# DPP - Daily Practice Problems

**Name :**

**Date :**

**Start Time :**

**End Time :**

# CHEMISTRY 11

**SYLLABUS :** Thermodynamics-2 (Laws of Thermochemistry, Entropy, Spontaneous Process)

**Max. Marks : 120** **Time : 60 min.**

## GENERAL INSTRUCTIONS

- The Daily Practice Problem Sheet contains 30 MCQ's. For each question only one option is correct. Darken the correct circle/ bubble in the Response Grid provided on each page.
- You have to evaluate your Response Grids yourself with the help of solution booklet.
- Each correct answer will get you 4 marks and 1 mark shall be deduced for each incorrect answer. No mark will be given/ deducted if no bubble is filled. Keep a timer in front of you and stop immediately at the end of 60 min.
- The sheet follows a particular syllabus. Do not attempt the sheet before you have completed your preparation for that syllabus. Refer syllabus sheet in the starting of the book for the syllabus of all the DPP sheets.
- After completing the sheet check your answers with the solution booklet and complete the Result Grid. Finally spend time to analyse your performance and revise the areas which emerge out as weak in your evaluation.

**DIRECTIONS (Q.1-Q.21) : There are 21 multiple choice questions. Each question has 4 choices (a), (b), (c) and (d), out of which ONLY ONE choice is correct.**

**Q.1** Calculate the enthalpy change ($\Delta H$) for the following reaction

$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(g)$

Given average bond energies of various bonds

i.e. $C-H$, $C \equiv C$, $O=O$, $C=O$, $O-H$ as 414, 810, 499, 724 and 460 kJ/mole respectively

(a) – 1861 kJ (b) – 1361 kJ

(c) – 1261 kJ (d) – 1161 kJ

**Q.2** The gaseous endothermic reaction : $P + Q \rightarrow 3R$ at 327°C takes place spontaneously, because-

(a) $\Delta H < 0; \Delta S < 0$ (b) $\Delta H > 0; \Delta S > 0$

(c) $\Delta H < 0; \Delta S > 0$ (d) $\Delta H > 0; \Delta S < 0$

**Q.3** Calculate the standard free energy change for the reaction

$N_2(g) + 3H_2(g) \rightarrow 2NH_3$ at 298K

Given $\Delta H^\circ = -92.4$ kJ and $\Delta S^\circ = -198.3\ JK^{-1}$

(a) – 23.106 kJ (b) – 43.506 kJ

(c) – 33.306 kJ (d) – 13.306 kJ

**Q.4** Calculate the entropy change in melting 1 mole of ice at 273K, $\Delta H_f^\circ = 6.025$ kJ/mole-

(a) 11.2 $JK^{-1}\ mol^{-1}$ (b) 22.1 $JK^{-1}\ mol^{-1}$

(c) 15.1 $JK^{-1}\ mol^{-1}$ (d) 5.1 $JK^{-1}\ mol^{-1}$

**Q.5** Determine the standard free energy change for the following reaction at 298 K

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

Given $\Delta G^\circ_f$ for $N_2(g)$, $H_2(g)$ and $NH_3(g)$ are 0, 0 and $-16.66\ kJ\ mol^{-1}$

(a) –11.22 kJ (b) –22.22 kJ

(c) –33.32 kJ (d) –44.44 kJ

**RESPONSE GRID** 1. ⓐⓑⓒⓓ 2. ⓐⓑⓒⓓ 3. ⓐⓑⓒⓓ 4. ⓐⓑⓒⓓ 5. ⓐⓑⓒⓓ

*Space for Rough Work*

**Q.6** Calculate the standard enthalpy change for a reaction

$$CO_2(g) + H_2(g) \rightarrow CO(g) + H_2O\,(g)$$

given that $\Delta H_f^0$ for $CO_2(g)$, $CO(g)$ and $H_2O(g)$ as $-393.5$, $-110.5$ and $-241.8$ kJ/mol respectively.

(a) $-31.2$kJ (b) $-21.2$ kJ (c) $-11.2$kJ (d) $+41.2$kJ

**Q.7** Determine the enthalpy change for the reaction

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

Bond energies for C–H, C – Cl, Cl – Cl, H–Cl are 412, 338, 242, 431 kJ/mol. respectively.

