

THERMODYNAMICS & THERMOCHEMISTRY

I -CHEMISTRY

- C_v values for monoatomic and diatomic gases respectively are

a) $\frac{1}{2}R, \frac{3}{2}R$ b) $\frac{3}{2}R, \frac{5}{2}R$ c) $\frac{5}{2}R, \frac{7}{2}R$ d) $\frac{3}{2}R, \frac{3}{2}R$
- Which of the following relationship is correct?

a) $\Delta G^\circ = -RT \ln K$ b) $K = e^{-\Delta G^\circ/RT}$ c) $K = 10^{-\Delta G^\circ/2.303 RT}$ d) All are correct
- ΔH and ΔS for the reaction
 $\text{Ag}_2\text{O}_{(s)} \rightarrow 2\text{Ag}_{(s)} + 1/2 \text{O}_{2(g)}$ are $30.56 \text{ kJ mol}^{-1}$ and $66.00 \text{ JK}^{-1} \text{ mol}^{-1}$ respectively. The temperature at which the free energy change for the reaction will be zero is

a) 463 K b) 35440 K c) 20 K d) 483 K
- The free energy and entropy change in kJ per mole when liquid water boils at 1 atmosphere are respectively (latent heat of water = 2.0723 kJ g^{-1})

a) 0, 0 b) 0.1, 0.1 c) 0.1, 0 d) 0, 0.1
- What entropy change is involved in the isothermal expansion of 5 mol of ideal gas from a volume of 10 L to 100 L at 300 K?

a) -95.7 J K^{-1} b) $+95.7 \text{ J K}^{-1}$ c) 28.72 kJ K^{-1} d) -28.72 kJ K^{-1}
- A plot of $\ln K$ against $1/T$ (abscissa) is expected to be a straight line with intercept on ordinate axis equal to

a) $\frac{\Delta S^\circ}{2.303R}$ b) $\frac{\Delta S^\circ}{R}$ c) $-\frac{\Delta S^\circ}{R}$ d) $R \times \Delta S^\circ$
- The standard Gibbs free energy change (ΔG°) at 25°C for the dissociation of $\text{N}_2\text{O}_{4(g)}$ to $\text{NO}_{2(g)}$ is (given, equilibrium cont. = 0.15, $R = 8.314 \text{ J K/mol}$)

a) 1.1 kJ b) 4.7 kJ c) 8.1 kJ d) 38.2 kJ
- Calculate the change in entropy when two gram atoms of aluminium is heated from 25°C to 600°C . The atomic heat capacity of aluminium is given by $C_p = 5.006 + 0.0026 T \text{ cal/deg}$.

a) 23.75 cal/deg b) 6.8 eu c) 13.75 eu d) 10 eu
- Which of the following is not correct for ideal gas?

a) $\left(\frac{\partial U}{\partial V}\right)_T = 0$ b) $\left(\frac{\partial H}{\partial P}\right)_T = 0$ c) $\left(\frac{\partial T}{\partial P}\right)_H = 0$ d) $\left(\frac{\partial P}{\partial T}\right)_V = 0$
- Consider the reaction $2\text{NO}_{(g)} + \text{O}_{2(g)} \rightarrow 2\text{NO}_{2(g)}$
Predict whether the reaction is spontaneous at 298 K.
 $\Delta_f G^\circ(\text{NO}) = 86.69 \text{ kJ/mol}$, $\Delta_f G^\circ(\text{NO}_2) = 51.84 \text{ kJ/mol}$

