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ALTIS VORTEX

1. A gas is allowed to expand in a well insulated container against a constant external pressure of 2.5 atm from an initial volume of 2.50 L to a final volume of 4.50 L. The change in internal energy ΔU of the gas in joules will be: (Delhi-2017)
- a. +505 J b. 1136.25 J
c. -500 J d. -505 J
2. For a given reaction, $\Delta H = 35.5 \text{ KJ mol}^{-1}$ and $\Delta S = 83.6 \text{ JK}^{-1} \text{ mol}^{-1}$. The reaction is spontaneous at: (Assume that ΔH and ΔS do not vary with temperature) (Delhi-2017)
- a. $T > 298 \text{ K}$ b. $T < 425 \text{ K}$
c. $T > 425 \text{ K}$ d. All temperatures
3. Of the following, the largest value of entropy at 25°C and 1 atm is that of: (Gujarat-2017)
- a. CH_4 b. H_2
c. C_2H_6 d. C_2H_2
4. Under isothermal and reversible conditions, the term "free energy" in thermodynamics signifies: (Gujarat-2017)
- a. Expansion work done on the system
b. Non-expansion work done by the system
c. Expansion work done by the system
d. Non expansion work done on the system
5. For a sample of perfect gas when its pressure is changed isothermally from P_i to P_f , the entropy change is given by: (2016 - II)
- a. $\Delta S = nRT \ln \left(\frac{P_f}{P_i} \right)$
b. $\Delta S = nRT \ln \left(\frac{P_i}{P_f} \right)$
c. $\Delta S = nR \ln \left(\frac{P_f}{P_i} \right)$
d. $\Delta S = nR \ln \left(\frac{P_i}{P_f} \right)$
6. The correct thermodynamic conditions for the spontaneous reaction at all temperatures is: (2016 - I)
- 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 $\Delta S > 0$
7. Consider the following liquid-vapour equilibrium.
- Liquid \rightleftharpoons Vapour. Which of the following relations is correct? (2016-I)
- a. $\frac{d \ln P}{dT^2} = \frac{-\Delta H_v}{T^2}$ b. $\frac{d \ln P}{dT} = \frac{-\Delta H_v}{RT^2}$
c. $\frac{d \ln G}{dT^2} = \frac{\Delta H_v}{RT^2}$ d. $\frac{d \ln P}{dT} = \frac{-\Delta H_v}{RT}$
8. The heat of combustion of carbon to CO_2 is -393.5 kJ/mol . The heat released upon formation of 35.2 g of CO_2 from carbon and oxygen gas is: (2015)
- a. + 315 kJ b. - 630 kJ
c. - 3.15 kJ d. - 315 kJ
9. For the reaction, $\text{X}_2\text{O}_4(\text{l}) \rightarrow 2\text{XO}_2(\text{g})$, $\Delta U = 2.1 \text{ kcal}$, $\Delta S = 20 \text{ cal K}^{-1}$ at 300 K. Hence, ΔG is: (2014)
- a. -2.7 kcal b. 9.3 kcal
c. -9.3 kcal d. 2.7 kcal
10. In which of the following reactions, standard reaction entropy change (ΔS°) is positive and standard Gibb's energy change (ΔG°) decreases sharply with increasing temperature? (2012 Pre)
- a. $\frac{1}{2} \text{C (graphite)} + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$
b. $\text{C (graphite)} + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{CO}(\text{g})$
c. $\text{CO}(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$
d. $\text{Mg}(\text{s}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{MgO}(\text{s})$

11. The enthalpy of fusion of water is 1.435 kcal/mol. The molar entropy change for the melting of ice at 0°C: (2012 Pre)
- a. 0.526 cal/mol K b. 10.52 cal/mol K
c. 21.04 cal/mol K d. 5.260 cal/mol K
12. Standard enthalpy of vapourisation ΔH_{vap} 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: (2012 Pre)
- a. +40.66 b. +37.56
c. -43.76 d. +43.76
13. Consider the following process
- $$\frac{1}{2} \text{A} \rightarrow \text{B} \quad \Delta H(\text{kJ/mol}) = +150$$
- $$3\text{B} \rightarrow 2\text{C} + \text{D} \quad \Delta H(\text{kJ/mol}) = -125$$
- $$\text{E} + \text{A} \rightarrow 2\text{D} \quad \Delta H(\text{kJ/mol}) = +350$$
- For $\text{B} + \text{D} \rightarrow \text{E} + 2\text{C}$, ΔH will be: (2011 Mains)
- a. -325 kJ/mol b. 325 kJ/mol
c. 525 kJ/mol d. -175 kJ/mol
14. Enthalpy change for the reaction, $4\text{H}(\text{g}) \rightarrow 2\text{H}_2(\text{g})$ is -869.6 kJ. The dissociation energy of H - H bond is: (2011 Pre)
- a. +217.4 kJ b. -434.8 kJ
c. -869.6 kJ d. +434.8 kJ
15. 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: (2011 Pre)
- a. 100 J mol⁻¹ K⁻¹ b. 10 J mol⁻¹ K⁻¹
c. 1.0 J mol⁻¹ K⁻¹ d. 0.1 J mol⁻¹ K⁻¹
16. Which of the following is correct option for free expansion of an ideal gas under adiabatic condition? (2011 Pre)
- a. $q = 0, \Delta T < 0, w \neq 0$
b. $q = 0, \Delta T \neq 0, w = 0$
c. $q \neq 0, \Delta T = 0, w \neq 0$
d. $q = 0, \Delta T = 0, w = 0$
17. The following two reactions are known:
- $$\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + 3\text{CO}_2(\text{g}); \Delta H = -26.8 \text{ kJ}$$
- $$\text{FeO}(\text{s}) + \text{CO}(\text{g}) \rightarrow \text{Fe}(\text{s}) + \text{CO}_2(\text{g}); \Delta H = -16.5 \text{ kJ}$$
- The value of ΔH for the following reaction $\text{Fe}_2\text{O}_3(\text{s}) + \text{CO}(\text{g}) \rightarrow 2\text{FeO}(\text{s}) + \text{CO}_2(\text{g})$ is: (2010 Mains)
- a. +10.3 kJ b. -43.3 kJ
c. -10.3 kJ d. +6.2 kJ
18. For vaporization of water at 1 atmospheric pressure, the values of ΔH and ΔS are 40.63 kJ mol⁻¹ and 108.8 JK⁻¹mol⁻¹, respectively. The temperature when Gibbs energy change (ΔG) for this transformation will be zero, is: (2010 Mains)
- a. 273.4 K b. 393.4 K
c. 373.4 K d. 293.4 K
19. Three moles of an ideal gas expanded spontaneously into vacuum. The work done will be: (2010 Mains)
- a. Infinite b. 3 Joules
c. 9 Joules d. Zero
20. Match List-I (Equations) with List-II (Type of process) and select the correct option: (2010 Mains)
- | List-I Equations | List-II Type of processes |
|------------------------------------|-----------------------------------|
| A. $K_p > Q$ | (i) Non-spontaneous |
| B. $\Delta G^\circ < RT \ln Q$ | (ii) Equilibrium |
| C. $K_p = Q$ | (iii) Spontaneous and endothermic |
| D. $T = \frac{\Delta H}{\Delta S}$ | (iv) Spontaneous |
- a. A(i) B(ii) C(iii) D(iv)
b. A(iii) B(iv) C(ii) D(i)
c. A(iv) B(i) C(ii) D(iii)
d. A(ii) B(i) C(iv) D(iii)
21. For an endothermic reaction, energy of activation is E_a and enthalpy of reaction is ΔH (both of these in kJ/mol). Minimum value of E_a will be: (2010 Pre)
- a. Equal to zero b. Less than ΔH
c. Equal to ΔH d. More than ΔH
22. Standard entropies of X_2 , Y_2 and XY_3 are 60, 40 and 50 JK⁻¹mol⁻¹ respectively. For the reaction
- $$\frac{1}{2}\text{X}_2 + \frac{3}{2}\text{Y}_2 \rightleftharpoons \text{XY}_3, \Delta H = 30 \text{ kJ}$$
- to be at equilibrium, the temperature should be: (2010 Pre)
- a. 500 K b. 750 K
c. 1000 K d. 1250 K
23. The pressure exerted by 6.0 g of methane gas in a 0.03 m³ vessel at 129°C is: (Atomic masses : C = 12.01, H = 1.01 and R = 8.314 JK⁻¹ mol⁻¹): (2010 Mains)
- a. 215216 Pa b. 13409 Pa
c. 41648 Pa d. 31684 Pa

