO_2

$2NO_2 \longrightarrow N_2O_4$

$\Delta H = \Delta H_f(N_2O_4) - 2 \times \Delta H_f(NO_2)$

$= 2–16$

$= -14.0$ kCal mol^{-1}

109. Ans. (4)

Given-

$\frac{1}{2}M_2O \longrightarrow M + \frac{1}{4}O_2$ ...(1) $\Delta H = 120$ kCal

$4M + O_2 \rightarrow 2M_2O$ $(\Delta H = (1) \times 4)$

$\Delta H = -480$ kCal

110. Ans. (4)

7.8 g

$n = \frac{7.8}{78} = 0.1$ mol

For 1 mol, $\Delta H = -3264.4$ KJ mol^{-1}

0.1 mol, $\Delta H = -3264.4 \times 0.1$

$= -326.4$ KJ mol^{-1}

111. Ans. (1)

$\text{Calorific value} = \frac{\text{Heat of combustion}}{\text{Molar mass}}$

$CH_4 = \frac{212.8}{16} = 13.3 \leftarrow$ best fuel

$C_2H_6 = \frac{373}{30} = 12.43$

$C_2H_4 = \frac{337}{28} = 12.03$

$C_2H_2 = \frac{310.5}{26} = 11.94$

112. Ans. (3)

$\Delta H_f(CO) = -110$ KJ mol^{-1}

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

$\Delta H_f(CO_2) = -394$ KJ mol^{-1}

$C + O_2 \rightarrow CO_2$

$\Delta H_C(C) = \Delta H_f(CO_2) = -394$ KJ mol^{-1}

113. Ans. (2)

$\Delta H_f(C_2H_4) = 52$,

$2C + 2H_2 \rightarrow C_2H_4$ ...(1)

$\Delta H_f(CO_2) = -394$,

$C + O_2 \rightarrow CO_2$ ...(2)

$\Delta H_f(H_2O) = -286$

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$ ...(3)

$\Delta H_C(C_2H_4) = C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$

$\Delta H_C(C_2H_4) = \Delta H_f(\text{product}) - \Delta H_f(\text{reactant})$

$= [2\times(-394)+2(-286)]-[52]$

$= (-788-572)-(52)$

$= -1412$ KJ mol^{-1}

114. Ans. (1)

$C_6H_6 + \frac{1}{2}SO_2 \rightarrow 6CO_2 + 3H_2O$ $\Delta H = -3267$KJ .... (1)

$\Delta H_f(CO_2) = -393.5$

$C + O_2 \rightarrow CO_2$ ... (2)

$\Delta H_f(H_2O) = -285.83$

$H_2 + \frac{1}{2}O_2 + H_2O$ ... (3)

$\Delta H_f(C_6H_6)$

$6C + 3H_2 \rightarrow C_6H_6$ ... (4)

Eq. (4) = ((2)×6−(1)+(3)×3)

$\Delta H = (-2361) + (-857.49) - (-3267)$

$= 48.51$

115. Ans. (3)

112L

$n = \frac{112}{22.4} \Rightarrow 5$

Equal volume of H_2 & CO

56L H_2 & 56L CO

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$, $\Delta H = -241.8$ KJ

$\Delta H = -241.8 \times 2.5 = -604.5$ KJ

$CO + \frac{1}{2}O_2 \rightarrow CO_2$; ΔH ; -283 KJ

$\left(n = \frac{56}{22.4} = 2.5\text{mol}\right)$

n= 2.5 mol, $\Delta H = -283\times2.5 = -707.5$KJ

Total heat evolved = −707.5−604.5=1312 KJ

116. Ans. (1)

Energy required = 2870 kCal

ΔH_C(cane sugar) = −1349 kCal/mol

Daily consumption of sugar cane = $\frac{2870}{1349}$ = 2.12

So, weight of $C_{12}H_{22}O_{11} = 2.12 \times 342 = 728$ g

117. Ans. (3)

$C_2H_2 + H_2 \rightarrow C_2H_4$

ΔH_f^0 (H_2O) = −68.3

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$ $\Delta H_f = -68.3$

$\Delta H_C[H_2] = -68.3$

$\Delta H_C^0(C_2H_2) = -337.2$

$\Delta H_C^0(C_2H_4) = -363.7$

Heat change of reaction:-

(ΔH_C) reactant − (ΔH_C) product

= [−337.2 − 68.3] − [−363.7]

= [−41.8] kCal

118. Ans. (2)

Heat of combustion of a substance is always :-

Always negative.