a) Spontaneous b) Non-spontaneous c) Equilibrium d) Cannot predict

11. The values of ΔH and ΔS for a reaction are respectively 30 kJ mol^{-1} and $100 \text{ JK}^{-1} \text{ mol}^{-1}$. Then the temperature above which the reaction will become spontaneous is
 a) 300 K b) 30 K c) 100 K d) 300°C
12. For the reaction, $\text{CaCO}_{3(s)} = \text{CaO}_{(s)} + \text{CO}_{2(g)}$
 Partial pressure of CO_2 at 1000 K is 0.003 atm.
 $\Delta G^\circ = 27.2 \text{ kcal}$. Calculate the value of ΔG .
 a) 12.6 kcal b) 15.6 kcal c) 13.4 kcal d) 14.2 kcal
13. Given: $\text{C}_2\text{H}_{6(g)} \rightarrow 2\text{C}_{(g)} + 6\text{H}_{(g)}$; $\Delta H = 712 \text{ kcal}$
 The C – C bond energy is 112 kcal, what is the C – H bond energy?
 a) 88 kcal b) 12 kcal c) 50 kcal d) 100 kcal
14. For a particular reversible reaction at temperature T, ΔH and ΔS were found to be both positive, If T_e is the temperature at equilibrium, the reaction would be spontaneous when
 a) $T = T_e$ b) $T_e > T$ c) $T > T_e$ d) $T_e = 5T$
15. One mole of an ideal gas at 300 K is expanded Isothermally from an initial volume of 1 L to 10L. The ΔE for this process is $R = 2 \text{ cal mol}^{-1} \text{ K}^{-1}$
 a) 163.7 cal b) zero c) 1381.1 cal d) 9 L atm'
16. For the reaction $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, $\Delta H = -571$. Bond energy of H – H = 435 and O = O = 498. Then the average bond energy of O-H bond will be
 a) 484 b) -484 c) 271 d) -271
17. If one mole of monoatomic ideal gas is heated at constant pressure of 1 atm from 25°C , the change in internal energy in calories will be
 a) 124.2 cal b) -49.7 cal c) 74.5 cal d) 173.9 cal
18. For the allotropic change represented by the equation $\text{C (graphite)} \rightarrow \text{C (diamond)}$, the enthalpy change $\Delta H = 1.9 \text{ kJ}$. If 6 g of diamond and 6 g of graphite are separately burnt to yield CO_2 , the heat liberated in first case is
 a) less than in the second case by 1.9 kJ
 b) more than in the second case by 11.4 kJ
 c) more than in the second case by 0.95 kJ
 d) less than in the second case by 11.4 kJ
19. The decreasing order of calorific values is
 a) $\text{C}_2\text{H}_2 > \text{C}_4\text{H}_{10} > \text{C}_3\text{H}_8 > \text{C}_2\text{H}_4$ b) $\text{C}_4\text{H}_{10} > \text{C}_3\text{H}_8 > \text{C}_2\text{H}_4 > \text{C}_2\text{H}_2$
 c) $\text{C}_2\text{H}_4 > \text{C}_3\text{H}_8 > \text{C}_4\text{H}_{10} > \text{C}_2\text{H}_2$ d) $\text{C}_3\text{H}_8 > \text{C}_2\text{H}_4 > \text{C}_4\text{H}_{10} > \text{C}_2\text{H}_2$
20. Which of the following reaction is said to be enthalpy driven?
 a) Endothermic reaction with positive entropy change and high temperature
 b) Endothermic reaction with negative entropy change and low temperature
 c) Exothermic reaction with positive entropy change and high temperature
 d) Exothermic reaction with negative entropy change and low temperature