(a) $-215$ kJ/mol (b) $-115$ kJ/ mol
(c) $-130$ kJ/mol (d) $-145$ kJ/mol

**Q.8** Calculate the heat change in the reaction

$$4NH_3\,(g) + 3O_2\,(g) \rightarrow 2N_2(g) + 6H_2O(\ell)$$

at 298 K, given that the heats of formation at 298 K for $NH_3$ (g) and $H_2O\,(\ell)$ are $-46.0$ and $-286.0$ kJ mol$^{-1}$ respectively.

(a) $-1861$ kJ (b) $-1361$ kJ (c) $-1261$ kJ (d) $-1532$kJ

**Q.9** The heat of hydration of solutions of anhydrous $CuSO_4$ and hydrated $CuSO_4.\,5H_2O$ are $-66.5$ and $11.7$ kJ mol$^{-1}$ respectively. Calculate the heat of hydration of $CuSO_4$ to $CuSO_4.\,5H_2O$.

(a) $-78.2$kJ (b) $-81.2$kJ (c) $-68.2$kJ (d) $-60.1$kJ

**Q.10** Calculate the entropy change ($\Delta S$) per mole for the following reactions:

(i) Combustion of hydrogen in a fuel cell at 298 K

$H_2(g) + ½\,O_2(g) \rightarrow H_2O\,(g)$

$\Delta H = -241.6$ kJ, $\Delta G = -228.4$ kJ

(ii) Vaporisation of methanol at its normal boiling point.

$CH_3OH\,(\ell) \rightarrow CH_3OH\,(g)$

$\Delta H_{vap} = 23.9$ kJ, boiling point $= 338$K

(iii) Fusion of ice at its normal melting point,

Ice → Water

$\Delta H_{fusion} = 6.025$ kJ, melting point $= 0°C$

(a) $-44.3$ JK$^{-1}$, 70 JK$^{-1}$, 22 JK$^{-1}$
(b) $-24.3$ JK$^{-1}$, 60 JK$^{-1}$, 11 JK$^{-1}$
(c) $-14.3$ JK$^{-1}$, 50 JK$^{-1}$, 32 JK$^{-1}$
(d) $-54.3$ JK$^{-1}$, 40 JK$^{-1}$, 21 JK$^{-1}$

**Q.11** One mole of an ideal diatomic gas ($C_v$ = 5 cal) was transformed from initial 25°C and 1 L to the state when temperature is 100°C and volume 10L. Then for this process (R = 2 calories/mol/K) (take calories as unit of energy and kelvin for temp.). Which of the following is incorrect?

(a) $\Delta H = 525$
(b) $\Delta S = 5 \ln \dfrac{373}{298} + 2 \ln 10$
(c) $\Delta E = 525$
(d) $\Delta G$ of the process cannot be calculated using given information

**Q.12** $C_{(graphite)} + O_2(g) \rightarrow CO_2(g); \Delta H = -94.05$ kcal mol$^{-1}$

$C_{(diamond)} + O_2(g) \rightarrow CO_2(g);\ \Delta H = -94.50$ kcal mol$^{-1}$

therefore,

(a) $C_{(graphite)} \rightarrow C_{(diamond)};\ \Delta H^\circ_{298K} = -450$ cal mol$^{-1}$
(b) $C_{(diamond)} \rightarrow C_{(graphite)};\ \Delta H^\circ_{298K} = +450$ cal mol$^{-1}$
(c) Graphite is the stable allotrope
(d) Diamond is harder than graphite

**Q.13** Consider the reactions

$C(s) + 2H_2(g) \rightarrow CH_4(g), \Delta H = -x$ kcal

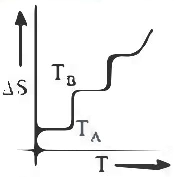
$C(g) + 4H(g) \rightarrow CH_4(g), \Delta H = -x_1$ kcal

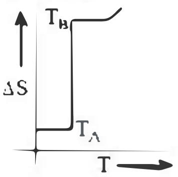
$CH_4\,(g) \rightarrow CH_3(g) + H(g), \Delta H = +y$ kcal