119. Ans. (2)

$\Delta H_C(C) = -94.4$ kCal

$C(s) + O_2 \rightarrow CO_2$

$\Delta H_f(CO_2)$

$C + O_2 \rightarrow CO_2$

$\Delta H_f(CO_2) = \Delta H_C(C) = -94.4$ kCal

120. Ans. (2)

$\Delta H_C(CH_4)$ → For 1 mol combustion of CH_4

n of $CH_4 = \frac{0.4\text{ g}}{16} \Rightarrow \frac{1}{40}$mol

for $\frac{1}{40}$ mol, Heat liberated = 2.25 kCal

for 1 mol $\Rightarrow \frac{0.25}{1/40} = 0.25 \times 40 = -10$ kCal

121. Ans. (2)

–680 kCal heat evolved when $C_6H_{12}O_6$ = 1 mol

for 170 kCal heat evolved mole of $C_6H_{12}O_6 = \frac{1}{680}\times 170 = \frac{1}{4}$ mol

$\frac{1}{4}\left(C_6H_{12}O_6 + 9O_2 \rightarrow 6CO_2 + 6H_2O\right)$

CO_2 mol = $\frac{6}{4} = \frac{3}{2} = 1.5$ mol

Weight of $CO_2 = 1.5 \times 44 = 66$ g

122. Ans. (3)

For combustion, 1 mole reactant taken and $CO_2(g)$ & $H_2O(l)$ will be formed.

123. Ans. (1)

$\Delta H_f(CO_2)$ = –94 kCal

$C + O_2 \rightarrow CO_2$ ΔH = –94 kCal

$C + O_2 \rightarrow CO_2$

mass = 3g

$n = \frac{3}{12} = \frac{1}{4}$ mol

for 1 mol C, ΔH = –94kCal

for $\frac{1}{4}$ mol C, $\Delta H = \frac{-94}{4}$ = –23.5

124. Ans. (3)

1 mol NH_4OH + 1 mol HCl

When acid and base both are strong heat liberate = 13.7 kCal

One is weak then heat is less than 13.7 kCal

125. Ans. (2)

Given $H^+ + OH^- = H_2O$ + 13.7 kCal

1 mol H_2SO_4 with strong base neutralisation

2mol H^+ + 2mol OH^- = –13.7 × 2

$\Rightarrow$ –27.4 Kcal

126. Ans. (2)

Valency factor of NaOH = 1

Valency factor of dibasic acid = 2

Thus, only 1 gm equivalent of dibasic acid is neutralised by 1gm equivalent of NaOH, for reaction

$H^+ + OH^- \rightarrow H_2O$, ΔH = –13.7 KJ/mol

Hence the heat of neutralisation of a strong dibasic acid in dilute solution by NaOH = –13.7 kCal/eq

127. Ans. (1)

When 5 ml strong acid reacting with 5 ml strong base heat evolved = ΔH

$\Delta H = ms\Delta T$

$\Delta T = \frac{\Delta H}{ms}$

When 10 ml of each are mixed

$\Delta T = \frac{2\Delta H}{2ms} = \frac{\Delta H}{ms}$

So temp rise will be same.

128. Ans. (2)

$HCl + NaOH \rightarrow NaCl + H_2O$ ... (1)

ΔH_{neut} = –55.9 KJ/mol

$HCN + NaOH \rightarrow NaCN + H_2O$ ... (2)

ΔH_{neut} = –12.1 KJ/mol

(2) – (1)

$HCN + NaCl + NaOH + H_2O \rightarrow NaCN + H_2O + HCl + NaOH$

ΔH = –12.1 + 55.9

ΔH = 43.8 KJ/mol

Energy of dissociation for 1 mole of HCN

129. Ans. (1)

$H^+ + OH^-$, heat exchange during reaction

= –13.7 kCal

130. Ans. (1)

$NaOH + HCl \rightarrow NaCl + H_2O$

Heat of neutralisation.