31. 12.0 g of carbon completely reacted with oxygen to form CO and CO₂ at 25°C and constant pressure. The heat liberated is 75.0 kcal. The mass of oxygen reacted is (Given $\Delta_f H^\circ$: CO₂ = -95 kcal mol⁻¹, CO = -24 kcal mol⁻¹)
- a) 13.75 g b) 27.5 g c) 33.3 g d) 41.25 g
32. 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
- a) 11.09×10^8 Pa b) 11.09×10^7 Pa c) 11.09×10^6 Pa d) 11.09×10^5 Pa
33. During an event, an athlete is given 100 g of glucose which has energy equivalent to 1560 kJ. He utilizes 50% of the gained energy in the event. In order to avoid storage of energy in the body, the weight of water he needs to perspire would be nearly (given enthalpy of vaporisation of water is 44 kJ/mol)
- a) 120 g b) 220 g c) 320 g d) 420 g
34. The enthalpy change involved in the oxidation of glucose is -2880 kJ mol⁻¹. 25% of this energy is available for muscular work. If 100 kJ of muscular work is needed to walk one kilometer, the maximum distance that a person will be able to walk after eating 100 g of glucose will be
- a) 2.9 km b) 3.8 km c) 4.8 km d) 5.8n km
35. One mole of a liquid (1 bar, 100 mL) is taken in an adiabatic container and the pressure increases steadily to 100 bar. Then at constant pressure of 100 bar, volume decreases by 1 mL. Enthalpy change during the process will be
- a) 980 J b) 990 J c) 970 J d) 950 J
36. For hypothetical reversible reaction
- $$\frac{1}{2}A_2(g) + \frac{3}{2}B_2(g) \longrightarrow AB_3(g); \quad \Delta H = -20 \text{ kJ}$$
- if standard entropies of A₂, B₂ and AB₃ are 60, 40 and 50 JK⁻¹ mol⁻¹ respectively. Then above reaction will be at equilibrium at
- a) 400 K b) 500 K c) 350 K d) 600 K
37. The enthalpy change (ΔH) for the reaction N₂ (g) + 3 H₂ (g) \longrightarrow 2 NH₃ (g) is 92.38 kJ at 298 K. The internal energy change ΔU at 298 K is
- a) -92.38 kJ b) -87.42 kJ c) -97.34 kJ d) -89.9 kJ
38. Sodium chloride is soluble in water but not in benzene because
- a) $\Delta H_{\text{hydration}} < \Delta H_{\text{lattice energy in water}}$ and $\Delta H_{\text{hydration}} > \Delta H_{\text{lattice energy in benzene}}$
- b) $\Delta H_{\text{hydration}} > \Delta H_{\text{lattice energy in water}}$ and $\Delta H_{\text{hydration}} < \Delta H_{\text{lattice energy in benzene}}$
- c) $\Delta H_{\text{hydration}} = \Delta H_{\text{lattice energy in water}}$ and $\Delta H_{\text{hydration}} < \Delta H_{\text{lattice energy in benzene}}$
- d) $\Delta H_{\text{hydration}} < \Delta H_{\text{lattice energy in water}}$ and $\Delta H_{\text{hydration}} = \Delta H_{\text{lattice energy in benzene}}$

39. Assuming that water vapour is an ideal gas, the internal energy change (ΔU) when 1 mole of water is vapourised at 1 bar pressure and 100°C (Given: molar enthalpy of vapourisation of water at bar and $373\text{ K} = 41\text{ kJ mol}^{-1}$ and $R = 8.3\text{ J mol}^{-1}$ will be
- a) 41.00 kJ mol^{-1} b) 4.100 kJ mol^{-1} c) $3.7908\text{ kJ mol}^{-1}$ d) $37.904\text{ kJ mol}^{-1}$
40. For the reaction $A \rightarrow B$, $\Delta H = +24\text{ kJ/mole}$
 For the reaction $B \rightarrow C$, $\Delta H = -18\text{ kJ/mole}$
 The decreasing order of enthalpy of A, B, C follows the order :
- a) A, B, C b) B, C, A c) C, B, A d) C, A, B
41. 1.0 g sample of substance A at 100°C is added to 100 mL of H_2O at 25°C . Using separate 100 mL portions of H_2O , the procedure is repeated with substance B and then with substance C. How will the final temperatures of the water compare?

Substance	Specific heat
A	$0.60\text{ J g}^{-1}\text{ }^\circ\text{C}^{-1}$
B	$0.40\text{ J g}^{-1}\text{ }^\circ\text{C}^{-1}$
C	$0.20\text{ J g}^{-1}\text{ }^\circ\text{C}^{-1}$

- a) $T_C > T_B > T_A$ b) $T_B > T_A > T_C$ c) $T_A > T_B > T_C$ d) $T_A = T_B = T_C$
42. A cylinder of a gas contains 11.6 kg of butane. If a normal family needs $2.0 \times 10^4\text{ kJ}$ of energy per day for cooking, how long will the cylinder last? (Given : $\Delta H_{\text{combustion}}$ of butane = -2600 kJ mol^{-1})
- a) 26 days b) 39 days c) 52 days d) 65 days
43. When 20 mL of a strong acid is added to 20 mL of an alkali, the temperature rises by 5°C . If 200 mL of each liquid are mixed, the temperature rise would be
- a) 5°C b) 50°C c) 20°C d) 0.5°C
44. One gram sample of NH_4NO_3 is decomposed in a bomb calorimeter. The temperature of the calorimeter increases by 6.12 K . The heat capacity of the system is 1.23 kJ/g/deg . What is the molar heat of decomposition for NH_4NO_3 ?
- a) -7.53 kJ/mol b) -398.1 kJ/mol c) -16.1 kJ/mol d) -602 kJ/mol
45. When 0.2 g of butanol -1 was burnt in a suitable apparatus, the heat evolved was sufficient to raise the temperature of 200 g of water by 5°C . The heat of combustion of butanol -1 in kcal/mole will be (mol mass of butanol -1 = 74)
- a) 14.8 b) 74 c) 37 d) 370
46. What amount of ice will remain when 52 g ice is added to 100 g of water at 313 K . Specific heat of water is 1 cal/g and latent heat of fusion of ice is 80 cal/g .
- a) 2 g b) 4 g c) 6 g d) 52 g