24. From the following bond energies:

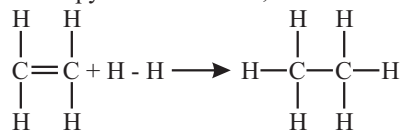
$$\text{H} - \text{H} \text{ bond energy: } 431.37 \text{ kJ mol}^{-1}$$

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

$$\text{C} - \text{C} \text{ bond energy: } 336.49 \text{ kJ mol}^{-1}$$

$$\text{C} - \text{H} \text{ bond energy: } 410.50 \text{ kJ mol}^{-1}$$

Enthalpy for the reaction,



will be:

(2009)

- a. $-243.6 \text{ kJmol}^{-1}$ b. $-120.0 \text{ kJmol}^{-1}$
c. $-553.0 \text{ kJmol}^{-1}$ d. $-1523.6 \text{ kJmol}^{-1}$

25. The values of ΔH and ΔS for the reaction, $\text{C}_{(\text{graphite})} + \text{CO}_2(\text{g}) \rightarrow 2\text{CO}(\text{g})$ are 170 kJ and 170 JK^{-1} , respectively. This reaction will be spontaneous at:

(2009)

- a. 910 K b. 1000 K
c. 510 K d. 710 K

26. For the gas phase reaction, $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$, Which of the following conditions is correct?

(2008)

- 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 $\Delta S < 0$

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

(2008)

- a. 245 kJmol^{-1} b. 93 kJmol^{-1}
c. -245 kJmol^{-1} d. -93 kJmol^{-1}

28. Which of the following are not state functions?

(2008)

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

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

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

29. Given that bond energies of $\text{H} - \text{H}$ and $\text{Cl} - \text{Cl}$ are 430 kJ mol^{-1} and 240 kJ mol^{-1} respectively and ΔH_f for HCl is -90 kJ mol^{-1} , bond enthalpy of HCl is:

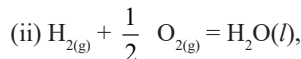
(2007)

- a. 380 kJ mol^{-1} b. 425 kJ mol^{-1}
c. 245 kJ mol^{-1} d. 290 kJ mol^{-1}

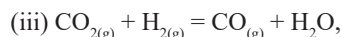
30. Consider the following reactions:



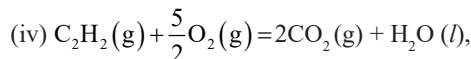
$$\Delta H = -X_1 \text{ kJ mol}^{-1}$$



$$\Delta H = -X_2 \text{ kJ mol}^{-1}$$



$$\Delta H = -X_3 \text{ kJ mol}^{-1}$$



$$\Delta H = -X_3 \text{ kJ mol}^{-1}$$

Enthalpy of formation of $\text{H}_2\text{O}(l)$ is: (2007)

- a. $+X_3 \text{ kJ mol}^{-1}$ b. $-X_4 \text{ kJ mol}^{-1}$
c. $+X_1 \text{ kJ mol}^{-1}$ d. $-X_2 \text{ kJ mol}^{-1}$

31. The enthalpy and entropy change for the reaction: $\text{Br}_2(l) + \text{Cl}_2(\text{g}) \rightarrow 2\text{BrCl}(\text{g})$ are 30 kJmol^{-1} and $105 \text{ JK}^{-1} \text{ mol}^{-1}$ respectively. The temperature at which the reaction will be in equilibrium is:

(2006)

- a. 285.7 K b. 373 K
c. 250 K d. 400 K

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

(2006)

- a. If $\Delta G_{\text{system}} < 0$, the process is not spontaneous
b. If $\Delta G_{\text{system}} > 0$, the process is spontaneous
c. If $\Delta G_{\text{system}} = 0$, the system has attained equilibrium
d. If $\Delta G_{\text{system}} = 0$, then system is still moving in a particular direction

33. Assume each reaction is carried out in an open container. For which reaction will $\Delta H = \Delta E$?

(2006)

- a. $2\text{CO}(\text{g}) + \text{O}_{2(\text{g})} \rightarrow 2\text{CO}_{2(\text{g})}$
b. $\text{H}_{2(\text{g})} + \text{Br}_{2(\text{g})} \rightarrow 2\text{HBr}(\text{g})$
c. $\text{C}(\text{s}) + 2\text{H}_2\text{O}(\text{g}) \rightarrow 2\text{H}_{2(\text{g})} + \text{CO}_{2(\text{g})}$
d. $\text{PCl}_{5(\text{g})} \rightarrow \text{PCl}_{3(\text{g})} + \text{Cl}_{2(\text{g})}$

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

(2006)

- a. $-358.5 \text{ kJ mol}^{-1}$ b. $-508.9 \text{ kJ mol}^{-1}$
c. $-208.1 \text{ kJ mol}^{-1}$ d. $-269.9 \text{ kJ mol}^{-1}$