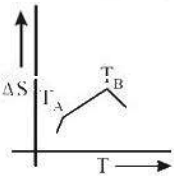
The bond energy of C – H bond is

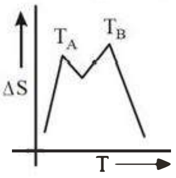
(a) y kcal mol$^{-1}$ (b) $x_1$ kcal mol$^{-1}$
(c) $x/4$ kcal mol$^{-1}$ (d) $x_1/4$ kcal mol$^{-1}$

**Q.14** If for a given substance, melting point is $T_B$ and freezing point is $T_A$, then correct variation shown by graph between entropy change and temperature is

(a)


(b)


(c)


(d)


**RESPONSE GRID**

6. ⓐⓑⓒⓓ 7. ⓐⓑⓒⓓ 8. ⓐⓑⓒⓓ 9. ⓐⓑⓒⓓ 10. ⓐⓑⓒⓓ
11. ⓐⓑⓒⓓ 12. ⓐⓑⓒⓓ 13. ⓐⓑⓒⓓ 14. ⓐⓑⓒⓓ

**Q.15** When a solid melts reversibly

(a) H decreases (b) G increases

(c) E decreases (d) S increases

**Q.16** The enthalpies of formation of $Al_2O_3$ and $Cr_2O_3$ are $-1596$ kJ and $-1134$ kJ respectively, ΔH for the reaction

$2Al + Cr_2O_3 \rightarrow 2Cr + Al_2O_3$ is

(a) $-2730$ kJ (b) $-462$ kJ

(c) $-1365$ kJ (d) $+2730$ kJ

**Q.17** "The quantity of heat which must be supplied to decompose a compound into its element is equal to the heat evolved during the formation of that compound from the elements." This statement is known as

(a) Hess's law

(b) Joule's law

(c) Le-chatelier's principle

(d) Lavoisier and Laplace law

**Q.18** For the reaction

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l), \Delta H = -285.8 \text{ kJ mol}^{-1}$$

$$\Delta S = -0.163 \text{ kJ mol}^{-1}K^{-1}.$$

What is the value of free energy change at 27°C for the reaction ?

(a) $-236.9$ kJ $mol^{-1}$ (b) $-281.4$ kJ $mol^{-1}$

(c) $-334.7$ kJ $mol^{-1}$ (d) $+334.7$ kJ $mol^{-1}$

**Q.19** The direct conversion of A to B is difficult, hence it is carried out by the following shown path :

$$\begin{tikzcd} C \arrow[r] & D \arrow[d] \\ A \arrow[u] & B \end{tikzcd}$$

Given

$\Delta S_{(A\rightarrow C)} = 50$ e.u. , $\Delta S_{(C\rightarrow D)} = 30$ e.u., $\Delta S_{(B\rightarrow D)} = 20$ e.u.,

where e.u. is the entropy unit, then $\Delta S_{(A\rightarrow B)}$ is

(a) +60e.u (b) +100 e.u (c) -60e.u. (d) $-100$ e.u.

**Q.20** The $H_2O(g)$ molecule dissociates as

(i) $H_2O(g) \rightarrow H(g) + OH(g); \Delta H = 490$ kJ

(ii) $OH(g) \rightarrow H(g) + O(g); \Delta H = 424$ kJ

The average bond energy (in kJ) for water is

(a) 490 (b) 424 (c) 914 (d) 914/2

**Q.21** Following reaction occurring in an automobile

$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$. The sign of ΔH, ΔS and ΔG would be

(a) +, –, + (b) –, +, – (c) –, +, + (d) +, +, –

**DIRECTIONS (Q.22-Q.24) : In the following questions, more than one of the answers given are correct. Select the correct answers and mark it according to the following codes:**

**Codes :**

**(a)** 1, 2 and 3 are correct **(b)** 1 and 2 are correct

**(c)** 2 and 4 are correct **(d)** 1 and 3 are correct

**Q.22** In which of the following entropy increases ?

(1) Rusting of iron

(2) Melting of ice

(3) Vaporization of camphor

(4) Crystallization of sucrose from solution

**Q.23** For a reaction $\Delta H = 9.08 \text{kJ mol}^{-1}$ and $\Delta S = 35.7 \text{ JK}^{-1} \text{ mol}^{-1}$ which of the following statements is incorrect for the reaction?