131. Ans. (1)

$\Delta H_C(C_2H_4)$, $\Delta H_C(C_2H_6)$, $\Delta H_C(H_2)$

$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$ ... (1)

$C_2H_6 + \frac{7}{2}O_2 \rightarrow 2CO_2 + 3H_2O$ ... (2)

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$ ... (3)

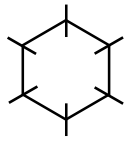
$C_2H_4 + H_2 \rightarrow C_2H_6$

(1) + (3) – (2)

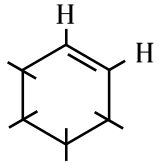
= (–1409.5) + (–285.6) – (–1558.3)

= –136.8 KJ

132. Ans. (1)



$C_6H_{12} + 9O_2 \rightarrow 6CO_2 + 6H_2O$ ... (1)



$C_6H_{10} + \frac{17}{2} O_2 \rightarrow 6CO_2 + 5H_2O$ ... (2)

$H_2 + \frac{1}{2} \rightarrow H_2O$ ... (3)

ΔH_{Hyd} (Cyclohexene)

$C_6 + H_{10} + H_{12} \rightarrow C_6H_{12}$ ... (4)

For eq. (4) $\Rightarrow$ (1) + (3) + (2)

= (+3920) + (–241) – 3800

= –121 KJ/mol

133. Ans. (2)

Bond energy of a molecule is +ve

134. Ans. (4)

Dissociate 1 mol gaseous bond into separate gaseous atoms.

$HCl(g) \rightarrow H(g) + Cl(g)$

135. Ans. (2)

Bond energy for dissociation ↑, Bond strongest.

136. Ans. (3)

For 1 mol bond energy of H_2 = ?

Gaseous $H(H_2)$

For 2 mol/energy = 208 kCal

Mole = $\frac{4}{2}$ = 2 mol

For 1 mol energy = 104 Kcal

137. Ans. (2)

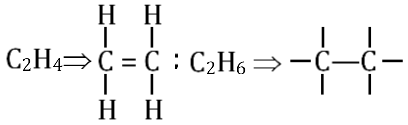
Bond energy of HCl = ?

$\Delta H_{reaction} = (BE)_R - (BE)_P$

182 = [430 + 242] – [2HCl]

HCl = 245 KJ/mol

138. Ans. (3)



$\Delta H_r = (BE)_R - (BE)_P$

$\Delta H_r = [103 + (145+4\times99)] - [80+6\times99]$

$\Delta H_r = -30$ kCal mol^{-1}

139. Ans. (2)

$\Delta H_{BDE}\ H_2 = 436$

$\Delta H_{BDE}\ N_2 = 941.8$

$N_2 + 3H_2 \rightarrow 2NH_3$

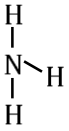
$\Delta H = 2\times\Delta H_f = 2\times(-46 KJ/mol)$

$\Rightarrow$ –92 KJ/mol

$\Delta H = \sum BE_R - \sum BE_P$

–92 = [941.8 + (3×436)] – [6x]

x = 390.3 KJ/mol



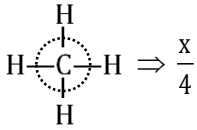
(Let BE of NH = x)

140. Ans. (3)

ΔH = +ve Endo

Average bond energy of C-H bond

$= \frac{\text{dissociation energy of } CH_4}{\text{number of bonds broken}}$



(C + 4H)

141. Ans. (4)

$H-OH \rightarrow H(g) + OH(g)$ $\Delta H = 498$

$OH(g) \rightarrow O(g) + H(g)$ $\Delta H = 428$

Two bond O–H broken so bond enthalpy of O–H bond

$= \frac{498+428}{2} = 463$ KJ mol^{-1}

142. Ans. (3)

ICl(g) + I(g) + Cl(g)

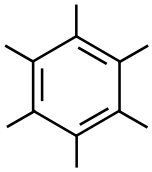
$\Delta H_{BDE}(ICl) = \sum \Delta H_f \text{product} - \sum \Delta H_f \text{ Reactant}$

= (121.34 + 106.96) – (17.57)

= 210.73 J mol^{-1}

143. Ans. (3)

Bond energy of formation of because



3C = C + 3C–C + 6C–H bond

$\Rightarrow$ 6(615) + 3(347.3) + 6(416.2)

$\Rightarrow$ 5384.1 KJ mol^{-1}

Resonance energy = $\Delta H_{dissociation} - \Delta H_f$

= 5535–5384.1

= 151 KJ mol^{-1}

144. Ans. (1)

Enthalpy of formation (Element in standard/reference state)

145. Ans. (1)

$Cl_2(g) \rightarrow 2Cl(g)$

For dissociation energy should be given generally positive.