35. The absolute enthalpy of neutralisation of the reaction:
 $\text{MgO(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{O(l)}$ will be: (2005)
- Less than $-57.33 \text{ kJ mol}^{-1}$
 - $-57.33 \text{ kJ mol}^{-1}$
 - Greater than $-57.33 \text{ kJ mol}^{-1}$
 - $57.33 \text{ kJ mol}^{-1}$
36. Which of the following pairs of a chemical reaction is certain to result in a spontaneous reaction? (2005)
- Exothermic and decreasing disorder
 - Endothermic and increasing disorder
 - Exothermic and increasing disorder
 - Endothermic and decreasing disorder
37. A reaction occurs spontaneously if: (2005)
- $T\Delta S < \Delta H$ and both ΔH and ΔS are +ve
 - $T\Delta S > \Delta H$ and both ΔH and ΔS are +ve
 - $T\Delta S = \Delta H$ and both ΔH and ΔS are +ve
 - $T\Delta S > \Delta H$ and ΔH is +ve and ΔS is -ve
38. If the bond energies of H – H, Br – Br and H – Br are 433, 192 and 364 kJ mol^{-1} respectively, the ΔH° for the reaction (2004)
- $$\text{H}_2\text{(g)} + \text{Br}_2\text{(g)} \rightarrow 2\text{HBr(g)}$$
- is:-
- + 103 kJ
 - + 261 kJ
 - 103 kJ
 - 261 kJ
39. Considering entropy (S) as a thermodynamic parameter, the criterion for the spontaneity of any process is:- (2004)
- $\Delta S_{\text{system}} - \Delta S_{\text{surroundings}} > 0$
 - $\Delta S_{\text{System}} > 0$ only
 - $\Delta S_{\text{surroundings}} > 0$ only
 - $\Delta S_{\text{system}} + \Delta S_{\text{surroundings}} > 0$
40. Standard enthalpy and standard entropy changes for the oxidation of ammonia at 298 K are $-382.64 \text{ kJ mol}^{-1}$ and $-145.6 \text{ JK}^{-1} \text{ mol}^{-1}$, respectively. Standard Gibbs energy change for the same reaction at 298 K is:- (2004)
- $-339.3 \text{ kJ mol}^{-1}$
 - $-439.3 \text{ kJ mol}^{-1}$
 - $-523.2 \text{ kJ mol}^{-1}$
 - $-221.1 \text{ kJ mol}^{-1}$
41. 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 (2004)
- 608 J
 - + 304 J
 - 304 J
 - 6 J
42. For the reaction:
 $\text{C}_3\text{H}_8\text{(g)} + 5\text{O}_2\text{(g)} \rightarrow 3\text{CO}_2\text{(g)} + 4\text{H}_2\text{O(l)}$ at constant temperature, $\Delta H - \Delta E$ is :- (2003)
- +RT
 - 3RT
 - +3RT
 - RT
43. The densities of graphite and diamond at 298 K are 2.25 and 3.31 g cm^{-3} , respectively. If the standard free energy difference (ΔG°) is equal to 1895 J mol^{-1} , the pressure at which graphite will be transformed into diamond at 298 K is: (2003)
- $9.92 \times 10^8 \text{ Pa}$
 - $9.92 \times 10^7 \text{ Pa}$
 - $9.92 \times 10^6 \text{ Pa}$
 - $9.92 \times 10^5 \text{ Pa}$
44. What is the entropy change (in $\text{J K}^{-1} \text{ mol}^{-1}$) 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): (2003)
- 20.13
 - 2.013
 - 2.198
 - 21.98
45. Formation of a solution from two components can be considered as:-
- Pure solvent \rightarrow separated solvent molecules, ΔH_1
 - Pure solvent \rightarrow separated solvent molecules, ΔH_2
 - Separated solvent and solute molecules solution, ΔH_3
- Solution so formed will be ideal if: (2003)
- $\Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 + \Delta H_3$
 - $\Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 - \Delta H_3$
 - $\Delta H_{\text{soln}} = \Delta H_1 - \Delta H_2 - \Delta H_3$
 - $\Delta H_{\text{soln}} = \Delta H_3 - \Delta H_1 - \Delta H_2$
46. For which one of the following equations is $\Delta H^\circ_{\text{react}}$ equal to $\Delta H^\circ_{\text{p}}$ for the product: (2003)
- $\text{N}_2\text{(g)} + \text{O}_3\text{(g)} \rightarrow \text{N}_2\text{O}_3\text{(g)}$
 - $\text{CH}_4\text{(g)} + 2\text{Cl}_2\text{(g)} \rightarrow \text{CH}_2\text{Cl}_2\text{(l)} + 2\text{HCl(g)}$
 - $\text{Xe(g)} + 2\text{F}_2\text{(g)} \rightarrow \text{XeF}_4\text{(g)}$
 - $2\text{CO(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{CO}_2\text{(g)}$
47. The molar heat capacity of water at constant pressure, C_p , is $75 \text{ JK}^{-1} \text{ mol}^{-1}$. 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: (2003)
- 1.2 K
 - 2.4 K
 - 4.8 K
 - 6.6 K
48. Unit of entropy is: (2002)
- $\text{JK}^{-1} \text{ mol}^{-1}$
 - J mol^{-1}
 - $\text{J}^{-1}\text{K}^{-1} \text{ mol}^{-1}$
 - JK mol^{-1}

49. In a closed insulated container a liquid is stirred with a paddle to increase the temperature, which of the following is true? (2002)
- $\Delta E = W \neq 0, q = 0$
 - $\Delta E = W = q \neq 0$
 - $\Delta E = 0, W = q \neq 0$
 - $W = 0 \Delta E = q \neq 0$
50. 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): (2002)
- 92.1
 - 0
 - 4
 - 9.2
51. Heat of combustion ΔH^0 for C(s), H₂(g) and CH₄(g) are - 94, - 68 and - 213 Kcal/mol, then ΔH^0 for C (s) + 2H₂(g) → CH₄(g) is: (2002)
- 17 Kcal
 - 111 Kcal
 - 170 Kcal
 - 85 Kcal
52. Which reaction is not feasible? (2002)
- 2KI + Br₂ → 2KBr + I₂
 - 2 KBr + I₂ → 2KI + Br₂
 - 2 KBr + Cl₂ → 2KCl + Br₂
 - 2H₂O + 2F₂ → 4HF + O₂
53. Change in enthalpy for reaction
2H₂O₂ (l) → 2H₂O (l) + O₂(g)
If heat of formation of H₂O₂(l) and H₂O (l) are -188 & - 286 KJ/mol respectively: (2001)
- 196 KJ/mol
 - + 196 KJ/mol
 - + 948 KJ/mol
 - 948 KJ/mol
54. When 1 mol 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? (2001)
- q = w = 500J, ΔU = 0
 - q = ΔU = 500J, w = 0
 - q = w = 500 J, AU = 0
 - ΔU = 0, q = w = - 500 J
55. Enthalpy of $\text{CH}_4 + \frac{1}{2}\text{O}_2 \rightarrow \text{CH}_3\text{OH}$ is negative. If enthalpy of combustion of CH₄ and CH₃OH are x and y respectively. Then which relation is correct? (2001)
- x > y
 - x < y
 - x = y
 - x ≥ y
56. $\text{PbO}_2 \rightarrow \text{PbO}$; $\Delta G_{298} < 0$
 $\text{SnO}_2 \rightarrow \text{SnO}$; $\Delta G_{298} > 0$
Most probable oxidation state of Pb & Sn will be: (2001)
- Pb⁺⁴, Sn⁺⁴
 - Pb⁺², Sn⁺²
 - Pb⁺², Sn⁺⁴
 - Pb⁺⁴, Sn⁺²
57. For the disproportionation of copper
 $2\text{Cu}^+ \rightarrow \text{Cu}^{+2} + \text{Cu}$, E⁰ is:- (Given E⁰ for Cu⁺²/Cu is 0.34 V & E⁰ for Cu⁺²/Cu⁺ is 0.15 V) (2000)
- 0.49 V
 - 0.19 V
 - 0.38 V
 - 0.38 V
58. Cell reaction is spontaneous when: (2000)
- ΔG⁰ is negative
 - ΔG⁰ is positive
 - ΔE⁰_{Red} is positive
 - ΔE⁰_{red} is negative
59. $2\text{Zn} + \text{O}_2 \rightarrow 2\text{ZnO}$; ΔG⁰ = - 616 J
 $2\text{Zn} + \text{S}_2 \rightarrow 2\text{ZnS}$; ΔG⁰ = - 293 J
 $\text{S}_2 + 2\text{O}_2 \rightarrow 2\text{SO}_2$; ΔG⁰ = - 408 J
ΔG⁰ for the following reaction is: (2000)
- $$2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2$$
- 731 J
 - 1317 J
 - 501 J
 - + 731 J
60. At 27° C, latent heat of fusion of a compound is 2930 J/mol. Entropy change is: (2000)
- 9.77 J/mol-K
 - 10.77 J/mol-K
 - 9.07 J/mol-K
 - 0.977 J/mol-K
61. For the reaction $\text{C}_2\text{H}_5\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(l)$ which one is true? (2000)
- ΔH = ΔE - RT
 - ΔH = ΔE + RT
 - ΔH = ΔE + 2RT
 - ΔH = ΔE - 2RT
62. In an endothermic reaction, the value of ΔH is: (1999)
- Zero
 - Positive
 - Negative
 - Constant
63. In the reaction
 $\text{S}(s) + \frac{3}{2}\text{O}_2(g) \rightarrow \text{SO}_3(g) + 2x \text{ kcal}$
and $\text{SO}_2(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{SO}_3(s) + y \text{ kcal}$
the heat of formation of SO₂ is: (1999)
- (x + y)
 - (x - y)
 - (2x + y)
 - (2x - y)
64. At 25°C and 730 mm pressure, 380 ml of dry oxygen was collected. If the temperature is constant, what volume will the oxygen occupy at 760 mm pressure? (1999)
- 365 ml
 - 2g
 - 10g
 - 20g