$$\Delta S = nR \ln \frac{V_2}{V_1} = 2.30 nR \log \frac{V_2}{V_1}$$

$$= 2.303 \times 5 \text{ mol} \times (8.314 \text{ JK}^{-1} \text{ mol}^{-1}) \log \frac{100\text{L}}{10\text{L}}$$

$$= 95.7 \text{ J K}^{-1}$$

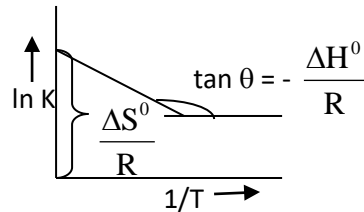
6. (b)

$$\Delta G^0 = -RT \ln K \text{ or } \Delta H^0 - T\Delta S^0 = -RT \ln K$$

$$\ln K = \frac{-\Delta H}{RT} + \frac{\Delta S^0}{R}$$

Comparing with $Y = mx + c$

$$\therefore \text{Y intercept is } \frac{\Delta S^0}{R}$$



7. (b)

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

8. (c)

The process takes place at constant pressure.

$$\text{Hence, } \Delta Q = C_p \Delta T$$

$$\text{and } \Delta S = \int_{T_1}^{T_2} \frac{nC_p dT}{T} = 2 \int_{298}^{873} \left(\frac{5.006}{T} + 0.0026 \right) dT$$

$$= 2 \left[5.006 \ln \frac{873}{298} + 0.0026(873 - 298) \right]$$

$$= 2 [5.381 + 1.495] = 13.75 \text{ cal/deg or } 13.75 \text{ eu}$$

9. (d)

The forces of attraction between the molecules of an ideal gas are negligible so

$$\left(\frac{\partial U}{\partial V} \right)_T = 0$$

$$\left(\frac{\partial H}{\partial P} \right)_T = 0 \text{ and } \left(\frac{\partial T}{\partial P} \right)_H = 0$$

10. (a)

For the reaction: $2\text{NO}_{(g)} + \text{O}_{2(g)} \rightarrow 2\text{NO}_{2(g)}$

$$\Delta G^0 = 2\Delta_f G^0 (\text{NO}_2) - [2\Delta_f G^0 (\text{NO}) + \Delta_f G^0 (\text{O}_2)]$$

$$= (2 \times 51.85) - (2 \times 86.69 + 0) = 103.68 - 173.38$$

$$= -69.7 \text{ kJ/mol; } \Delta G^0 \text{ is negative}$$

So, the reaction is spontaneous.

11. (a)

Given $\Delta H = 30 \text{ kJ mol}^{-1}$ or $30,000 \text{ J mol}^{-1}$

$$\Delta S = 100 \text{ JK}^{-1} \text{ mol}^{-1}$$

$$T = \frac{\Delta H}{\Delta S} = \frac{30,000}{100} = 300 \text{ K} \quad (\text{at equilibrium})$$

The reaction will be spontaneous above 300 K.

12. (b)



$$K_p = p_{\text{CO}_2} = 0.003 \text{ atm}$$

$$\begin{aligned}\Delta G &= \Delta G^0 + 2.303 RT \log K_p \\ &= 27200 + 2.303 \times 2 \times 1000 \times \log 0.003 \\ &= 15.6 \text{ kcal. (R = 2 cal mol}^{-1} \text{ K}^{-1})\end{aligned}$$

13. (d)

Bond energy of 6C – H bonds = 712 – 112 = 600 kcal

14. (c)

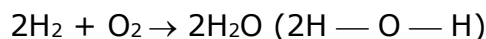
$\Delta G = \Delta H - T\Delta S$. At equilibrium, $\Delta G = 0$. Also, ΔG should be negative for a spontaneous reaction. So, $T > T_e$ in order to make ΔG negative, because both ΔH and ΔS are positive.