146. Ans. (3)

Addition of solvent to solution.

Magnitude of react of solution depends upon the reaction (specific heat)

It doesn't depend on solute or solvent. So the specific enthalpy remain constant.

147. Ans. (2)

$H_2(g) \rightarrow 2H(g)$ $\Delta H = 104$ kCal

Atomisation of H_2 is H–H, $\Delta H = 104$ kCal is correct answer.

148. Ans. (4)

From $S_R \rightarrow S_M$

$S_R - S_M$ = +ve (Endothermic)

149. Ans. (4)

$\Delta H_4 - \Delta H_R$

$\Rightarrow$ –9.91 + 8.78

$\Rightarrow$ –1.13 KJ

150. Ans. (3)

Heat of reaction

151. Ans. (2)

For addition of I(g) in $I_2(g)$, Heat will be released so ΔH = –ve

152. Ans. (1)

$A(s) \xrightarrow{T} A(l)$; $\Delta H = x$ ... (1)

$A(l) \xrightarrow{T} A(g)$; $\Delta H = y$ ... (2)

$A(s) \longrightarrow A(g)$ ... (3)

Eq. (3) = Eq.(1) + Eq.(2)

$\Delta H = x + y$

153. Ans. (3)

Different intermediate reactions.

Enthalpy change depends upon nature, states of reactant products.

154. Ans. (2)

$C(g) + \frac{1}{2}O_2 \rightarrow CO$, $\Delta H = -110.5$ ...(1)

$CO + \frac{1}{2}O_2 \rightarrow CO_2$ $\Delta H = -283.2$ ...(2)

$C_{(graphite)} + O_2 \rightarrow CO_2$ ... (3)

Eq. (3) = (1) + (2)

= –110.5 – 283.2

ΔH = –393.7 KJ

155. Ans. (2)

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$, $\Delta H = -68.39$ ... (1)

$K + H_2O$ + water $\rightarrow$ KOH(aq.) + $\frac{1}{2}H_2$;

$\Delta H = -48$kCal ... (2)

KOH + water $\rightarrow$ KOH(aq)

$\Delta H = -14$ kCal ... (3)

$K + \frac{1}{2}O_2 + \frac{1}{2}H_2 \longrightarrow KOH$ ... (4)

Eq. (4) = (2) + (1) – (3)

= –68.39–48+14

156. Ans. (3)

$C + O_2 \rightarrow CO_2$ $\Delta H = -94.2$ ...(1)

$H_2 + \frac{1}{2}O_2 \rightarrow CO_2$ $\Delta H = -68.3$ ...(2)

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

$\Delta H = -210.8$ ...(3)

$C_{(s)} + 2H_2 \rightarrow CH_4$ ...(4)

Eq. (4) = (1) + (2×(2))–(3)

ΔH = –20 kCal

157. Ans. (3)

$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

$\Delta H^0 = -44kCal$ ...(1)

$2Na + 2HCl \rightarrow NaCl + H_2$

$\Delta H^0 = -152$ kCal ...(2)

$Na(s) + 0.5Cl_2(g) + NaCl(s)$ $\Delta H^0 = ?$ ... (3)

Eq. (3) = $\frac{Eq.(1)}{2} + \frac{Eq.(2)}{2} \Rightarrow \frac{-44}{2} - \frac{-152}{2}$

= –98 kCal

158. Ans. (2)

$S + \frac{3}{2}O_2 \rightarrow SO_3$ $\Delta H = -2x$ ... (1)

$SO_2 + \frac{1}{2}O_2 \rightarrow SO_3$ $\Delta H = -y$ ... (2)

$S + O_2 \rightarrow SO_2$ ... (3)

Eq.(3) = Eq.(1)–(2)