65. Identify the correct statement regarding entropy: (1998)
- At absolute zero temperature, entropy of a perfectly crystalline substance is taken to be zero
 - At absolute zero of temperature, the entropy of a perfectly crystalline substance is +ve
 - At absolute zero of temperature, the entropy of all crystalline substance is to be zero
 - At 0° C, the entropy of a perfectly crystalline substance is taken to be zero
66. One mole of an ideal gas at 300K is expanded isothermally from an initial volume of 1L to 10L. The ΔE for this process is ($R = 2 \text{ cal mol}^{-1} \text{ K}^{-1}$): (1998)
- 163.7 cal
 - Zero
 - 1381.1 cal
 - 9 L atm
67. Given that $\text{C} + \text{O}_2 \rightarrow \text{CO}_2, \Delta H^0 = -x \text{ kJ}$
 $2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2, \Delta H^0 = -y \text{ kJ}$
 The enthalpy of formation of carbon monoxide will be: (1997)
- $\frac{y-2x}{2}$
 - $2x-y$
 - $y-2x$
 - $\frac{2x-y}{2}$
68. Which of the following is the correct equation? (1996)
- $\Delta U = \Delta W + \Delta Q$
 - $\Delta U = \Delta Q - W$
 - $\Delta W = \Delta U + \Delta Q$
 - None of these
69. If enthalpies of formation for $\text{C}_2\text{H}_4(\text{g})$, $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{l})$ at 25°C and 1 atm pressure are 52, -394 and -286 kJ/mol respectively, then enthalpy of combustion of $\text{C}_2\text{H}_4(\text{g})$ will be: (1995)
- +141.2 kJ/mol
 - +1412 kJ/mol
 - 141.2 kJ/mol
 - 1412 kJ/mol
70. A chemical reaction is catalyzed by a catalyst X. Hence X (1995)
- Reduces enthalpy of the reaction
 - Does not affect equilibrium constant of reaction
 - Decreases rate constant of the reaction
 - Increases activation energy of the reaction.
71. Standard state Gibb's free energy change for isomerization reaction *cis*-2-pentene \rightleftharpoons *trans*-2-pentene is -3.67 kJ/mol at 400 K. If more *trans*-2-pentene is added to the reaction vessel, then: (1995)
- Equilibrium remains unaffected
 - Equilibrium is shifted in the forward direction
 - More *cis*-2-pentene is formed
 - Additional *trans*-2-pentene is formed.
72. For a reaction to occur spontaneously: (1995)
- ΔH must be negative
 - ΔS must be negative
 - $(\Delta H - T\Delta S)$ must be negative
 - $(\Delta H + T\Delta S)$ must be negative.
73. The average kinetic energy of an ideal gas, per molecule in S.I. units, at 25°C will be: (1995)
- $6.17 \times 10^{-20} \text{ J}$
 - $7.16 \times 10^{-20} \text{ J}$
 - $61.7 \times 10^{-21} \text{ J}$
 - $6.17 \times 10^{-21} \text{ J}$
74. During isothermal expansion of an ideal gas, its: (1991, 94)
- Internal energy increases
 - Enthalpy decreases
 - Enthalpy remains unaffected
 - Enthalpy reduces to zero.
75. Following reaction occurring in an automobile $2\text{C}_8\text{H}_{18(\text{g})} + 25\text{O}_{2(\text{g})} \rightarrow 16\text{CO}_{2(\text{g})} + 18\text{H}_2\text{O}_{(\text{g})}$. The sign of ΔH , ΔS and ΔG would be: (1994)
- , +, +
 - +, +, -
 - +, -, +
 - Wd. -, +, -
76. For the reaction $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3, \Delta H = ?$ (1991)
- $\Delta E + 2RT$
 - $\Delta E - 2RT$
 - $\Delta H = RT$
 - $\Delta E - RT$
77. If ΔH is the change in enthalpy and ΔE , the change in internal energy accompanying a gaseous reaction, then: (1990)
- ΔH is always greater than ΔE
 - $\Delta H < \Delta E$ only if the number of moles of the products is greater than the number of moles of the reactants
 - ΔH is always less than ΔE
 - $\Delta H < \Delta E$ only if the number of moles of products is less than the number of moles of the reactants.

