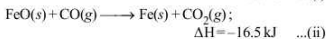
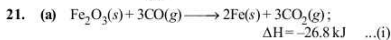


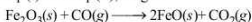
Speed Test-34

1. (b) In insulated system no heat or matter is exchanged, $q = 0$. $\therefore \Delta E = W$. The work done on the system increases internal energy.
2. (c) $\Delta H = H_2 - H_1 = (E_2 + P_2 V_2) - (E_1 + P_1 V_1)$
 $= (E_2 - E_1) + (P_2 V_2 - P_1 V_1)$
 $= 30 + 4 \times 5 - 2 \times 3 = 44 \text{ L atm}$
3. (d) Given $\Delta H = 41 \text{ kJ mol}^{-1} = 41000 \text{ J mol}^{-1}$
 $T = 100^\circ\text{C} = 273 + 100 = 373 \text{ K}$
 $n = 1$
 $\Delta U = \Delta H - \Delta nRT = 41000 - (1 \times 8.314 \times 373)$
 $= 37898.88 \text{ J mol}^{-1} \approx 37.9 \text{ kJ mol}^{-1}$
4. (a) Process is isothermal reversible expansion, hence $\Delta U = 0$, therefore $q = -W$.
 Since $q = +208 \text{ J}$, $W = -208 \text{ J}$
5. (b) An intensive property is a physical quantity whose value does not depend on the amount of the substance for which it is measured.
6. (d) Kirchhoff's equation, $\frac{\Delta H_2 - \Delta H_1}{T_2 - T_1} = \Delta C_p$
7. (a) Heat capacity at constant volume (q_v) = ΔE
 Heat capacity of constant pressure (q_p) = ΔH
 $\Delta H = \Delta E + \Delta nRT$ or $\Delta H - \Delta E = \Delta nRT$
 $\Delta n = \text{no. of moles of gaseous products}$
 $\quad - \text{no. of moles of gaseous reactants}$
 $= 12 - 15 = -3$
 $\Delta H - \Delta E = -3 \times 8.314 \times 298 \text{ J} = -7.43 \text{ kJ}$
8. (b) Desired equation is $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$, $\Delta H = ?$
 (Equation II - Equation I)
 $\Delta H = -245.5 \text{ kJ} - (-286 \text{ kJ}) = 40.5 \text{ kJ}$
9. (a) $\text{C}_2\text{H}_5\text{OH}(\text{l}) + 3\text{O}_2(\text{g}) \longrightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$
 Bomb calorimeter gives ΔU of the reaction
 Given, $\Delta U = -1364.47 \text{ kJ mol}^{-1}$
 $\Delta n_g = -1$
 $\Delta H = \Delta U + \Delta n_g RT = -1364.47 - \frac{1 \times 8.314 \times 298}{1000}$
 $= -1366.93 \text{ kJ mol}^{-1}$
10. (a) $\Delta S_{(A \rightarrow B)} = \Delta S_{(A \rightarrow C)} + \Delta S_{(C \rightarrow D)} - \Delta S_{(B \rightarrow D)}$
 $= 50 + 30 - 20 = 60 \text{ e.u.}$
11. (a) $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$; $\Delta H = 20 \text{ kcal} \dots (1)$
 $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$; $\Delta H = -40 \text{ kcal} \dots (2)$
12. (b) Enthalpy of reaction
 $= B.E_{(\text{Reactant})} - B.E_{(\text{Product})}$
 $= [B.E_{(\text{C}=\text{C})} + 4 B.E_{(\text{C}-\text{H})} + B.E_{(\text{H}-\text{H})}]$
 $\quad - [B.E_{(\text{C}-\text{C})} + 6 B.E_{(\text{C}-\text{H})}]$
 $= [606.1 + (4 \times 410.5) + 431.37] - [336.49 + (6 \times 410.5)]$
 $= -120.0 \text{ kJ mol}^{-1}$
13. (a) $\Delta G = \Delta H - T\Delta S$, ΔH is +ve, ΔS is +ve; $T\Delta S > \Delta H$ for spontaneous process. It will make ΔG -ve
14. (d) $\Delta H = \Delta E + \Delta nRT$ For $\Delta H \neq \Delta E$, $\Delta n \neq 0$
 Where $\Delta n = \text{no. of moles of gaseous products} - \text{no. of moles of gaseous reactants}$
 (a) $\Delta n = 2 - 2 = 0$
 (b) $\Delta n = 0$ (\because they are either in solid or liquid state)
 (c) $\Delta n = 1 - 1 = 0$ (\because C is in solid state)
 (d) $\Delta n = 2 - 4 = -2$
 \therefore (d) is correct answer
15. (b) The species in its elemental form has zero standard molar enthalpy of formation at 298 K. At 298K, Cl_2 is gas while Br_2 is liquid.
16. (b) High pressure reduces volume, decreases entropy, hence ΔS negative.
17. (d) $\Delta S^\circ = 2S^\circ_{\text{HCl}} - (S^\circ_{\text{H}_2} + S^\circ_{\text{Cl}_2})$
 $= 2 \times 186.7 - (130.6 + 223.0)$
 $= 19.8 \text{ JK}^{-1} \text{ mol}^{-1}$
18. (a) $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$
 For a spontaneous reaction $\Delta G^\circ < 0$
 or $\Delta H^\circ - T\Delta S^\circ < 0 \Rightarrow T > \frac{\Delta H^\circ}{\Delta S^\circ}$
 $\Rightarrow T > \frac{179.1 \times 10^3}{160.2} > 1117.9 \text{ K} \approx 1118 \text{ K}$
19. (d) For the reaction, $\text{C}_{(\text{s})} + \frac{1}{2} \text{O}_{2(\text{g})} \longrightarrow \text{CO}$
 $\Delta H = \Delta U + \Delta nRT$ or $\Delta H - \Delta U = \Delta nRT$
 $\Delta n = 1 - \frac{1}{2} = \frac{1}{2}$; $\Delta H - \Delta U = \frac{1}{2} \times 8.314 \times 298$
 $= 1238.78 \text{ J mol}^{-1}$
20. (c) $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$; $-RT \ln K = \Delta H^\circ - T\Delta S^\circ$

