CHEMICAL THERMODYNAMICS

INTRODUCTION

1. Identify the intensive quantity from the following -
(1) Enthalpy and temperature
(2) Volume and temperature
(3) Enthalpy and volume
(4) Temperature and refractive index

2. Which of the following is an extensive property
(1) Mass (2) Enthalpy
(3) Energy (4) All of these

3. For an adiabatic process which of the following relations is correct
(1) $\Delta E = 0$ (2) $P\Delta V = 0$
(3) $q = 0$ (4) $q = +W$

4. In which of the following process work is independent of path :
(1) Isothermal (2) Isochoric
(3) Adiabatic (4) Isobaric

5. Out of boiling point (I), entropy (II), pH (III) and emf of a cell (IV), intensive properties are :
(1) I, III, IV (2) I, II
(3) I, II, III (4) All of these

6. Enthalpy of 1 mole monoatomic ideal gas is equals to :–
(1) $\frac{3}{2}RT$ (2) $\frac{5}{2}RT$ (3) RT (4) 2 RT

7. Which statement is true for reversible process :-
(1) It takes place in single step
(2) Driving force is much greater than opposing force
(3) Work obtain is minimum
(4) None

FIRST LAW OF THERMODYNAMICS ($\Delta E = q + W$)

8. If work done by the system is 300 joule when 100 cal. heat is supplied to it. The change in internal energy during the process is :-
(1) – 200 Joul (2) 400 Joul
(3) 720 Joul (4) 120 Joul

9 A system has internal energy equal to E_1, 450 J of heat is taken out of it and 600 J of work is done on it. The final energy of the system will be -
(1) $(E_1 + 150)$ (2) $(E_1 + 1050)$
(3) $(E_1 - 150)$ (4) None of these

10. The work done by a system is 8J when 40J heat is supplied to it. The change in internal energy of the system during the process :
(1) 32 J (2) 40 J (3) 48 J (4) –32 J

ENTHALPY [$\Delta H = \Delta E + P\Delta V / \Delta H = \Delta E + \Delta n_g RT$]

11. For a gaseous reaction,
$A(g) + 3B(g) \longrightarrow 3C(g) + 3D(g)$
ΔE is 17 kCal at 27°C assuming $R = 2\ Cal\ K^{-1}\ mol^{-1}$, the value of ΔH for the above reaction is:
(1) 15.8 Kcal (2) 18.2 Kcal
(3) 20.0 Kcal (4) 16.4 Kcal

12. Which of the following statements is correct for the reaction; $CO(g) + \frac{1}{2} O_2(g) \longrightarrow CO_2(g)$ at constant temperature and pressure
(1) $\Delta H = \Delta E$ (2) $\Delta H < \Delta E$
(3) $\Delta H > \Delta E$ (4) None of the above

13. For the reaction $Ag_2O(s) \longrightarrow 2Ag(s) + \frac{1}{2}O_2(g)$, which one of the following is true :
(1) $\Delta H = \Delta E$ (2) $\Delta H = \frac{1}{2}\Delta E$
(3) $\Delta H < \Delta E$ (4) $\Delta H > \Delta E$

14. For the reversible isothermal expansion of one mole of an ideal gas at 300 K, from a volume of 10 dm^3 to 20 dm^3, ΔH is -
(1) 1.73 kJ (2) –1.73 kJ
(3) 3.46 kJ (4) Zero

15. The reaction :-
$$NH_2CN_{(S)} + \frac{3}{2}O_{2(g)} \rightarrow N_{2(g)} + CO_{2(g)} + H_2O_{(\ell)}$$
was carried out in a bomb caloriemeter. The heat released was 743 kJ mol^{-1}. The value of ΔH_{300k} for this reaction would be :-
(1) – 740.5 kJ mol^{-1} (2) – 741.75 kJ mol^{-1}
(3) – 743.0 kJ mol^{-1} (4) – 744.25 kJ mol^{-1}

16. The heat of combustion of ethanol determined in a bomb calorimeter is – 670.48 kCal $mole^{-1}$ at 27°C. What is ΔH at 27°C for the reaction :–
(1) – 335.24 kCal (2) – 671.08 kCal
(3) – 670.48 kCal (4) + 670.48 kCal

17. For which of the following reactions ΔH is less than ΔE :–
(1) $C_{12}H_{22}O_{11}(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(\ell)$
(2) $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$
(3) $N_2O_4(g) \rightarrow 2NO_2(g)$
(4) $N_2(g) + O_2(g) \rightarrow 2NO(g)$

18. For a reaction $2X(s) + 2Y(s) \rightarrow 2C(\ell) + D(g)$

The q_p at 27°C is – 28 kCal mol^{-1}.