Answer Key

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
d	c	c	b	d	d	b	a	a	b	d	a	d	d	a	d	d
18	19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34
c	d	c	d	b	c	b	b	c	d	a	b	d	a	c	b	c
35	36	37	38	39	40	41	42	43	44	45	46	47	48	49	50	51
b	b	b	c	d	a	a	b	none	d	a	c	b	a	b	d	a
52	53	54	55	56	57	58	59	60	61	62	63	64	65	66	67	68
b	a	b	b	d	c	a	a	a	a	b	d	a	c	b	a	b
69	70	71	72	73	74	75	76	77								
d	b	c	c	d	c	d	b	d								

EXPLANATIONS

1. (d) NCERT (XI) Ch - 6, Pg. 160

$$\Delta U = q + W$$

For adiabatic process, $q = 0$

$$\therefore \Delta U = W \Rightarrow = -P \cdot \Delta V$$

$$= -2.5 \text{ atm} \times (4.50 - 2.50) \text{ L}$$

$$= -2.5 \times 2 \text{ L-atm} \Rightarrow -5 \times 101.3 \text{ J}$$

$$= -506.5 \text{ J} \Rightarrow -505 \text{ J}$$

2. (c) NCERT (XI) Ch - 6, Pg. 178

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

For spontaneous reaction, ΔG must be negative. For negative value of ΔG , ΔH should be greater than $T\Delta S$. It is possible when $T > 425 \text{ K}$.

$$\Delta H = 35.5 \text{ KJ/mol} = 35500$$

$$T\Delta S = (425)(83.6) = 35530$$

$$T\Delta S > \Delta H$$

Hence, the reaction is spontaneous

3. (c) NCERT (XI) Ch - 6, Pg. 175

The molecule having more number of bonds have largest value of entropy

$\therefore \text{C}_2\text{H}_6$ have large value of entropy.

4. (b) NCERT (XI) Ch - 6, Pg. 178

Under isothermal reversible conditions, the term “free-energy” in thermodynamics signifies that “No expansion work done by the system”.

$$-\Delta G_{(\text{system})} = W_{\text{non-expansion}}$$

5. (d) NCERT (XI) Ch - 6 Pg. 178

$$\Delta S = nRT \ln \frac{P_i}{P_f}$$

For isothermal process $T_i = T_f = 0$

$$\therefore \Delta S = nR \ln \frac{P_i}{P_f}$$

6. (d) NCERT (XI) Ch - 6, Pg. 178

For any spontaneous process, $\Delta G = (-ve)$ & $\Delta S = (+ve)$ that is increase in entropy

So, at $\Delta H < 0$ & $\Delta S > 0$ at all temperatures according the reaction will be spontaneous.

7. (b) NCERT (XI) Ch - 6

$$\frac{d \ln P}{dT} = \frac{-\Delta H_v}{RT^2} \quad (\text{this is clausius clapeyron equation.})$$

8. (a) NCERT (XI) Ch - 6, Pg. 165



when heat of combustion is given then

$$\Delta H_r = H_R - H_P$$

$$= 0 - (-393.5)$$

$$= 393.5 \text{ KJ/mole}$$

So heat of formation of $\text{CO}_2 = 393.5 \text{ KJ/mole}$
or we can say

Heat released on formation of 44 gm/O_2

$$= 393.5 \text{ KJ}$$

Heat released on formation of 35.2 gm of

$$\text{CO}_2 = \frac{393.5}{44} \times 35.2$$

$$= 314.8 \Rightarrow \approx 315 \text{ KJ}$$

9. (a) NCERT (XI) Ch - 6, Pg. 161

According to 1st law of law of thermodynamics

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

$$\Delta H = 2.1 + 2 \times 2 \times 300$$

$$= 3.3 \text{ kcal} = 3.3 - 300(20) = -2.7 \text{ kcal}$$

According to 3rd law of thermodynamics:

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

10. (b) NCERT (XI) Ch - 6, Pg. 176

As C changes from solid to gas, randomness increases. A significant decrease in Gibbs free energy will be observed.

11. (d) NCERT (XI) Ch - 6, Pg. 178

According to 3rd law of thermodynamics:

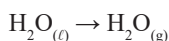
$$\Delta S_{\text{sys.}} = \frac{\Delta H}{T}$$

$$\Delta S = \frac{1.435 \times 10^3}{273} = 5.26 \text{ cal / K mol.}$$

12. (a) NCERT (XI) Ch - 6, Pg. 166

$$\Delta H_{\text{vap}}^{\circ} \text{ for water at } 100^{\circ}\text{C} = 40.66 \text{ KJ/mole}$$

At $100^{\circ}\text{C} / 373 \text{ K}$



$$\Delta n_g = 1$$

we know that $\Delta H = \Delta E + \Delta n_g RT$

$$\Delta E = \Delta H - \Delta n_g RT$$

$$\Delta E = (40.66 \times 10^3 \text{ J}) - (1 \times 8.314 \times 373)$$

$$\Delta E = 37599 \text{ J/mole} = 37.56 \text{ KJ/mole}$$

13. (d) NCERT (XI) Ch - 6, Pg. 173 - 175

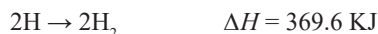
According to the data given -



$$\Delta H = 150 \times 2 + (-125) - 350$$

$$= -175 \text{ KJ}$$

14. (d) NCERT (XI) Ch - 6, Pg. 172



$$\text{For dissociation enthalpy of H-H} = \frac{869.6}{2} = 434.8 \text{ KJ}$$

15. (a) NCERT (XI) Ch - 6, Pg. 178

According to 3rd law of thermodynamics $\Delta S = \frac{\Delta H}{T}$

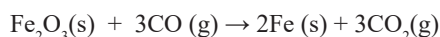
$$\Delta S_{\text{vap}} = \frac{\Delta H_{\text{vap}}}{T} = \frac{30 \text{ kJ}}{300} = 100 \text{ J mol}^{-1} \text{ K}^{-1}$$

16. (d) NCERT (XI) Ch - 6 Pg. 160

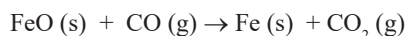
For free expansion of ideal gas

$$q = 0, \Delta T = 0 \text{ \& } w = 0$$

17. (d) NCERT (XI) Ch - 6, Pg. 171-173



$$\Delta H = -26.8 \text{ KJ} \quad \text{--- (1)}$$



$$\Delta H = -16.5 \text{ KJ} \quad \text{--- (2)}$$

For ΔH of $\text{Fe}_2\text{O}_3(\text{g}) + \text{CO}(\text{g}) \rightarrow 2\text{FeO}(\text{s}) + \text{CO}_2(\text{g})$

equation (1) - $2 \times$ (2)

$$= -26.8 - 2 \times 16.5 = +6.2 \text{ KJ}$$

18. (c) NCERT (XI) Ch - 6, Pg. 179

Gibbs free energy for transformation = 0

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

$$\Delta T = \frac{40.63 \times 10^3}{108.8} = 373.4 \text{ K}$$

19. (d) Work done in vacuum = 0 as $P_{\text{ext}} = 0$.

20. (c) NCERT (XI) Ch - 6, Pg. 178-179

a \rightarrow (iv) b \rightarrow (i) c \rightarrow (ii) d \rightarrow (iii)

$K_p > Q \leftarrow$ Reaction moves in forward direction.

$[\Delta G < RT \ln Q]$ $\Delta G = (+ve)$ reaction becomes non-spontaneous

$K_p = Q$ Reaction attains equilibrium

$T > \frac{\Delta H}{\Delta S} = \Delta H = (+ve) \Rightarrow$ endothermic reaction

When $\Delta H < T\Delta S$ reaction is spontaneous.

21. (d) NCERT (XI) Ch - 6, Pg. 115

For endothermic reaction, E_a activation energy

will be more than ΔH (difference between $E_f - E_b$), activation energy of forward reaction and backward.