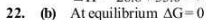
$$nK = -\frac{\Delta H^\circ - T\Delta S^\circ}{RT}$$



eq. (i) $- 2 \times$ eq. (ii), we get



$\Delta H = -26.8 + 33.0 = +6.2 \text{ kJ}$



Hence, $\Delta G = \Delta H - T_e \Delta S = 0$

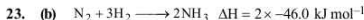
$\therefore \Delta H = T_e \Delta S$ or $T_e = \frac{\Delta H}{\Delta S}$

For a spontaneous reaction

ΔG must be negative which is possible only if

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

$\therefore \Delta H < T\Delta S$ or $T > \frac{\Delta H}{\Delta S}$; $T_e < T$



Let x be the bond enthalpy of $\text{N}-\text{H}$ bond then

[Note : Enthalpy of formation or bond formation enthalpy is given which is negative but the given reaction involves bond breaking hence values should be taken as positive.]

$\Delta H = \Sigma \text{Bond energies of reactants} - \Sigma \text{Bond energies of products}$

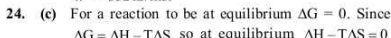
$2 \times -46 = 712 + 3 \times (436) - 6x$

$-92 = 2020 - 6x$

$6x = 2020 + 92$

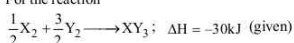
$6x = 2112$

$x = 352 \text{ kJ/mol}$



or $\Delta H = T\Delta S$

For the reaction



Calculating ΔS for the above reaction, we get

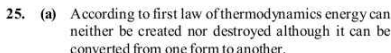
$\Delta S = 50 - \left[\frac{1}{2} \times 60 + \frac{3}{2} \times 40 \right] \text{ JK}^{-1}$

$= 50 - (30 + 60) \text{ JK}^{-1} = -40 \text{ JK}^{-1}$

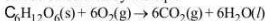
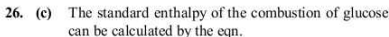
At equilibrium, $T\Delta S = \Delta H$ [$\because \Delta G = 0$]

$\therefore T \times (-40) = -30 \times 1000$ [$\because 1 \text{ kJ} = 1000 \text{ J}$]

or $T = \frac{-30 \times 1000}{-40}$ or 750 K



NOTE: Carnot cycle is based upon this principle but during the conversion of heat into work some mechanical energy is always converted to other form of energy hence this data violates 1st law of thermodynamics.