(1) Reversible and isothermal

(2) Reversible and exothermic

(3) Spontaneous and exothermic

(4) Spontaneous and endothermic

**Q.24** Which of the following are correct statements ?

(1) When ΔG is negative, the process is spontaneous

(2) When ΔG is zero, the process is in a state of equilibrium

(3) When ΔG is positive, the process is non-spontaneous

(4) When ΔG is positive, the process is spontaneous

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| **RESPONSE GRID** | **15.** ⓐⓑⓒⓓ | **16.** ⓐⓑⓒⓓ | **17.** ⓐⓑⓒⓓ | **18.** ⓐⓑⓒⓓ | **19.** ⓐⓑⓒⓓ |
| | **20.** ⓐⓑⓒⓓ | **21.** ⓐⓑⓒⓓ | **22.** ⓐⓑⓒⓓ | **23.** ⓐⓑⓒⓓ | **24.** ⓐⓑⓒⓓ |

**DIRECTIONS (Q.25-Q.27) : Read the passage given below and answer the questions that follows :**

J.W. Gibbs and H.Von Helmholtz had given two equations which are known as Gibbs-Helmholtz equation. One equation can be expressed in terms of change in free energy (ΔG) and enthalpy (ΔH) while other can be expressed in terms of change in internal energy (ΔE) and work function (ΔW)

$$\Delta G = \Delta H + T\,.\,d\left(\frac{\Delta G}{dT}\right)_P \qquad \text{............ (1)}$$

$$\Delta W = \Delta E + T\,.\,d\left(\frac{\Delta W}{dt}\right)_V \qquad \text{............. (2)}$$

Where T is temperature

Equation (1) is obtained at constant pressure while equation (2) is obtained at constant volume. It is observed that for the reaction

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

Free energy change at 25°C is –33 kJ while at 35°C it is –28 kJ which are at a constant pressure.

**Q.25** What would be the difference between enthalpy change at 25°C and 35°C for a given reaction ?

(a) 4 kJ (b) 5 kJ (c) 3 kJ (d) zero

**Q.26** What would be the free energy change at 30°C ?

(a) 30.5 kJ (b) 33 kJ
(c) –28 kJ (d) –30.5 kJ

**Q.27** Internal energy change at 25°C is $\Delta E_1$ while at 35°C it is $\Delta E_2$ then –

(a) $\Delta E_1 = \Delta E_2$ (b) $\Delta E_2 > \Delta E_1$
(c) $\Delta E_1 > \Delta E_2$ (d) None of these

**DIRECTIONS (Q. 28-Q.30) : Each of these questions contains two statements: Statement-1 (Assertion)and Statement-2 (Reason). Each of these questions has four alternative choices, only one of which is the correct answer. You have to select the correct choice.**

(a) Statement-1 is True, Statement-2 is True; Statement-2 is a correct explanation for Statement-1.
(b) Statement-1 is True, Statement-2 is True; Statement-2 is NOT a correct explanation for Statement-1.
(c) Statement -1 is False, Statement-2 is True.
(d) Statement -1 is True, Statement-2 is False.

**Q.28 Statement 1 :** Entropy of ice is less than water.
**Statement 2 :** Ice has cage like structure.

**Q.29 Statement 1 :** There is no reaction known for which ΔG is positive, yet it is spontaneous.
**Statement 2 :** For photochemical reactions ΔG is positive.

**Q.30 Statement 1 :** Enthalpy and entropy of any elementary substance in the standard state are taken as zero.
**Statement 2 :** Even at zero degree absolute, the constituent particles are not completely motionless.

| | |
|---|---|
| **RESPONSE GRID** | **25.** ⓐⓑⓒⓓ **26.** ⓐⓑⓒⓓ **27.** ⓐⓑⓒⓓ **28.** ⓐⓑⓒⓓ **29.** ⓐⓑⓒⓓ<br>**30.** ⓐⓑⓒⓓ |

| DAILY PRACTICE PROBLEM SHEET 11 - CHEMISTRY | | | |
|---|---|---|---|
| Total Questions | 30 | Total Marks | 120 |
| Attempted | | Correct | |
| Incorrect | | Net Score | |
| Cut-off Score | 32 | Qualifying Score | 52 |
| Success Gap = Net Score – Qualifying Score | | | |
| Net Score = (Correct × 4) – (Incorrect × 1) | | | |

***Space for Rough Work***

# DAILY PRACTICE PROBLEMS

# CHEMISTRY SOLUTIONS

# (11)

**(1)** (a) $2[H-C\equiv C-H]+5(O=O)$
$\rightarrow 4(O=C=O)+2(H-O-H)$

The reaction involves the breaking of 4 C – H bonds, 2C ≡ C bonds, 5 O = O bonds and formation of 8 C = O bonds and 4 O – H bonds.