The q_V is ---------- kCal mol^{-1} :–

(1) – 27.4 (2) + 27.4

(3) – 28.6 (4) 28.6

WORK DONE IN DIFFERENT PROCESS

19. The work done in ergs for a reversible expansion of one mole of an ideal gas from a volume of 10 litres to 20 litres at 25^0C is :

(1) $-2.303 \times 8.31 \times 10^7 \times 298 \log 2$

(2) $-2.303 \times 0.0821 \times 298 \log 2$

(3) $-2.303 \times 0.0821 \times 298 \log 0.5$

(4) $-2.303 \times 2 \times 298 \log 2$

20. Two moles of an ideal gas expand spontaneouly into vacuum. The work done is :–

(1) Zero (2) 2 J (3) 4 J (4) 8 J

21. One mole of a gas occupying $3dm^3$ expands against a constant external pressure of 1 atm to a volume of 13 lit. The workdone is :–

(1) – 10 atm dm^3 (2) – 20 atm dm^3

(3) – 39 atm dm^3 (4) – 48 atm dm^3

ENTROPY/SECOND LAW OF THERMODYNAMICS

22. An adiabatic reversible process is one in which :-

(1) Temperature of the system does not change

(2) The system is not closed to heat transfer

(3) There is no entropy change

(4) None of these

23. Entropy means

(1) Disorderness (2) Randomness

(3) Orderness (4) both 1 & 2

24. Change in entropy is negative for

(1) Bromine $(\ell) \longrightarrow$ Bromine (g)

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

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

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

25. In which reaction ΔS is positive :-

(1) $H_2O (\ell) \rightarrow H_2O (s)$

(2) $3O_2 (g) \rightarrow 2O_3 (g)$

(3) $H_2O (\ell) \rightarrow H_2O (g)$

(4) $N_2(g) + 3H_2(g) \rightarrow 2NH_3 (g)$

26. The enthalpy of vaporization for water is 186.5 kJ mol^{-1}, the entropy of its vaporization will be -

(1) 0.5 kJK^{-1} mol^{-1} (2) 1.0 kJK^{-1} $mole^{-1}$

(3) 1.5 kJ K^{-1} $mole^{-1}$ (4) 2.0 kJK^{-1} $mole^{-1}$

27. If 900J/g of heat is exchanged at boiling point of water, then what is increase in entropy?

(1) 43.4 $JK^{-1}mole^{-1}$ (2) 87.2 $JK^{-1}mole^{-1}$

(3) 900 $JK^{-1}mole^{-1}$ (4) Zero

28. In a spontaneous irreversible process the total entropy of the system and surroundings

(1) Remains constant (2) Increases

(3) Decreases (4) Zero

29. Calculate the entropy of $Br_2(g)$ in the reaction $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$, $\Delta S° = 20.1 JK^{-1}$ given, entropy of H_2 and HBr is 130.6 and 198.5 J $mol^{-1} K^{-1}$:-

(1) 246.3 JK^{-1} (2) 123.15 JK^{-1}

(3) 24.63 JK^{-1} (4) 20 KJK^{-1}

30. In which of the following case entropy decreases–

(1) Solid changing to liquid

(2) Expansion of a gas

(3) Crystals dissolve

(4) Polymerisation

31. Entropy of an adiabatic reversible process is:-

(1) Positive (2) Zero

(3) Negative (4) Constant

GIBBS FREE ENERGY

32. For a reaction at 25°C enthalpy change (ΔH) and entropy change (ΔS) are -11.7×10^3 $Jmol^{-1}$ and –105 J mol^{-1} K^{-1} respectively. The reaction is :