$\Rightarrow -2x + y$

$= -(2x-y)$

159. Ans. (1)

$S + O_2 \rightarrow SO_2$ $\Delta H = -298.2$ ... (1)

$SO_2 + \frac{1}{2}O_2 \rightarrow SO_3$ $\Delta H = -98.7$ ... (2)

$SO_3 + H_2O \rightarrow H_2SO_4$ $\Delta H = -130.2$ ... (3)

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$ $\Delta H = -287.3$ ... (4)

$H_2 + S + 2O_2 \rightarrow H_2SO_4$ ... (5)

Eq.(5) = (1) + (2) + (3) + (4)

$\Delta H = -298.2 - 98.7 - 130.2 - 287.3$

$\Delta H = -814.4$ KJ

160. Ans. (4)

$Zn + \frac{1}{2}O_2 \rightarrow ZnO$ $\Delta H = -84000$ Cal ...(1)

$Hg + \frac{1}{2}O_2 \rightarrow HgO$ $\Delta H = -21700$ Cal ...(2)

$Zn + HgO \rightarrow ZnO + Hg$ ... (3)

Eq(3) = Eq(1)–(2)

= –84000 + 21700 = –62300 Cal

161. Ans. (4)

$2C + 2O_2 \rightarrow 2CO_2$ $\Delta H = -787$ ... (1)

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$ $\Delta H = -286$ ... (2)

$C_2H_5 + \frac{5}{2}O_2 \rightarrow 2CO_2 + H_2O$

$\Delta H = -1310$KJ ... (3)

$2C + H_2 \rightarrow C_2H_2$ ... (4)

Eq.(4) = (1) + (2) – (3)

$\Delta H = +237$ KJ

162. Ans. (2)

$A + \frac{1}{2}O_2 \rightarrow AO$ $\Delta H = -50$ kCal ...(1)

$AO + \frac{1}{2}O_2 \rightarrow AO_2$ $\Delta H = 100$ kCal ...(2)

$A + O_2 \rightarrow AO_2$ ... (3)

Eq. (3) = Eq(1) + (2)

= –50+100

= 50 Kcal

163. Ans. (2)

$C + O_2 \rightarrow CO_2$ $\Delta H = -94$ kCal ... (1)

$CO + \frac{1}{2}CO_2$ $\Delta H = -67.7$ kCal ... (2)

$C + \frac{1}{2}O_2 \rightarrow CO$ ... (3)

Eq.(3) = Eq.(1)–(2)

$\Delta H = -94 + 67.7$

$\Delta H = -26.3$

26.3 kCal heat will be produced.

164. Ans. (1)

$H_2O(l) \rightarrow H_2O(g)$

$\Delta H = (-241.84) - (-285.77)$

$\Delta H = 43.93$ KJ mol^{-1}

165. Ans. (4)

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O(l)$ $\Delta H = -68.32$ kCal ... (1)

$H_2O(l) \rightarrow H_2O(g)$ $\Delta H = 10.52$ kCal ... (2)

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O(g)$ ... (3)

Eq.(3) = Eq.(1) + (2)

= –68.32 + 10.52

= –57.8 Kcal

166. Ans. (1)

$CuSO_{4(s)} \xrightarrow{H_2O} CuSO_{4(aq)}$

$\Delta H = -15.89$ ...(1)

$CuSO_4 \cdot 5H_2O_{(s)} \xrightarrow{H_2O} CuSO_{4(aq)}$

$\Delta H = 2.8$ Kcal ...(2)

$CuSO4_{(s)} \xrightarrow{H_2O} CuSO_4 \cdot 5H_2O_{(s)}$ ...(3)

Eq.(3) = Eq.(1) – (2)

$\Delta H = (-15.89) - (2.8)$

$\Delta H = -18.69$ K cal

167. Ans. (4)

$C + O_2 \rightarrow CO_2$ $\Delta H_{comb}(C)$ ... (1)

$CO + \frac{1}{2}O_2 \rightarrow CO_2$ $\Delta H_{comb}(CO)$ ... (2)

$\Delta H_g(CO)$

$C + \frac{1}{2}O_2 \rightarrow CO$ ... (3)

Eq.(3) = Eq.(1) – (2)