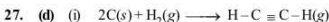
$\Delta H_c^\circ = 6 \times \Delta H_f^\circ(\text{CO}_2) + 6 \times \Delta H_f^\circ(\text{H}_2\text{O}) - \Delta H_f^\circ[\text{C}_6\text{H}_{12}\text{O}_6]$

$\Delta H_c^\circ = 6(-400) + 6(-300) - (-1300)$

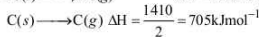
$\Delta H_c^\circ = -2900 \text{ kJ/mol}$

For one gram of glucose, enthalpy of combustion

$\Delta H_c^\circ = -\frac{2900}{180} = -16.11 \text{ kJ/g}$



$\Delta H = 225 \text{ kJ mol}^{-1}$



From equation (i):

$225 = [2 \times \Delta H_{\text{C}(s)} \rightarrow \text{C}(g)] + 1 \times \text{BE}_{\text{H}-\text{H}}$

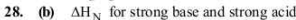
$-[2 \times \text{BE}_{\text{C}-\text{H}} + 1 \times \text{BE}_{\text{C}=\text{C}}]$

$225 = [1410 + 1 \times 330] - [2 \times 350 + 1 \times \text{BE}_{\text{C}=\text{C}}]$

$225 = [1410 + 330] - [700 + \text{BE}_{\text{C}=\text{C}}]$

$225 = 1740 - 700 - \text{BE}_{\text{C}=\text{C}}$

$\text{BE}_{\text{C}=\text{C}} = 815 \text{ kJ mol}^{-1}$



$= -13.7 \text{ kcal eq}^{-1}$

$\Delta H_{\text{ion}}(\text{CH}_3\text{COOH})$

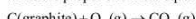
$= -12.5 - (-13.7) = 1.2 \text{ kcal mol}^{-1}$

$\Delta H_{\text{ion}}(\text{NH}_4\text{OH})$

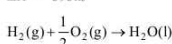
$= -10.5 - (-13.7) - \Delta H_{\text{ion}}(\text{CH}_3\text{COOH})$

$= 13.7 - 10.5 - 1.2$

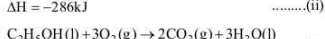
$= 2 \text{ kcal mol}^{-1}$



$\Delta H = -393 \text{ kJ} \dots (i)$

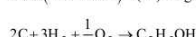


$\Delta H = -286 \text{ kJ} \dots (ii)$



$\Delta H = -1360 \text{ kJ} \dots (iii)$

From $(2 \times I + 3 \times II) - (III)$ we get



$\therefore 2(-393 \text{ kJ}) + 3(-286 \text{ kJ}) - (-1360 \text{ kJ}) = -284 \text{ kJ}$



$= 1 - 180 = -179 \text{ kcal mol}^{-1}$

$$\text{Then } \Delta H_{\text{hyd}}(\text{Na}^+) + \Delta H_{\text{hyd}}(\text{Cl}^-) = -179$$

$$\text{or } \Delta H_{\text{hyd}}(\text{Na}^+) + \frac{2}{3} \Delta H_{\text{hyd}}(\text{Na}^+) = -179$$

$$\text{or } \Delta H_{\text{hyd}}(\text{Na}^+) = -107.4 \text{ kcal mol}^{-1}$$

31. (a) For non spontaneous reaction

$$\Delta G = +ve$$

$$\Delta G = \Delta H - T\Delta S \text{ and}$$

$$\Delta S = 121 \text{ J K}^{-1}$$

$$\text{For } \Delta G = +ve$$

ΔH has to be positive. Hence the reaction is endothermic.