(1) Spontaneous (2) Non spontaneous

(3) At equilibrium (4) Can't say anything

33. If $\Delta H > 0$ and $\Delta S > 0$, the reaction proceeds spontaneously when :-

(1) $\Delta H > 0$ (2) $\Delta H < T \Delta S$

(3) $\Delta H = T\Delta S$ (4) None

34. The enthalpy change for a given reaction at 298 K is –x cal mol^{-1}. If the reaction occurs spontaneously at 298 K, the entropy change at that temperature

(1) Can be negative but numerically larger than x/298 Cal K^{-1} mol^{-1}

(2) Can be negative but numerically smaller than x/298 Cal K^{-1} mol^{-1}

(3) Cannot be negative

(4) Cannot be positive

35. For the reaction $Ag_2O(s) \longrightarrow 2Ag(s) + \frac{1}{2} O_2(g)$ the value of $\Delta H = 30.56$ kJ mol^{-1} and $\Delta S = 66$ JK^{-1} mol^{-1}. The temperature at which the free energy change for the reaction will be zero is :–

(1) 373 K (2) 413 K
(3) 463 K (4) 493 K

36. What is the free energy change ΔG, when 1.0 mole of water at 100°C and 1 atm pressure is converted into steam at 100°C and 1 atm pressure :–

(1) 540 Cal (2) –9800 Cal
(3) 9800 Cal (4) 0 Cal

37. The Vant Hoff equation is :

(1) $\Delta G° = RT \log_e K_P$ (2) $-\Delta G° = RT \log_e K_P$
(3) $\Delta G° = RT^2 \ell n K_P$ (4) None

38. The Gibbs free energy change of a reaction at 27°C is –26 kCal. and its entropy change is – 60 Cals K. ΔH for the reaction is :–

(1) – 44 kCals. (2) – 18 kCals.
(3) 34 kals. (4) – 24 kCals.

HESS LAW

39. Given $C(s) + O_2(g) \longrightarrow CO_2(g) + 94.2$ kCal

$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(\ell) + 68.3$ kCal

$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(\ell) + 210.8$ kCal

The heat of formation of methane in Kcal will be

(1) –45.9 (2) –47.8
(3) –20.0 (4) –47.3

40. If, $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$; $\Delta H^0 = -44$ kCal

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

$\Delta H = -152$ kCal

Then, $Na(s) + 0.5\ Cl_2(g) \longrightarrow NaCl(s)$; $\Delta H^0 = ?$

(1) 108 kCal (2) 196 kCal
(3) – 98 kCal (4) 54 kCal

41. If $S + O_2 \longrightarrow SO_2$; $\Delta H = -298.2$

$SO_2 + \frac{1}{2} O_2 \longrightarrow SO_3$; $\Delta H = -98.7$

$SO_3 + H_2O \longrightarrow H_2SO_4$; $\Delta H = -130.2$

$H_2 + \frac{1}{2} O_2 \longrightarrow H_2O$; $\Delta H = -287.3$

Then the enthalpy of formation of H_2SO_4 at 298 K is -

(1) –814.4 kJ (2) –650.3 kJ
(3) –320.5 kJ (4) –433.5 kJ

42. Given that :

$Zn + \frac{1}{2} O_2 \rightarrow ZnO + 84000$ Cal 1

$Hg + \frac{1}{2} O_2 \rightarrow HgO + 21700$ Cal 2

The heat of reaction (ΔH) for,

$Zn + HgO \rightarrow ZnO + Hg$ is :-

(1) 105700 Cal (2) 62300 Cal
(3) –105700 Cal (4) –62300 Cal

43. The heat of reaction for

$A + \frac{1}{2} O_2 \rightarrow AO$ is - 50 kCal and

$AO + \frac{1}{2} O_2 \rightarrow AO_2$ is 100 kCal.

The heat of reaction for $A + O_2 \rightarrow AO_2$ is:-

(1) – 50 kCal (2) + 50 kCal
(3) 100 kCal (4) 150 kCal

44. Using the following thermochemical data:

$C(S) + O_2(g) \rightarrow CO_2(g)$, $\Delta H = -94.0$ kCal

$H_2(g) + 1/2O_2(g) \rightarrow H_2O(\ell)$, $\Delta H = -68.0$ kCal

$CH_3COOH(\ell) + 2O_2(g) \rightarrow 2CO_2(g) + 2H_2O(\ell)$, $\Delta H = -210.0$ kCal

The heat of formation of acetic acid is:-

(1) 116.0 kCal (2) - 116.0 kCal
(3) - 114.0 kCal (4) + 114.0 kCal

45. $H_2(g) + \frac{1}{2}O_2(g) = H_2O(\ell)$; $\Delta H_{298K} = -68.32$kCal. Heat of vapourisation 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

(1) 10.52 (2) –78.84
(3) +57.80 (4) –57.80

46. The heat of solution of anhydrous $CuSO_4$ and $CuSO_4.5H_2O$ are –15.89 and 2.80 kCal mol^{-1} respectively. What will be the heat of hydration of anhydrous $CuSO_4$?