$\Delta H^0_g(CO, g)$

$= \Delta H^0_{comb}(C, \text{graphite}) - \Delta H^0_{comb}(CO, g)$

Exercise-II (Previous Year Questions)

1. Ans. (2)

$X_2O_{4(\ell)} \longrightarrow 2 \times O_{2(g)}$

$\Delta U = 2.1$ K cal, $\Delta S = 20$ cal K^{-1} $T = 300K$

$\Delta H = \Delta U + \Delta_{ng}RT$

$\Delta H = 2.1 + \frac{(2)\times(2)(300)}{1000}$

$\Delta H \Rightarrow 3.3$ K cal mol^{-1}

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

$\Delta G = 3.3 - \frac{(300)(20)}{1000}$

$\Delta G = 3.3 - 6$

$\Delta G \Rightarrow -2.7$ Kcal mol^{-1}

2. Ans. (2)

$\Delta G = \Delta G^0 + 2.303\ RT\ \log K$

At Equilibrium

$\Delta G = 0$

$\Delta G^0 = -2.303\ RT\ \log K$

3. Ans. (3)

$C + O_2 \rightarrow CO_2$ $\Delta H_{comb.} = -393.5$ KJ mol^{-1}

$C_{(graphite)} + O_2 \rightarrow CO_2$ $\Delta H_f = ?$

$\Delta H_{comb.} = \Delta H_f = -393.5$ KJ mol^{-1}

for 1 mol CO_2 formation = 44 g

for 44g CO_2, $\Delta H = -393.5$

35.2 g CO_2, $\Delta H = \frac{-393.5}{44} \times 35.2$

$\Delta H = -315$ KJ

4. Ans. (3)

For spontaneous reaction $\Delta G < 0$

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

For always $\Delta G < 0$

$\Delta H < T\Delta S$

$\Delta S > 0$, $\Delta H < 0$

5. Ans. (4)

$\Delta S = nc_v \ln \frac{T_1}{T_2} + nR \ln \frac{p_1}{p_2}$

For isothermal $T_1 = T_2$

$\Delta S = nR \ln \frac{p_1}{p_2}$

6. Ans. (1)

$\Delta H = 35.5$ KJ mol^{-1}, $\Delta S = 83.6$ J K^{-1} mol^{-1}

For spontaneity $\Delta G < 0$

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

$\Delta H - T\Delta S < 0$

$\Delta H < T\Delta S$

$T > \frac{\Delta H}{\Delta S} = \frac{35.5\times10^3}{83.6} = 424.64$

$T > 425$ K

7. Ans. (2)

$P_{ext} = 2.5$ atm, $V_i = 2.5$ L, $V_f = 4.5$ L

$\Delta U = ?$

From 1st law –

$\Delta U = q + w$ $(q = 0)$

$\Delta U = -P\Delta V = -2.5$ atm $(4.5 - 2.5)$ L

$\Rightarrow -5$ atm L 1 L. atm = 101.3J

$= -5 \times 101.3$ J

$\Delta U = -506.5 \simeq 505$ J

8. Ans. (3)

$\Delta H_f = -200$ KJ mol–1

ΔH_{BDE} $X_2 : Y_2 : XY$

$1 : 0.5 : 1$

Let $a : a/2 : a$

$$\frac{1}{2}X_2 + \frac{1}{2}Y_2 \longrightarrow XY$$

$\Delta H_{reaction} = \Sigma(\Delta H_{BE})_R - \Sigma(\Delta H_{BE})_P$

-200 KJ mol^{-1}

$$\Rightarrow \left[\frac{1}{2}\Delta H(x-x)+\frac{1}{2}\Delta H(y-y)\right]-\left[\Delta H_{xy}\right]$$

$$-200=\left[\frac{1}{2}(a)+\frac{1}{2}\frac{(a)}{2}\right]-[a] \Rightarrow a = 800$$