$\Delta H$ = [Total energy required to break bonds] – [Energy given out in forming the bonds]

$= [4\Delta H_{C-H} + 2\,\Delta H_{C\equiv C} + 5\,\Delta H_{O=O}] - [8\,\Delta H_{C=O} + 4\,\Delta H_{O-H}]$

$=(4\times 414+2\times 810+5\times 499)-(8\times 724+4\times 460)$

$=5771-7632=-1861$ kJ

**(2)** (b) $\Delta G = \Delta H - T\Delta S$

For an endothermic reaction $\Delta H = +$ ve

[less molecules → more molecules],

$\Delta S = +$ ve

$\therefore T\Delta S > \Delta H$, and $\Delta G = -$ ve

Hence, the reaction occurs spontaneously.

**(3)** (c) $\Delta G^{\bullet} = \Delta H^{\circ} - T\Delta S^{\circ}$

$=-92.4-298\times(-198.3)\times 10^{-3}=-33.306$ kJ.

**(4)** (b) $\Delta S_f = \dfrac{\Delta H_f}{T} = \dfrac{6025\text{ J mol}^{-1}}{273\text{K}} = 22.1\text{ JK}^{-1}\text{ mol}^{-1}$

**(5)** (c) $\Delta G^{\circ} = 2\times\Delta G^{\circ}{}_{NH_3} - \Delta G^{\circ}{}_{N_2} - (\Delta G^{\circ}{}_{H_2}\times 3)$

$= 2\times(-16.66)-0-0=-33.32$ kJ

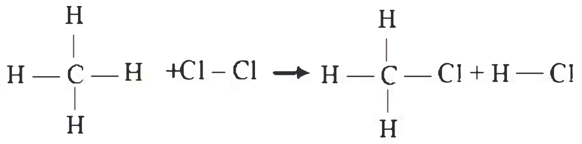
**(6)** (d) $\Delta H^{\circ} = \Sigma\Delta H^{\bullet}_{f\,(products)} - \Sigma\Delta H^{\circ}_{f\,(reactants)}$

$= [\Delta H^{\bullet}_{f(H_2O)} + \Delta H^{\bullet}_{f(CO)}] - [\Delta H^{\bullet}_{f(CO_2)} + \Delta H^{\bullet}_{f(H_2)}]$

$\Delta H^{\circ} = [-241.8-110.5]-[-393.5+0]$

$=-352.3+393.5 \quad =+41.2$ kJ

**(7)** (b)



$$H-\underset{H}{\overset{H}{C}}-H + Cl-Cl \longrightarrow H-\underset{H}{\overset{H}{C}}-Cl + H-Cl$$

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

$= [4(B.E.)_{C-H} + 1(B.E.)_{Cl-Cl} - \{3(B.E.)_{C-H} + 1(B.E.)_{C-Cl} + 1(B.E.)_{H-Cl}\}]$

$=412+242-338-431=-115$ kJ/mol

**(8)** (d) $\Delta H^{\circ}$ for the reaction

$4NH_3(g)+3O_2(g)\rightarrow 6H_2O(\ell)+2N_2(g)$

$\Delta H^{\circ} = \Delta H^{\circ}_f(\text{products}) - \Delta H^{\circ}_f(\text{reactants})$

$= \{6H^{\bullet}_f[H_2O(\ell)] + 2\Delta H^{\circ}_f[N_2(g)]\} - \{4\Delta H[NH_3(g)] + 3\Delta H_f[O_2(g)]\}$

$\Delta H^{\circ}_f[H_2O(\ell)] = -286.0\text{ kJ mol}^{-1}$

$\Delta H^{\circ}_f[NH_3(g)] = 0$ and $\Delta H^{\circ}_f[N_2(g)] = 0$ (by convention)