(1) –18.69 kCal (2) 18.69 kCal
(3) –28.96 kCal (4) 28.96 kCal

47. Which of the following expressions is true:-

(1) $H_f^0 (CO,g) = \frac{1}{2} \Delta H_f^0 (CO_2,g)$

(2) $\Delta H_f^0 (CO,g) = \Delta H_f^0 (C,graphite) + \frac{1}{2} \Delta H_f^0 (O_2,g)$

(3) $\Delta H_f^0 (CO,g) = \Delta H_f^0 (CO_2,g) - \frac{1}{2} \Delta H_f^0 (O_2,g)$

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

ANSWER KEY

CHEMICAL THERMODYNAMICS

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	4	4	3	3	1	2	4	4	1	1	2	2	4	4	2	2	2	3	1	1
Que.	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**	**31**	**32**	**33**	**34**	**35**	**36**	**37**	**38**	**39**	**40**
Ans.	1	3	4	4	3	1	1	2	1	4	4	2	2	2	3	4	2	1	3	3
Que.	**41**	**42**	**43**	**44**	**45**	**46**	**47**													
Ans.	1	4	2	3	4	1	4													

SOLUTION

1. Intensive equantity : Propherties that does not depends upon amount ie Temperature, Refractive index

2. Extensive properties depends upon amount of matter. ie mass, enthalpy, energy etc.

3. For adiabatic process q = 0

4. For adiabatic process q = 0

$\omega = \Delta U = nC_V\Delta T$

5. Intensive properties : Boiling point, pH, emf of cell.

6. 1 mole mono atomic ideal gas

$$\Delta H = nC_p\Delta T = 1\times\frac{5}{2}R\Delta T$$

$$= \frac{5}{2}RT$$

7. Reversible process take place in ∞ step. Driving force is almost equal to opposing force and work obtained maximum.

8. $\Delta U = q + w$

$= (100\text{Cal}) + (-300\text{ Joule})$

$= (100 \times 4.2\text{ joule}) - 300\text{ joule}$

$= 120\text{ Joule}$

9. $\Delta U = q + \omega$

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

$E_2 = E_1 + 150$

10. $\Delta U = q + \omega$

$= (+40J) + (-8J)$

$= 32\text{ J}$

11. $A(g) + 3B(g) \rightarrow 3C(g) + 3D(g)$

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

$\Delta H = \Delta U + \Delta ng\,RT$

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

$= 18.2\text{ Kcal}$

12. $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$

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

$\Delta H = \Delta E - \frac{1}{2}RT$

$\Delta H < \Delta E$

13. $Ag_2O(s) \longrightarrow 2Ag(s) + \frac{1}{2}O_2(g);\ \Delta n_g = \frac{1}{2}$