9. Ans. (4)

$T = 300K$, $V_1 = 0.1$ L, $V_2 = 0.25$ L, $P_{ext} = 2$ bar

$w = ?$

$w = -P\Delta V = -2bar\ (0.25 - 0.1) = -30J$

10. Ans. (2)

$V_1 = 10^{-3}m^3$, $V_2 = 10^{-2}m^3$, $T = 300$ K,

$P = 10^{-5}$ N/m^2

$w = -P\Delta V$

$\Rightarrow -10^5\ (10^{-2} - 10^{-3})$

$\Rightarrow -10^5 \times (0.009)$

$\Rightarrow -900$ J

11. Ans. (4)

Constant temp. throughout the process. So, $T_A = T_C$

12. Ans. (2)

For free expansion of ideal gas, adiabatic

$w = 0, \quad \Delta T = 0, \quad q = 0$

13. Ans. (2)

$Kc = 2 \times 10^{13}$, $\Delta G^0 = ?$

$\Delta G^0 = -RT\ln K$

$= -8.314 \times 300 \times \ln (2\times10^{13})$

14. Ans. (1)

$2Cl\ (g) \longrightarrow Cl_2\ (g)$

For association, ΔH exothermic

$\Delta H < 0, \Delta S < 0$

n is decreasing

15. Ans. (1)

$\Delta H = 30$ KJ mol^{-1} $\quad T = 450$ K, $\Delta S = ?$

For spontaneity $\Delta G < 0$

$\Delta H < T\Delta S$

$$T > \frac{\Delta H}{\Delta S}$$

$$\Delta S > \frac{\Delta H}{T} \Rightarrow \frac{36\times10^3}{450}$$

$\Delta S > 66.66$

$\Delta S = 70$

16. Ans. (1)

$\Delta H = -109$ KJ mol–1

$H_2 + Br_2 \rightarrow 2HBr \quad \Delta H_{BE}(H_2) = 435$ KJ mol^{-1}

$\Delta H_{reaction} = \Sigma(\Delta H_{BE})_R - \Sigma(\Delta H_{BE})_P$

$\Delta H_{BE}(Br_2) = 192$ KJ mol^{-1}

$-109 \Rightarrow [435 + 192] - [2 \times \Delta H_{HBr}]$

$\Delta H_{HBr} \Rightarrow 368$ KJmol^{-1}

17. Ans. (2)

For 1 mol of an ideal gas, $C_p - C_v = R$

(As per theory)

18. Ans. (3)

For isothermal, $\Delta U = 0$

$\Delta S_T \neq 0$

19. Ans. (1)

Work done = area under the curve

$\Rightarrow 1$

20. Ans. (4)

$\Delta U = q + w$

For isothermal

$\Delta U = 0$

21. Ans. (2)

$A + B \rightleftharpoons C + D$

$[A] = 2$ mol L^{-1}

$[B] = 3$ mol L^{-1}

$[C] = 10$ mol L^{-1}

$[D] = 6$ mol L^{-1}

$\Delta G^0 = -2.303\ RT \log K_{eq}$

$$= -2.303RT \log \frac{[C][D]}{[A][B]}$$

$$= -2.303 \times 2 \times 300 \times \log \frac{10\times6}{2\times3}$$

$= -2.303 \times 2 \times 300 \times \log 10$

$= -1381.8$ cal

22. Ans. (1)

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

23. Ans. (4)

(A) Isothermal process $\rightarrow$ Temperature is constant(II)

(B) Isochoric process $\rightarrow$ Volume is constant(III)

(C) Isobaric process $\rightarrow$ Pressure is constant(IV)

(D) Adiabatic process $\rightarrow$ No exchange of heat(I)

24. Ans. (3)

(A) When liquid evaporates to vapour, randomness increases hence entropy increases.

(B) Temperature decreases $\Rightarrow$ entropy decreases

(C) $\Delta n_g = 2 - 0 = 2$

Δn_g = + ve $\rightarrow \Delta S$ = + ve $\Rightarrow$ entropy increases

(D) $\Delta n_g = 2 - 1 = 1$

Δn_g = +ve $\rightarrow \Delta S$ = + ve $\Rightarrow$ entropy increases

(A), (C), (D) are correct.