$\Delta H^{\bullet} = \{6(-286)+2(0)\} - \{4(-46.0)+3(0)\}$

$=-1716+184=-1532$ kJ

**(9)** (a) The required equation is :

$CuSO_4(s) + 5H_2O(\ell) \rightarrow CuSO_4.5H_2O(s)\ \Delta H = ?$

Given that

$CuSO_4(s) + (aq.) \rightarrow CuSO_4.(aq)\,;\ \Delta H_1 = -66.5$ kJ

$CuSO_4.5H_2O(s) + aq. \rightarrow CuSO_4.(aq)\,;\ \Delta H_2 = +11.7$ kJ

The process of hydration may be expressed as

$CuSO_4(s) \xrightarrow{\Delta H_1} CuSO_4(aq)$

$CuSO_4.5H_2O(s) \xrightarrow{\Delta H_2} CuSO_4(aq.)$

According to Hess's law

$\Delta H_1 = \Delta H + \Delta H_2$

$\Delta H = \Delta H_1 - \Delta H_2$

$=-66.5-11.7=-78.2$ kJ

**(10)** (a) (i) According to the reaction

$\Delta G = \Delta H - T\Delta S$ of $\Delta S = \dfrac{\Delta H - \Delta G}{T}$

Given that $\Delta G = -228.4$ kJ, $\Delta H = -241.6$ kJ and $T = 298$ K

Substituting the values, we get

$\Delta S = \dfrac{(-241.6\text{kJ})-(-228.4\text{kJ})}{298\text{K}}$

$=-0.0443\text{ kJK}^{-1} = -44.3\text{ JK}^{-1}$

(ii) Vaporisation represents a state of equilibrium between the liquid and the vapour state of the substance. At this state, $\Delta G = 0$ so

$\Delta H = T\Delta S$ or $\Delta S = \dfrac{\Delta H}{T}$

Here $\Delta H = 23.9$ kJ. $T = 338$ K

$\therefore\ \Delta S = \dfrac{23.9\text{kJ}}{338\text{K}} = 0.070\text{ kJK}^{-1} = 70\text{ JK}^{-1}$

(iii) Fusion represents a state of equilibrium between ice and water. At this state, $\Delta G = 0$ so

$\Delta H = T\Delta S$ or $\quad \Delta S = \dfrac{\Delta H}{T}$

Here $\Delta H = 6.025$ kJ and $T = 273$ K

$\Delta S = \dfrac{6.025\text{kJ}}{273\text{K}} = 0.0220\text{ kJK}^{-1} = 22\text{ JK}^{-1}$

**(11)** (d) $\Delta S = nC_v \ln\left(\dfrac{T_f}{T_i}\right) + nR\ln\left(\dfrac{V_f}{V_i}\right) = 5\ln\dfrac{373}{298} + 2\ln 10$

For 1 mol, $C_p - C_v = R \Rightarrow C_p = 2+5 = 7$ Cal

$\Delta H = nC_p\Delta T = 1\times 7\times 75 = 525$

$\Delta E = nC_v\Delta T = 1\times 5\times 75 = 375$

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

So, we cannot calculate $\Delta G$ from the given data.

**(12)** **(c)** Heat energy is also involved when one allotropic form of an element is converted into another. Graphite is the stabler allotrope because of heat of transformation of $C_{(diamond)} \rightarrow C_{(graphite)}$.

(i) $C_{(dia)} + O_{2(g)} = CO_{2(g)}, \Delta H = -94.5\ kcal$

(ii) $C_{(graphite)} + O_{2(g)} = CO_{2(g)}, \Delta H = -94.0\ kcal$

$\Delta H_{transformation} = -94.5 - (-94.0) = -0.5\ kcal.$

**(13)** **(a)** The bond energy of C – H bond will be the energy required to break one C – H in $CH_4$ i.e.

$CH_4(g) \rightarrow CH_3(g) + H(g), \Delta H = +y\ kcal\ mol^{-1}$

**(14)** **(a)** For a pure substance $T_A$ and $T_B$ represent the same temperature. Hence (a) is a correct choice.

**(15)** **(d)** When a solid melts, it changes into liquid and hence, entropy (s) increases.

**(16)** **(b)** $2Al + \frac{3}{2}O_2 \rightarrow Al_2O_3, \Delta H = -1596\ kJ$ ...(i)

$2Cr + \frac{3}{2}O_2 \rightarrow Cr_2O_3, \Delta H = -1134\ kJ$ ...(ii)

By (i) - (ii)

$2Al + Cr_2O_3 \rightarrow 2Cr + Al_2O_3, \Delta H = -462\ kJ.$

**(17)** **(d)** It is Lavoisier and Laplace law. This is another law of thermochemistry which was put forward before Hess's law.