15. (b)

For a finite change $\Delta E = C_v\Delta T$. For an isothermal process $\Delta T = 0$. Hence, $\Delta E = 0$.

16. (a)

For the reaction



$$\Delta_r H = 2 \times \text{BE} (\text{H}_2) + \text{BE} (\text{O}_2) - 4 \times \text{BE} (\text{O} - \text{H})$$

$$-571 = 2 \times 435 + 498 - 4 \times \text{BE} (\text{O} - \text{H}) \text{ or } \text{BE} (\text{O} - \text{H}) = 484$$

17. (c)

For monoatomic ideal gas,

$$C_v = \frac{2}{3} R, \quad C_p = \frac{5}{2} R$$

$$\text{But } C_p = \frac{\Delta H}{\Delta T}$$

$$\therefore \Delta H = C_p \Delta T \text{ mol}^{-1} \text{ (or } nC_p \Delta T \text{ for } n \text{ moles)}$$

$$= \frac{5}{2} \times 1.987 \times 25 = 124.2 \text{ cal}$$

$$\text{Work done, } w = -P \Delta V = -P (V_2 - V_1)$$

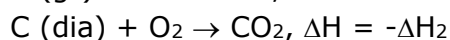
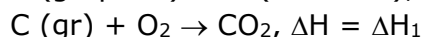
$$= -(PV_2 - PV_1) = -(n RT_2 - n RT_1)$$

$$= -n R (T_2 - T_1)$$

$$= -1 \times 1.987 (323 - 298) \text{ cal} = -49.7 \text{ cal}$$

$$\Delta U = q + w = 124.2 - 49.7 \text{ cal} = 74.5 \text{ cal}$$

18. (c)



$$(-\Delta H_1) - (-\Delta H_2) = 1.9 \text{ kJ or } \Delta H_2 - \Delta H_1 + 1.9 \text{ kJ}$$

For combustion of 6 g, $\Delta H_2 > \Delta H_1$ by

$$1.9/2 \text{ kJ} = 0.95 \text{ kJ}$$

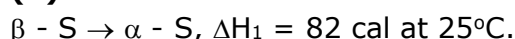
19. (b)

Greater the number of C and H atoms, greater is the heat of combustion per mole. Dividing by respective molecular weights, we get calorific values

20. (d)

For exothermic reaction, $\Delta H = -ve$. If ΔS is +ve, reaction is always spontaneous. However, if ΔS is -ve but T is so low that $T\Delta S < \Delta H$ or $\Delta H > T\Delta S$ in magnitude, process is spontaneous. This reaction is then said to be enthalpy driven.

21. (b)



Aim = To find ΔH_2 at 50°C

Applying Kirchoff's equation

$$\Delta H_2 - \Delta H_1 = \Delta C_p (T_2 - T_1)$$

$$\Delta C_p = C_p (\alpha - S) - C_p (\beta - S)$$

Converting specific heats into molar heat capacities

$$\Delta C_p = 0.171 \times 32 - 0.163 \times 32$$

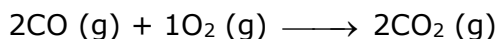
$$= 32 \times 0.008 \text{ cal mol}^{-1} = 0.256 \text{ cal mol}^{-1}$$

$$\therefore \Delta H_2 = \Delta H_1 + \Delta C_p (T_2 - T_1)$$

$$= 82 + 0.256 (25) = 82 + 6.4 \text{ cal mol}^{-1}$$

$$= 88.4 \text{ cal/g atom}$$

22. (b)



$$\text{Since } \Delta n_g = 2 - 3 = -2$$

Thus, $\Delta H < \Delta U$

23. (d)

$$\Delta H^\circ = \Delta H_f^\circ (\text{Al}_2\text{O}_3) - \Delta H_f^\circ (\text{Cr}_2\text{O}_3)$$

$$= (-1596) - (-1134) = -462 \text{ kJ}$$

24. (a)

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

$$= -2.303 \times 8.314 \times 300 \log 10$$

$$= -5744.1 \text{ J}$$

25. (c)

$$q = 1 \text{ kJ} = 1000 \text{ J}$$

$$m = 100 \text{ g} = \frac{100 \text{ g}}{18 \text{ g mol}^{-1}} = 5.55 \text{ mol}$$

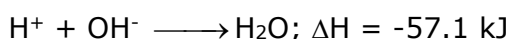
$$q = mC\Delta t$$

$$\Delta t = \frac{q}{m.C} = \frac{1000 \text{ J}}{5.55 \text{ mol} \times 75 \text{ JK}^{-1} \text{ mol}^{-1}}$$

$$= 2.4 \text{ K}$$

26. (b)

$$0.5 \text{ mol HNO}_3 \equiv 0.5 \text{ mol H}^+$$



$$\text{Here OH}^- = 0.3 \text{ mol}$$

Thus, OH^- is the limiting reactant.