The minimum value of ΔH can be obtained by putting

$$\Delta G = 0$$

$$\Delta H = T\Delta S = 298 \times 121 \text{ J}$$

$$= 36.06 \text{ kJ}$$

32. (c) The standard enthalpy of reaction is the enthalpy change for a reaction when all the participating substances are in their standard states. The standard state of a substance at a specified temperature is its pure form at 1 bar. For example, the standard state of liquid ethanol at 298 K is pure liquid ethanol at 1 bar. Standard state of solid iron at 500 K is pure iron at 1 bar.

33. (d)

$$34. (a) \Delta C_p = 2 \times 25.1 + 3 \times 75.3 - [103.8 + 3 \times 28.8]$$

$$= 85.9 \text{ J K}^{-1} \text{ mol}^{-1}$$

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

$$\frac{\Delta H_2 - (-33290)}{358 - 298} = 85.9$$

$$\Delta H = -28136 \text{ J mol}^{-1}$$

$$= -28.136 \text{ kJ mol}^{-1}$$

35. (b) The enthalpy of ionisation of weak acid is given by

$$\Delta H_{\text{ion}}(\text{HA})$$

$$= \Delta H_{\text{N(weak acid/strong base)}} - \Delta H_{\text{N(strong acid/strong base)}}$$

$$= -56.1 - (-57.3) = 1.2 \text{ kJ mol}^{-1}$$

$$\Delta H_{\text{(ionisation)}} = 1.5 \text{ kJ mol}^{-1}$$

Hence % ionisation in 1 M solution

$$= \frac{(1.5 - 1.2)}{1.5} \times 100 = 20$$

$$36. (d) \frac{1}{2} \text{N}_2(\text{g}) + \frac{3}{2} \text{F}_2(\text{g}) \longrightarrow \text{NF}_3; \Delta H = -113 \text{ kJ}$$

$$\text{or } \Delta H_{\text{N=N}} + \frac{3}{2} \Delta H_{\text{F-F}} - 3 \Delta H_{\text{N-F}} = -113 \text{ kJ}$$

Let $x \text{ kJ mol}^{-1}$ be the bond energy of F-F bond then bond energy of N=N bond = $6x$

$$\therefore \frac{1}{2} \times 6x + \frac{3}{2} \times x - 3 \times 273 = -113 \text{ kJ}$$

On solving,

$$x = 156.9 \text{ kJ mol}^{-1} \text{ and}$$

$$\text{N=N bond energy} = 6 \times 156.9 = 941.4 \text{ kJ mol}^{-1}$$

37. (b) Given ΔH



To calculate ΔH operate

$$2 \times \text{eq. (1)} + \text{eq. (2)} - \text{eq. (3)}$$

$$\Delta H = 300 - 125 - 350 = -175$$

$$38. (a) q = -W = 2.303nRT \log \frac{V_2}{V_1}$$

39. (c)

40. (a) The first reaction is exothermic and the second reaction is endothermic. On passing the mixture of O_2 and H_2O (steam) over coke while keeping the temperature constant ΔH of both the reactions must be same. Moles of O_2 needed to evolve 132 kJ

$$= \frac{0.5 \times 132}{110} = 0.6$$

Hence steam : O_2 ratio must be 1 : 0.6

41. (a) The greater the (negative value) of heat of neutralisation, the more is the strength of the acid. Hence, $\text{HCOOH} > \text{CH}_3\text{COOH} > \text{H}_2\text{S} > \text{HCN}$

42. (b) Dissolution of KNO_3 is endothermic, hence heat is absorbed and cooling is observed.

43. (d) The positive sign expresses when work is done on the system. Similarly, negative sign expresses when work is done by the system.

44. (b) This is combustion reaction, which is always exothermic hence

$$\Delta H = -ve$$

As the no. of gaseous molecules are increasing hence entropy increases so $\Delta S = +ve$.

$$\text{now } \Delta G = \Delta H - T\Delta S$$

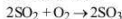
For a spontaneous reaction

$$\Delta G = -ve$$

which is possible in this case as $\Delta H = -ve$ and $\Delta S = +ve$.

45. (c) ΔS has negative value if number of gaseous moles decreases during a reaction, $\Delta n_g = -ve$

For the reaction



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