22. (b) NCERT (XI) Ch - 6, Pg. 172, 178

$$\Delta S^\circ = \Sigma \Delta S^\circ_{\text{products}} - \Sigma \Delta S^\circ_{\text{Reactants}}$$

$$= 50 - (30 + 60)$$

$$\Delta S^\circ = -40 \text{ J}^{-1} \text{ mol}^{-1}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ, \text{ in equilibrium } \Delta G^\circ = 0$$

$$T = \frac{\Delta H^\circ}{\Delta S^\circ} = \frac{-30 \times 10^{-1} \text{ KJ mol}^{-1}}{-40 \times \text{JK}^{-1} \text{ mol}^{-1}} \Rightarrow 750 \text{ K}$$

23. (c) NCERT (XI) Ch - 6, Pg. 161

According to the ideal gas equation:

$$PV = nRT$$

$$P = \frac{nRT}{V} = \frac{6}{19} \times \frac{8.34 \times 402}{0.03} = 41648 \text{ Pa}$$

24. (b) NCERT (XI) Ch - 6, Pg. 172

For the given reaction, enthalpy of reaction can be calculated as

$$= \text{B.E. (reactant)} - \text{B.E. (Product)}$$

$$= [\text{B.E.}_{(\text{C}-\text{C})} + \text{B.E.}_{(\text{H}-\text{H})} + 4 \times \text{B.E.}_{(\text{C}-\text{H})}]$$

$$- [\text{B.E.}_{(\text{C}-\text{C})} + 6 \times \text{B.E.}_{(\text{C}-\text{H})}]$$

$$= [606.10 + 431.37 + 4 \times 410.50]$$

$$- [336.49 + 6 \times 410.50]$$

$$= 2679.47 - 2799.49 = -120.02 \text{ kJ mol}^{-1}$$

25. (b) NCERT (XI) Ch - 6, Pg. 178

For the reaction to be spontaneous, $\Delta G = -ve$,
Given, $\Delta H = 170 \text{ kJ} = 170 \times 10^3 \text{ J}$

$$\Delta S = 170 \text{ JK}^{-1} \text{ mol}^{-1}$$

Applying, $\Delta G = \Delta H - T\Delta S$, the value of $\Delta G = -ve$ only when $T\Delta S > \Delta H$, which is possible only when $T = 1110 \text{ K}$.

$$\therefore \Delta G = 170 \times 10^3 - (1110 \times 170) = -18700 \text{ J}$$

Thus, reaction is spontaneous at $T = 1110 \text{ K}$

26. (c) NCERT (XI) Ch - 6, Pg. 178



This is an endothermic reaction ΔH (+ve)

Also the number of moles of product is greater than that of reactant, $\Delta S > 0$ spontaneity increases.

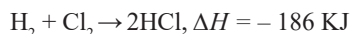
27. (d) NCERT (XI) Ch - 6, Pg. 173-174

According to the chemical equation:



$$\Delta H \text{ for } \text{H}_2 = 434 \text{ KJ}$$

$$\Delta H \text{ for } \text{Cl}_2 = 242 \text{ KJ}$$



$$\Delta H_f \text{ of HCl} = \frac{-186}{2} = -93 \text{ KJ mol}^{-1}$$

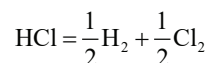
28. (a) NCERT (XI) Ch - 6, Pg. 156-157

Heat given to a system or by the system (Q) & work done (W) on the system or by the system are path functions, as they depend on the initial and final position.

$$\Delta U \text{ internal energy} = Q + W$$

$$\Delta G = \Delta H - T\Delta S \text{ are state functions.}$$

29. (b) NCERT (XI) Ch - 6, Pg. 172

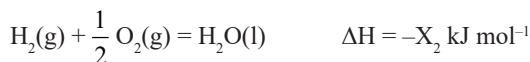


$$\Delta H = \Sigma \text{B.E.}_{(\text{products})} - \Sigma \text{B.E.}_{(\text{reactants})}$$

$$= \frac{1}{2}(430 + 240) - (-90) = \frac{1}{2} \times 670 + 90$$

$$= 335 + 90 = 425 \text{ kJ mol}^{-1}$$

30. (d) NCERT (XI) Ch - 6, Pg. 167



This reaction shows the formation of H_2O , and the X_2 represents the enthalpy of formation of H_2O , according to the definition, it suggests that the enthalpy of formation is the heat evolved or absorbed when one mole of substance is formed from its constituent atoms.

31. (a) NCERT (XI) Ch - 6, Pg. 178

According to 3rd Law of thermodynamics-

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

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

$$= \frac{30 \times 10^3 \text{ J}}{105} = 285.7 \text{ K}$$

32. (c) Criteria for spontaneity of a process in terms of ΔG is as following.

If $\Delta G = (-)ve \Rightarrow$ Process is spontaneous

If $\Delta G = (+)ve \Rightarrow$ Process does not occur in forward direction spontaneously but may occur spontaneously in backward direction.

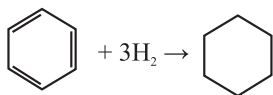
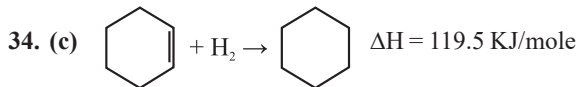
If $\Delta G = \text{zero} \Rightarrow$ system is in equilibrium.

$$33. \text{ (b) } \Delta H = \Delta E + \Delta n_g RT$$

if $\Delta n_g = \text{zero}$ then $\Delta H = \Delta E$

Δn_g for reaction, $\text{H}_{2(g)} + \text{B}_{2(g)} \rightarrow 2\text{HBr}_{(g)}$ is zero i.e.

$$\Delta H = \Delta E$$

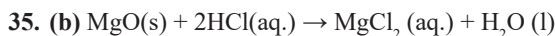


Enthalpy of hydrogenation of benzene

$$= (3 \times \Delta H) - (\text{Resonance energy})$$

$$= [3 \times (-119.5)] - [-150.4]$$

$$= -208.1 \text{ KJ/mole}$$



Enthalpy of neutralisation will be less than $-57.33 \text{ J mol L}^{-1}$ as per the experimental data. However, question is incomplete.

36. (b) For endothermic reaction ΔH is (+ve) & increasing disorderness or randomness increases the spontaneity of the reaction

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

$$(+ve) - (+ve) = (-ve) \text{ negative}$$

So, ΔG is (-ve) negative.