When 1 mol of OH^- ions are neutralised, heat released = 57.1 kJ

When 0.3 mol of OH^- ions are neutralised, heat released = $57.1 \times 0.3 = 17.13 \text{ kJ}$

27. (c)

$$\Delta H = \Delta U + w_{\text{exp}}$$

When $\Delta U = 0$, $\Delta H = w_{\text{exp}}$

For free expansion, $w_{\text{exp}} = 0$

$$\therefore \Delta H = 0$$

28. (b)

$A \rightarrow B$, $\Delta H = +ve$ means that $H(A) < H(B)$

$B \rightarrow C$, $\Delta H = -ve$ means that $H(B) > H(C)$ or $H(C) < H(B)$

Adding the given equations, we get

$A \rightarrow C$, $\Delta H = +6 \text{ kJ mol}^{-1}$. This means that $H(A) < H(C)$

Combining the above results,

$H(A) < H(C) < H(B)$ or $H(B) > H(C) > H(A)$

Hence, decreasing order of enthalpy is B, C, A.

29. **(c)**
 Specific heat is the amount of heat required to raise the temperature of 1 g of the substance through 1°C or it is the heat released when temperature of 1 g of the substance falls through 1°C. Greater the specific heat of the substance, greater is the heat released which is absorbed by water and hence greater is the rise in temperature of water or greater is the temperature of water. As specific heats of A, B and C are in the order $A > B > C$, therefore, $T_A > T_B > T_C$.
30. **(b)**
 $C_{12}H_{22}O_{11} + 12O_2 \rightarrow 12CO_2 + 11H_2O, \Delta H = -5645 \text{ kJ}$
 O_2 used per day = 640 g = $\frac{640}{32}$ moles = 20 moles
 \therefore Heat evolved in one day = $\frac{5645}{12} \times 20$
 = 9408 kJ
31. **(b)**
 (i) $C + O_2 \rightarrow CO_2, \Delta H = -95 \text{ kcal}$
 (ii) $C + \frac{1}{2}O_2 \rightarrow CO, \Delta H = -24 \text{ kcal}$
 Total mass of carbon burnt = 12 g = 1 mole
 Suppose 'a' mole of C form CO_2 . Then (1 - a) mole will form CO.
 Heat liberated from 'a' mole of C forming CO_2 = 95 a
 Heat liberated from (1 - a) mole of C forming CO = 24 (1 - a)
 Total heat liberated = 95 a + 24 (1 - a) = 75 (Given)
 or $95 a + 24 - 24 a = 75$ or $71 a = 51$ or $a = \frac{51}{71}$
 O_2 consumed in reaction (i) = a mole
 O_2 consumed in reaction (ii) = $\frac{1}{2}$ (1 - a) mole
 Total O_2 consumed = $a + \frac{1}{2}(1 - a)$
 = $\frac{a}{2} + \frac{1}{2} = \frac{51}{142} + \frac{1}{2} = \frac{122}{142} \text{ mol}$
 = $\frac{122}{142} \times 32 \text{ g} = 27.5 \text{ g}$
32. **(a)**
 $-\Delta F = w_{\max} = p \Delta V$
 $1895 = p \left[\left(\frac{12}{2.25} - \frac{12}{3.31} \right) \times 10^{-6} \text{ m}^3 \right]$
 = $P (5.333 - 3.625) \times 10^{-6}$
 = $P \times 1.708 \times 10^{-6}$
 or $P = \frac{1895}{1.708 \times 10^{-6}} \text{ Pa} = 1109 \times 10^6$
 = $11.09 \times 10^8 \text{ Pa}$
33. **(c)**
 Energy left unutilized = $\frac{1560}{2} \text{ kJ} = 780 \text{ kJ}$
 For losing 44 kJ of energy, water to be evaporated = 1 mole = 18 g
 \therefore For losing 780 kJ of energy, water to be evaporated = $\frac{18}{44} \times 780 \text{ g} = 319 \text{ g}$
34. **(c)**

$$\text{Energy available for muscular work} = \frac{25}{100} \times 2880 \text{ kJ mol}^{-1} = 720 \text{ kJ mol}^{-1}$$

$$\text{Molar mass of glucose} = 180 \text{ g mol}^{-1}$$

$$\therefore \text{Energy available for muscular work from 120 g glucose} = \frac{720}{180} \times 120 \text{ kJ} = 480 \text{ kJ}$$