**(18)** **(a)** $\Delta G = \Delta H - T\Delta S,\ T = 27 + 273 = 300\ K$

$\Delta G = (-285.8) - (300)(-0.163) = -236.9\ kJ\ mol^{-1}$

**(19)** **(a)** $\Delta S_{(A \rightarrow B)} = \Delta S_{(A \rightarrow C)} + \Delta S_{(C \rightarrow D)} - \Delta S_{(B \rightarrow D)}$

$= 50 + 30 - 20 = 60$ e.u.

**(20)** **(d)** eq. (i) + eq. (ii) find the required result and divide by 2.

**(21)** **(b)** For combustion reaction, $\Delta H$ is negative, $\Delta n = (16 + 18) - (25 + 2) = +7$, so $\Delta S$ is +ve. Reaction is spontaneous, hence $\Delta G$ is –ve.

**(22)** **(a)** On crystallization, randomness of the system decreases and hence, entropy increases.

**(23)** **(a)** $\Delta H$ = +ve and $\Delta S$ = +ve (given)

$\therefore$ Process will be spontaneous if T is high.

Also, as $\Delta H$ = +ve, thus, it is an endothermic process.

**(24)** **(a)**

**(25)** **(d)** $\Delta G_1 = \Delta H_1 + T_1\, d\left(\frac{\Delta G}{dT}\right)_P$

$$\Delta G_2 = \Delta H_2 + T_2\, d\left(\frac{\Delta G}{dT}\right)_P$$

$$\Delta G_2 - \Delta G_1 = (\Delta H_2 - \Delta H_1) + (T_2 - T_1)\left(\frac{\Delta G}{dT}\right)_P$$

$$(\Delta H_2 - \Delta H_1) = (\Delta G_2 - \Delta G_1) - (T_1 - T_2)\left[\frac{\Delta G_2 - \Delta G_1}{T_2 - T_1}\right] = 0$$

**(26)** **(d)** $\Delta S = -d\left(\frac{\Delta G}{dT}\right)_P = -\left[\frac{\Delta G_2 - \Delta G_1}{T_2 - T_1}\right]$

$$= -\left[\frac{-28 - (-33)}{308 - 298}\right] = -0.5\ kJ/K$$

$$\Delta G_1 = \Delta H_1 + T_1\, d\left(\frac{\Delta G}{dT}\right)_P$$

$-33 = \Delta H_1 + 298 \times 0.5$

$\Delta H_1 = -33 - 149 = -182$ kJ/mole

Now $\Delta G = \Delta H + T.d\left(\frac{\Delta G}{dt}\right)_P$

$\Delta G = -182 + 303 \times 0.5 = -30.5$ kJ

This is average of $\Delta G$ at 25°C and 35°C

**(27)** **(c)** $\Delta H_1 = \Delta E_1 + \Delta n_g RT$

$\Delta H_2 = \Delta E_2 + \Delta n_g RT_2$

$\Delta H_2 - \Delta H_1 = 0$ (as calculated in Ans. 25)

So, $\Delta H_2 = \Delta H_1$

So, $\Delta E_1 + \Delta n_g RT_1 = \Delta E_2 + \Delta n_g RT_2$

$(\Delta E_1 - \Delta E_2) = \Delta n_g R(T_2 - T_1)$

But $\Delta E_1 - \Delta E_2 > 0$

So, $\Delta E_1 > \Delta E_2$

**(28)** **(b)** Entropy of ice is less than water because water molecules in solid state lose kinetic energy and hence their tendency of movement minimises. Hence entropy decreases in solid state. The statement that ice has cage like structure is also correct but reason is not correct explanation for assertion.

**(29)** **(c)** Photochemical reactions may have +ve $\Delta G$.

**(30)** **(c)** Enthalpy is zero but entropy is not zero. Vibrational motion exists even at absolute zero.