25. Ans. (2)

In Reversible Isothermal process,

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

$$= -2.303 \times 1 \times 2 \times 298\ \log_{10}\left(\frac{20}{10}\right)$$

$= -413.14$ calories

26. Ans. (3)

Isothermal process $\Delta U = 0$

$\Delta U = q + w$ ___ (1)

$w = 0$ for vaccum

$\Delta U = q + 0 = 0$

27. Ans. (1)

For an endothermic reaction; q_p = (+) Ve

ΔH = (+) Ve

28. Ans. (1)

$A_2(g) + 3B_2(g) \rightarrow 2AB_3(g)\ \Delta H = +15$ kJ

$\Delta H = \Delta U + \Delta ngRT$

$$15 = \Delta U + (-2) \times \frac{25}{3} \times 10^{-3} \times 300$$

$\Delta U = (15 + 5)$ kJ $\simeq 19988.4$ J

Exercise-III (Analytical Questions)

1. Ans. (4)

A reversible process is a process in which all changes occurup at any part of the system are reversed when small always in variable are carried out in opposite direction.

Process takes place in infinite small steps and takes infinite time to complete process.

So option (4) is correct.

2. Ans. (4)

Combustion is an exothermic process.

The enthalpy of formation any standard stable state/reference state is zero.

(4) is correct

3. Ans. (3)

$C_{p_m} - C_{v_m} = R$ is applicable only for ideal gas and doesn't consider for non-ideal gas.

So (3) option correct

4. Ans. (4)

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

ΔH is calculated at constant pressure.

and constant temp. to decide spontaneity.

5. Ans. (2)

$(\Delta G)_{T,P} = 0$ is not correct for spontaneous reaction for spontaneity $\Delta G < 0$

6. Ans. (1)

(i) $\Delta S_{vap} = \frac{\Delta H_{vap}}{T}$

(ii) $\Delta U = q + w \Rightarrow \Delta U = w = nc_v\Delta T$

For expansion w = -ve

So temp. decreases

(iii) Crystalline solid states entropy is lowest.

(iv) Liquid freezes so intermolecular attraction increases, entropy decreases.

(1) Option correct

7. Ans. (2)

(a) liquid $\rightarrow$ Solid

Intermolecular attraction increases, entropy increases.

(b) T $\uparrow$ entropy $\uparrow$

(c) gaseous mole increases, entropy increases

(d) $H_{2(g)} \rightarrow 2H_{(g)}$

gaseous mole increases, entropy increases

(2) correct

8. Ans. (1)

$w = 0, q < 0$

$\Delta U = q + w$

$\Delta U = q$

$q = nc_v \Delta T$

q decreases, temp. decreases

option (1) correct

9. Ans. (4)

(A) $\Delta H_f^0(H_2) = \Delta H_f^0(Cl_2) = 0$

For reference state formation enthalpy is zero.

(R) Reason is correct

(4) option correct

10. Ans. (4)

As per theory all natural process are spontaneous and irreversible.

11. Ans. (2)

For spontaneity $\Delta G < 0$

$\Delta G = \Delta H - T\Delta S < 0$

For $\Delta G < 0$ ΔH = –ve

ΔS = +ve

Option (2) is correct

12. Ans. (3)

In isolated system, matter and energy transfer is not possible between system and surrounding.

Example $\Rightarrow$ Thermoflask

13. Ans. (4)

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

ΔH is contributory factor for ΔG, and as per theory.

Statement (1) is correct.

14. Ans. (2)

$\Delta G^0 = \Delta H^0 - T\Delta S^0$

+ve +ve

At high temp. ΔG decreases. So spontaneous.

with +ve entropy if $\Delta H^0 > T\Delta S^0$, $\Delta G^0 > 0$, so non spontaneous.

15. Ans. (2)

(2) option is correct as per theory.

16. Ans. (3)

(a) solid $\longrightarrow$ liquid entropy increase, intermolecular attraction decreases, at equilibrium $\Delta G = 0$

(b) $H_2O_{(s)} \xrightarrow[1atm]{10^\circ C} H_2O_{(l)}$

$T\uparrow, S\uparrow$

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

$\Delta G < 0$

(c) $H_2O_{(s)} \xrightarrow[1atm]{-10^\circ C} H_2O_{(l)}$

From solid to liquid, $\Delta S\uparrow$, $\Delta G = \Delta H - T\Delta S$

At less than zero degree celsius, can't convert solid into liquid without any external help, so ΔG should be +ve.

(d) $Ag^+ + HCl \longrightarrow AgCl_{(ppt)}$

So entropy decreases, process is spontaneous.