In 100 kJ of muscular work, distance walked = 1 km

$$\therefore \text{In 480 kJ of muscular work, distance walked} = \frac{1}{100} \times 480 = 4.8 \text{ km}$$

35. (b)

As the process is carried out under adiabatic conditions, $q = 0$. But $\Delta U = q + w$. Hence,
 $\Delta U = w$.

Further, at constant pressure of 100 bar, volume has decreased by 1 ml, therefore, work of contraction

$$= P \Delta V = 100 \text{ bar} \times 1 \text{ ml} = (100 \times 10^5 \text{ N m}^{-2}) (10^{-6} \text{ m}^3) = 10 \text{ J}$$

$$\text{Hence, } \Delta U = 10 \text{ J}$$

$$\Delta H = \Delta U + \Delta (PV)$$

$$\Delta (PV) = (P_2 V_2 - P_1 V_1)$$

$$\text{But } P_1 = 1 \text{ bar, } V_1 = 100 \text{ ml}$$

$$P_2 = 100 \text{ bar, } V_2 = 99 \text{ ml}$$

$$\therefore \Delta (PV) = 100 \text{ bar} \times 99 \text{ ml} - (1 \text{ bar} \times 100 \text{ ml})$$

$$= (9900 - 100) \text{ bar ml} = 9800 \text{ bar ml}$$

$$= 980 \text{ J}$$

$$\therefore \Delta H = 10 \text{ J} + 980 \text{ J} = 990 \text{ J}$$

$$(1 \text{ L atm} = 101.3 \text{ J}; 1 \text{ bar L} = 100 \text{ J}; 1 \text{ bar ml} = 0.1 \text{ J})$$

36. (b)

$$\Delta S = \Sigma S^\circ (p) - \Sigma S^\circ (r)$$

$$= [S^\circ (AB_3)] - \left[\frac{1}{2} S^\circ (A_2) + \frac{3}{2} S^\circ (B_2) \right] = 50 - \left(\frac{1}{2} \times 60 + \frac{3}{2} \times 40 \right) \text{ JK}^{-1} \text{ mol}^{-1}$$

$$\text{At equilibrium, } \Delta H = T \Delta S \therefore T = \frac{\Delta H}{\Delta S} = \frac{-20 \times 10^3 \text{ J}}{(50 - 1/2 \times 60 - 3/2 \times 40)}$$

$$T = 500 \text{ K}$$

37. (b)

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

For the reaction, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

$$\Delta H = 2 - 4 = -2$$

$$\Delta H = \Delta U + 8.314 \times 10^{-3} \times 298 \times (-2)$$

$$-92.38 = \Delta U - 4.955$$

$$-92.38 + 4.955 = \Delta U = 87.425 \text{ kJ}$$

38. (b) NaCl is soluble in H_2O but not in C_6H_6 due to

$\Delta H_{\text{hydration}} > \Delta H_{\text{lattice energy in water}}$ and $\Delta H_{\text{hydration}} < \Delta H_{\text{lattice energy in benzene}}$

39. (d)

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

$$41 = \Delta U + 8.314 \times 10^{-3} \times 373 \times 1$$

$$41 = \Delta U + 3.101 \text{ kJ}$$

$$\Delta U = 37.89 \text{ kJ}$$

40. (b)

In a reaction $A \rightarrow B$, $\Delta H = +ve$ means B has higher enthalpy (+24 kJ) than A.

In a reaction $B \rightarrow C$, $\Delta H = -ve$ means C is of lower enthalpy (18 kJ) enthalpy than B.

$\therefore A < C < B$ \therefore decreasing order is B, C, A.