$\Delta H = \Delta E + \Delta n_gRT$

$\Delta H = \Delta E + \frac{1}{2}\times RT$, $\Delta H > \Delta E$

14. For isothermal ideal gas

$\Delta H = nCp\Delta T = 0 \quad (\Delta T = 0)$

15. $NH_2CN(s) + \frac{3}{2}O_2(g) \rightarrow N_2(g) + CO_2(g) + H_2O(\ell)$

$\Delta ng = 1 + 1 - \frac{3}{2} = \frac{1}{2}$

Bomb calorimeter means

$\Delta E = -743\text{ kJ/mole}$

$\Delta H = \Delta E + \Delta ng\,RT$

$= -743\text{kJ} + \frac{1}{2} \times 8.314 \times 10^{-3}\text{ kJ} \times 300$

$= -7.41.75\text{ kJ/mole}$

16. $C_2H_5OH(\ell) + 3O_2(g) \rightarrow 2CO_2(g) + 3\,H_2O(\ell)$

$\Delta ng = -1$

Bomb calorimeter means

$\Delta E = -670.48\text{ kcal/mole}$

$\Delta H = \Delta E + \Delta ng\,RT$

$= -670.48\text{ kcal} - 1 \times 2 \times 10^{-3}\text{ kcal} \times 300$

$= -671.08\text{ kcal/mole}$

17. $\Delta H < \Delta E$ if $\Delta ng < 0$

i.e. $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

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

18. $2X(s) + 2Y(s) \rightarrow 2C(\ell) + D(g)$

$\Delta ng = 1$

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

$q_p = q_v + (1) \times 2 \times 10^{-3}\text{ kcal} \times 300$

$q_v = q_p - 0{\cdot}6\text{ KCal}$

$= -28\text{ kcal} - 0{\cdot}6\text{Kcal}$

$= -28.6\text{ Kcal}$

19. $W_{rev}(\text{isothermal}) = -nRT\ \ell n\left(\frac{v_2}{v_1}\right)$

$= -1 \times 8.31 \times 10^7\text{ ergs} \times 298.\ \ell n\left(\frac{20}{10}\right)$

$= -2.303 \times 8.31 \times 10^7 \times 298.\ \ell n(2)\text{ ergs}$

20. w(vacuum) = 0

21. $w_{isobaric} = -p_{ext}(v_2 - v_1)$

$= -1\text{atm }(13\text{ lit} - 3\text{lit})$

$= -10\text{ lit atm}$

$= -10\text{ atm dm}^3\ (1\text{dm}^3 = 1\text{ lit})$

22. Adiabatic reversible process $q = 0$

$\Delta S = 0$

23. Entropy means dis orderness or randomness

24. $\Delta S < 0$ if $\Delta ng < 0$ for reaction

$P\uparrow$ or $T\downarrow$ for state change

25. $\Delta S > 0$ if $\Delta ng > 0$ for reaction
or melting occures.

26. $$\Delta S_{vap} = \frac{\Delta H_{vap}}{T_{BP}} = \frac{186.5 kJ}{373} = 0.5 kJ / mole\text{-}K$$

$$\Delta S_{H_2O}(\text{at B.P.}) = \frac{\Delta H_{vap}}{T_{BP}} = \frac{9000\,J/gm}{373\,K}$$

$$= \frac{9000 \times 18}{373\,K} J/mole$$

= 43.4 J/mole-K

28. ΔS_{total} (irreversible) > 0

29. $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$

$$\Delta S^{\circ}_{Re^n} = \Sigma S^{\circ}_p - \Sigma S^{\circ}_R$$

$$20.1\ J/k = S^{\circ}_{H_2} + S^{\circ}_{Br_2} - 2 \times S^{\circ}_{HBr}$$

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

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

30. In polymerisation entrpy decreases

31. Adiabatic reversible process

$\Delta S_{total} = 0$

means entropy remains constant.

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

$= -11.7 \times 10^3 J - 298 \times -105 J$

$= -11700 + 298 \times 105$

> 0

So non spontaneous.

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

given $\Delta H > 0$

$\Delta S > 0$

$\Delta G = (+ve) - T.\Delta S$

ΔG will be –ve if $\Delta H < T.\Delta S$

34. For spontaneous process

$\Delta G < 0$

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

$(-x) - 298.\Delta S < 0$

$-298.\Delta S < +x$

$$\Delta S > \left(\frac{-x}{298}\right)$$

ΔS can be –ve but numericaly

smaller than $\frac{x}{298}$ cal/k

35. $Ag_2O(s) \rightarrow 2Ag(s) + -O\ (g)$

$\Delta G = 0$

$\Delta H - T.\Delta S = 0$

$$T = \frac{\Delta H}{\Delta S} = \frac{30.56 \times 10^3 J}{66\ J/k}$$

T = 463 K

36. $H_2O(\ell) \rightleftharpoons H_2O(g)$

1mole 1mole

(1atm, 100°C) (1atm, 100°C)