37. (b) For spontaneous reaction $\Delta H < 0$

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

$\Delta G = (-ve)$ for spontaneous reaction

$$\Delta S = (+ve), \Delta H = (+ve)$$

$$T\Delta S > \Delta H \Rightarrow \Delta G = (-ve)$$



$$433 + 192 \quad 2 \times 364$$

$$= 625 \quad = 728$$

Energy absorbed Energy released

$$\Delta H^0 = 728 - 625 = 103 \text{ kJ}$$

here $\Delta H^0 = -103 \text{ kJ}$

(-) represent released Heat

$$39. \text{ (d) } \Delta S = R \ln \frac{V_2}{V_1}$$

$$\Delta S_{\text{sys}} + \Delta S_{\text{surr}} > 0.$$

$$40. \text{ (a) } \Delta G = \Delta H - T\Delta S$$

$$-382.64 - 298 \left(\frac{-145.6}{1000} \right)$$

$$= -382.64 + 43.38$$

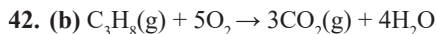
$$= -339.3 \text{ KJ mol}^{-1}$$

$$41. \text{ (a) } \text{Work} = -P_{\text{ext}} \times \text{Volume change}$$

$$= 3 \times 101.32 \times (6 - 4)$$

$$= 6 \times 101.32$$

$$= -607.92 \text{ J} \Rightarrow W = -608 \text{ J}$$



$$\Delta n_g = 3 - 6 = -3$$

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

$$\text{or } \Delta H - \Delta E = P\Delta V$$

$$\Delta H - \Delta E = \Delta n RT = -3RT$$

43. (None) $\Delta G = -P\Delta V = \text{work done (w)}$

$$\Delta V = \left(\frac{12}{3.31} - \frac{12}{2.25} \right) \times 10^{-3} \text{ L}$$

$$= 1.71 \times 10^{-3} \text{ L}$$

$$\Delta G = w = -(-1.71 \times 10^{-3}) \times P \times 101.3 \text{ J}$$

$$P = \frac{1895}{1.71 \times 10^{-3} \times 100.3}$$

$$= 10.93 \times 10^3 \text{ atm}$$

$$= 11.07 \times 10^8 \text{ pa}$$

No option is correct

$$44. \text{ (d) } S = \frac{q_{\text{rev}}}{T} = \frac{6000}{273} = 21.978 \text{ J/K}$$

45. (a) Heat of solution is defined as the amount of heat evolved or absorbed when one mole of the substance is dissolved in excess of the solvent. For hydrates salt & for salt which do not form hydrates, ΔH is +ve & for anhydrous salt, ΔH is -ve

$$46. \text{ (c) } \Delta H_{\text{reaction}}^0 = \Delta H_{\text{F}}^0(\text{XeF}_4) - [\Delta H_{\text{F}}^0(\text{Xe}) + 2\Delta H_{\text{F}}^0(\text{F}_2)]$$

Enthalpies of formation & elementary substances Xe & F_2 are taken as zero.

47. (b) Molar heat capacity = $75 \text{ JK}^{-1}\text{mol}^{-1}$

$$18 \text{ g of water} = 1 \text{ mole} = 75 \text{ JK}^{-1}\text{mol}^{-1}$$

$$1 \text{ g of water} = \frac{75}{18} \text{ JK}^{-1}$$

$$Q = m.c.\Delta t = \Delta t = \frac{10 \times 10}{75} = 2.4 \text{ K}$$

48. (a) Entropy (s) = $\frac{q}{T}$
 $= \text{JK}^{-1} \text{ mol}^{-1}$

49. (b) $q = \Delta E + W$
 As system is closed $q = 0$ and work is done on system
 $\therefore W \neq 0$
 Temperature and hence internal energy of the system increases.
 $\Delta E \neq 0$

50. (d) Change of entropy

$$ds = \frac{dq}{T}$$

$$\Delta S = C_v \ln \frac{T_2}{T_1} + R \ln \frac{V_2}{V_1} \dots\dots(1)$$

$$T_2 = T_1 = 27^\circ\text{C} = 300\text{K}$$

putting value in eq - (1)

$$\Delta S = 4.605 \text{ cal/mol}\cdot\text{K}$$

Entropy change for 2 mol of gas

$$= 2 \times 4.605 = 9.2 \text{ cal/K}$$

51. (a) (i) $\text{C (s)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$

$$\Delta H_i = -68 \times 2 \text{ k cal mol}^{-1}$$

(ii) $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}$

$$\Delta H_{ii} = -682 \text{ k cal mol}^{-1}$$

(iii) $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}$

$$\Delta H_{iii} = -213 \text{ kcal mol}^{-1}$$

(iv) $\text{C(s)} + 2\text{H}_2(\text{g}) \rightarrow \text{CH}_4(\text{g})$

$$\Delta H_{iv} = ?$$

Apply Hess's law

(i) + (ii) + (iii)

$$= (-94 - 68 \times 2 + 213) \text{ kcal} \Rightarrow -17 \text{ kcal}$$

52. (b) $2\text{KBr} + \text{I}_2 \rightarrow 2\text{KI} + \text{Br}_2$

As electronegative character decreases down the group, so option (b) is not feasible.

53. (a) $\Delta H_f^0 = \Sigma H_f^0(\text{product}) - \Sigma H_f^0(\text{reactant})$

for the given equation:



$$\Delta H_f^0 = 2 \times \Delta H_f^0(\text{H}_2\text{O}) - 2 \times \Sigma H_f^0(\text{H}_2\text{O}_2)$$

$$= 2 \times (-286 \text{ KJ mol}^{-1}) - 2 \times (-188) \text{ KJ mol}^{-1}$$

$$= -196 \text{ KJ mol}^{-1}$$

54. (b) we know that $\Delta H = \Delta E + \Delta PV$

$$\Delta H = \Delta E + P\Delta V + V\Delta P = 0$$

When $\Delta V = 0$; $W = 0$; $\Delta H = \Delta E + P\Delta V$

As $\Delta E = q + W$, $\Delta E = q$

In the present problem

$$\Delta H = 500\text{J}$$

$$\Delta H = \Delta E = 500\text{J}, q = 500\text{J}, W = 0$$

55. (b) $\text{CH}_4 + \frac{1}{2}\text{O}_2 \rightarrow \text{CH}_3\text{OH}$

$\Delta H = x - y$; given $\Delta H = \text{negative}$

Hence $x - y < 0$

56. (d) The sign and magnitude of Gibbs energy is a criterion of spontaneity for a process. When $\Delta G > 0$ or +ve, it means $G_{\text{product}} > G_{\text{reactant}}$, as $\Delta G = G_{\text{product}} + G_{\text{reactant}}$ the reaction will not take place spontaneously, i.e. the reaction should be spontaneous in reverse direction.