41. (c)

Greater the specific heat, greater is the heat released when is absorbed by water

$$\therefore T_A > T_B > T_C$$

42. (c)

ΔH_c for combustion of 1 mole of butane (mol mass = 58) = -2600 kJ

$$\therefore \text{Energy of 11.6 kJ of butane} = \frac{11,600}{58} = 200 \text{ moles} = \frac{200 \times 2600}{5,20,000} \text{ kJ}$$

If the family needs $2 \times 10^4 \text{ kJ} = 20,000 \text{ kJ}$

$$\text{No. of days} = \frac{5,20,000}{20,000} = 26 \text{ days.}$$

43. (a)

$$Q_1 = m_1 s t \quad m = 20 \text{ g} + 20 \text{ g} = 40 \text{ g} = 0.04 \text{ kJ}$$

$Q_1 = 0.04 \times s \times t_1$ ----- (1) when 200 ml each are mixed $m = 0.4 \text{ kJ}$ and

$$Q_2 = 10 Q_1$$

$$Q_2 = 0.4 \times s \times t_2 \text{ \& } Q_2 = 10Q_1$$

$$\therefore 10 Q_1 = 0.4 s \times t_2 \text{ ----- (2) Dividing (2) by (1) } \frac{10}{1} = \frac{10t_2}{t_1} \Rightarrow t_1 = t_2$$

44. (d)

Heat evolved by 1g of $\text{NH}_4\text{NO}_3 = 1.23 \times 6.12 = 7.529 \text{ kJ}$

Heat evolved by 1mole of $\text{NH}_4\text{NO}_3 = 7.5296 \times 80 = 602 \text{ kJ}$

45. (d)

Heat evolved from 0.2 g = $200 \times 5 = 1000 \text{ cal} = 1 \text{ kcal}$

Heat evolved from 1 mole, i.e., 74 g = $\frac{1}{0.2} \times 74 = 370 \text{ kcal}$

46. (a) $Q = mst = 100 \times 1 \times (313 - 273)$

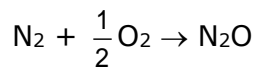
$$= 4000 \text{ cal}$$

Latent heat of fusion = 80 cal/g

$$\therefore \text{Mass of ice which can melt} = \frac{4000}{80} = 50 \text{ g}$$

$$\therefore \text{Amount of ice remaining} = 52 - 50 = 2 \text{ g}$$

47. (a)



$$H_R = \text{N} \equiv \text{N} + \frac{1}{2}\text{O} = \text{O} = 946 + \frac{1}{2} \times 498 = 946 + 249 = 1195 \text{ kJ}$$

$$H_P = \text{N} = \text{N} = \text{O} = 418 + 607 = 1025 \text{ kJ}$$

$$\Delta H = H_R - H_P = 1195 - 1025 = 170 \text{ kJ}$$

$$\text{The difference in } \Delta H_f = 170 - 82 = 88 \text{ kJ}$$

48. (a)

$$n \text{ for } 90 \text{ g of H}_2\text{O} = \frac{90}{18} = 5$$

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

$$\Delta H = \Delta U + 2 \times 373 \times 5$$

$$\Delta H - \Delta U = 3730 \text{ cal}$$

$$\Delta H = \frac{90 \times 540}{48600 \text{ cal}} \quad \therefore \Delta U = 48600 - 3730 = 44870 \text{ cal}$$

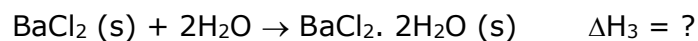
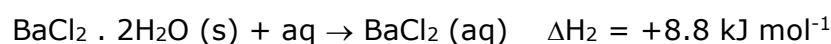
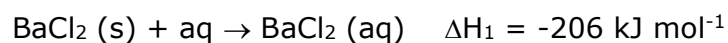
49. (b)

$$\text{Molar specific heat of monoatomic gas} = \frac{3}{2}R$$

$$\text{Molar specific heat of diatomic gas} = \frac{5}{2}R$$

$$\therefore \text{Molar specific heat of the mixture} = \frac{\frac{3}{2}R + \frac{5}{2}R}{2} = 2R = 2 \times 2 = 4 \text{ cal}$$

50. (d)



From Hess's law ,

$$\Delta H_1 = \Delta H_3 + \Delta H_2$$

$$\Delta H_1 - \Delta H_2 = \Delta H_3$$

$$-20.6 - 8.8 = -29.4 \text{ kJ}$$