Reversible phase transfer take place at equilibrium so $\Delta G = 0$

37. Vant Haff equation is

$\Delta G^{\circ} = -RT\ \ell n\ kp$

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

-26×10^3 cal $= \Delta H - 300 \times 60$ cal/k

$\Delta H = -26000$ cal $- 18000$ cal

$= -44$ Kcal

39. From given reaction it can be easily prodicted that

$\Delta H_f(CO_2) = \Delta H_C(C) = -94.2$ Kcal

$\Delta H_f(H_2O) = \Delta H_C(H_2) = -68.3$ Kcal

$\Delta H_C(CH_4) = -210.8$ Kcal

Now heat of formation of CH_4

$C(s) + 2H_2(g) \rightarrow CH_4(g)$; ΔH

$\Delta H_f\ (CH_4) = \Delta H$

$= \Sigma\ H_C(R) - \Sigma H_C(p)$

$= \Delta H_C(C) + 2 \times \Delta H_C(H_2) - \Delta H_C(CH_4)$

$= -94.2 + 2 \times -68.3 - (-210.8)$

$= -20$ kcal

40. From given equation

$$(III) = \frac{(II)}{2} + \frac{(I)}{2}$$

$$\Delta H^{\circ} = \left(\frac{-44}{2}\right) + \left(\frac{-152}{2}\right)$$

$= -22\ -76$

$= -98$ Kcal

41. Enthalpy of formation of H_2SO_4

$H_2(g) + S(s) + 2O_2(g) \rightarrow H_2SO_4$; ΔH ...(v)

(V) = (I) + (II) + (III) + (IV)

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

$\Delta H = -814.4$ kJ $= \Delta H_f(H_2SO_4)$.

42. From given equation

(III) = (I) – (II)

$\Delta G = -84000 - (-21700)$

$= -62300$ cal.

43. $A + \frac{1}{2}O_2 \rightarrow AO$; $\Delta H_1 = -50$ kcal

$AO + \frac{1}{2}O_2 \rightarrow AO_2$; $\Delta H_2 = 100$ kcal

$A + O_2 \rightarrow AO_2$; ΔH.

(III) = (I) + (II)

$\Delta H = -50 + 100 = +50$ kcal

44. from given reaction

$\Delta H_f(CO_2) = \Delta H_C(C) = -94$ kcal

$\Delta H_f(H_2O) = \Delta H_C(H_2) = -68$ kcal

$\Delta H_C(CH_3COOH) = -210$ kcal

Heat of formation of CH_3COOH

$2C(S) + 2H_2(g) + O_2(g) \rightarrow CH_3COOH$; ΔH

ΔH_f $(CH_3COOH) = \Delta H = \Sigma\Delta H_C(R)\ \Sigma\Delta H_C(P)$

$= 2 \times -94 + 2 \times -68 + 0 - (-210)$

$= -114$ Kcal

45. $H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(\ell)$ ...(I)

$\Delta H_1 = -68.32$ kcal

$H_2O(\ell) \longrightarrow H_2O(V)$ ...(II)

$\Delta H_2 = 10.52$ kcal

$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(V)$ ...(IV)

$\Delta H = \Delta H_f$ $(H_2O(v))$

(III) = (I) + (II)

$\Delta H_f(H_2O(v)) = -68.32 + 10.52$

$= -57.80$ Kcal

46. $CuSO_4 + aq \rightarrow CuSO_4(aq)$ ....(I)

$\Delta H_1 = -15.89$ kcal

$CuSO_4.5H_2O + aq \rightarrow CuSO_4(aq) + 5H_2O(\ell)$ ...(II)

$\Delta H_2 = +2.80$ kcal

$CuSO_4(s) + 5H_2O(\ell) \rightarrow CuSO_4.5H_2O(s)$ ...(III)

ΔH.

(III) = (I) – (II)

$\Delta H = -15.89 - 2.80$

$= -18.69$ kcal

47. $C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$; $\Delta H_f = (CO, g)$ ...(I)

$C(s) + O_2(g) \rightarrow CO_2(g)$; ΔH_f $(CO_2, g) = \Delta H_C$ $(C(s))$ ...(II)

ΔH_f $(O_2(g)) = 0$ ...(III)

$CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$; $\Delta H_C(CO(g))$ ...(IV)

(I) + (IV) = (II)

ΔH_f $(CO(g)) + \Delta H_C$ $(CO(g)) = \Delta H_C$ $(C(s))$