+ 4(more favourable) + 2

$\Delta G < 0$ or -ve, the reaction or change occurs spontaneously, $\text{PbO}_2 \rightarrow \text{PbO}; \Delta G < 0$

+4 +2

57. (c) For the reaction $2\text{Cu}^+ \rightarrow \text{Cu}^{2+} + \text{Cu}$

The cathode is Cu^+/Cu and anode is $\text{Cu}^+/\text{Cu}^{2+}$

Given, $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$; $E_1^0 = 0.34 \text{ V} \dots\dots(1)$

$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$; $E_2^0 = 0.15 \text{ V} \dots\dots(2)$

$\text{Cu}^+ + \text{e}^- \rightarrow \text{Cu}$; $E_3^0 = ? \dots\dots(3)$

$$\text{Now } \Delta G_1^0 = -nFE_1^0 = -2 \times 0.34 \times F$$

$$\Delta G_2^0 = -1 \times 0.15 \times F, \Delta G_3^0 = -1 \times E_3^0 \times F$$

$$\text{Again } \Delta G_1^0 = \Delta G_2^0 \Rightarrow -0.68F$$

$$= -0.15F - E_3^0 \times F$$

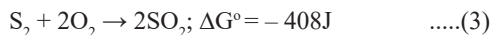
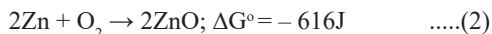
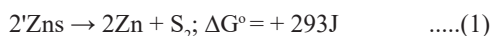
$$\Rightarrow E_3^0 = 0.68 - 0.15 = 0.53\text{V}$$

$$\text{As, } E_{\text{cell}}^0 = E_{\text{cathode}}^0 (\text{Cu}^+ / \text{Cu}) - E_{\text{anode}}^0 (\text{Cu}^{2+} / \text{Cu}^+)$$

$$= 0.53 - 0.15 = 0.38\text{V}$$

58. (a) For a cell reaction to be spontaneous ΔG^0 should be negative. As $\Delta G^0 = -nFE^0$, So the value will be negative only when E^0 is positive. n = number of electrons involved, F = Value of Faraday.

59. (a) For the reactions,



The ΔG° for the reaction,

$2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2$ can be obtained by adding eq. (1), (2) and (3)

$$\text{So, } \Delta G^\circ = 293 - 616 - 408 = -731\text{J}$$

60. (a) $\Delta S = \frac{Q}{T} = \frac{2930}{300} = 9.77\text{J/mol K}$

61. (a) $\Delta H = \Delta E + P\Delta V$

also $PV = nRT$ (ideal gas equation)

$$\text{or } P\Delta V = \Delta nRT$$

$\Delta n =$ change in number of gaseous moles

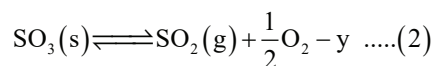
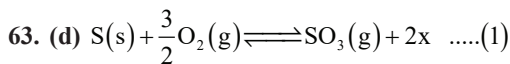
$$\therefore \Delta H = \Delta E + \Delta nRT = \Delta n_g \cdot 2-3 \Rightarrow -1$$

$$\Delta H = \Delta E - RT$$

62. (b) For endothermic rxn

$$E_R < E_P$$

$$\text{So, } \Delta H = E_P - E_R = +ve$$



On adding eq (1) & (2) we get



64. (a) *NCERT (XI) Ch - 6, Pg. 161*

According to Boyle's law:

$$P_1V_1 = P_2V_2$$

$$V_2 = \frac{730 \times 380}{760} = 365$$

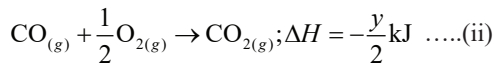
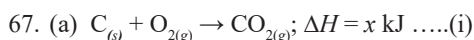
$$= 365 \text{ ml}$$

65. (c) At absolute zero temperature, entropy of a perfectly crystalline substance is taken to be zero. It is called third law of thermodynamics

66. (b) Isothermal means temperature is constant.

At constant temperature, change in internal energy (ΔE) remain constant.

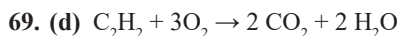
$$\text{So, } \Delta E = 0$$



By subtracting equation (ii) from (i) we get,

$$\Delta H = -x - \left(-\frac{y}{2}\right) = \frac{y-2x}{2} \text{ kJ}$$

68. (b) This is the mathematical relation of first law of thermodynamics. Here $\Delta U =$ Change in internal energy; $\Delta Q =$ Heat absorbed by the system and $W =$ Work done by the system.



$$\begin{aligned} \Delta H &= \Delta H_{\text{products}} - \Delta H_{\text{react}} \\ &= 2 \times (-394) + 2 \times (-286) - (52 + 0) \\ &= -1412 \text{ kJ/mol} \end{aligned}$$

70. (b) Since a catalyst affects equally on both forward and backward reactions, therefore it does not affect equilibrium constant of reaction.

71. (c) If more *trans*-2-pentene is added, then its concentration in right hand side will increase. But in order to maintain the constant K , concentration of *cis*-2-pentene will also increase. Therefore more *cis*-2-pentene will be formed.

72. (c) For a reaction to be spontaneous, ΔG (Gibbs free energy change) must be negative.

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

$\Delta H =$ change in enthalpy, $\Delta S =$ change in entropy.

73. (d) *NCERT (XI) Ch - 6*

$$\text{Temperature } (T) = 25^\circ\text{C} = 298 \text{ K.}$$

Therefore K.E. per molecule

$$= \frac{3RT}{2n} = \frac{3 \times 8.314 \times 298}{2 \times (6.02 \times 10^{23})}$$

$$= 6.17 \times 10^{-21} \text{ J}$$

74. (c) During isothermal expansion of an ideal gas, $\Delta T = 0$. Now we know $H = E + PV \therefore \Delta H = \Delta E + \Delta(PV) = \Delta E + \Delta(nRT) \therefore \Delta H = \Delta E + nR\Delta T; \Delta T = 0 + 0 = 0$

75. (d) (i) The given reaction is a combustion reaction, therefore ΔH is less than 0. Hence, ΔH is negative.

(ii) Since there is increase in the number of moles, therefore ΔS is positive

(iii) Since reaction is spontaneous, therefore ΔG is negative.

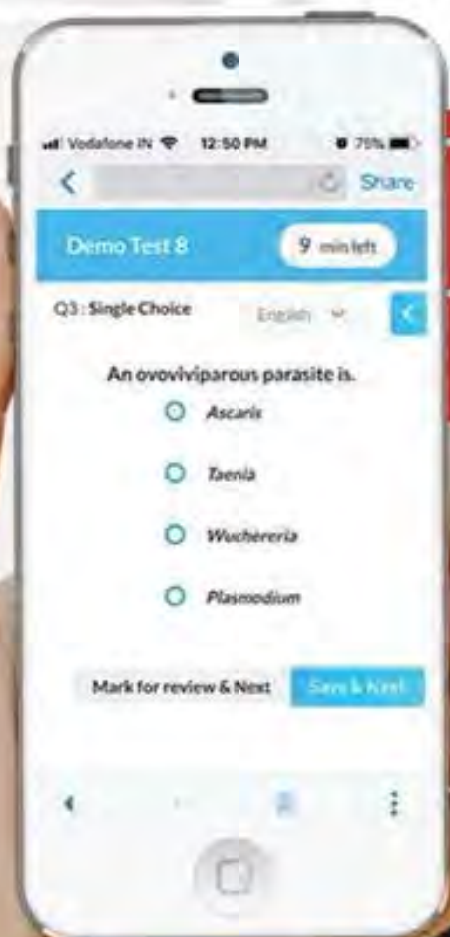
76. (b) $\Delta n_g = 2 - 4 = -2, \Delta H = \Delta E - 2RT$

77. (d) If $n_p < n_r; \Delta n_g = n_p - n_r = -ve$.
Hence, $\Delta H < \Delta E$.

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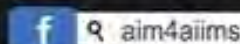
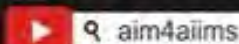
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