

THERMODYNAMICS

1. For the following reaction $X + Y \rightleftharpoons A + B$ at 300 K, it is found equilibrium constant equal to 103. Therefore, ΔG & ΔG° of the reaction at 300 K respectively are -
- (a) Both are equal to zero (b) 17.23 KJ & zero
 (c) Zero and -17.23 KJ (d) Both are equal to -17.23 KJ

Answer: (c)

$$\Delta G^\circ = -RT \ln K_{eq}$$

$$\Delta G^\circ = -2.303 RT \log K_{eq}$$

$$\Delta G^\circ = -2.303 \times 8.314 \times 300 \times 10^{-3} \times 3$$

$$\Delta G^\circ = -17.23 \text{ KJ}$$

Since, the given reaction is at equilibrium hence $\Delta G = 0$.

2. The calorific value $H_2(g)$ at STP is 12.78 KJ/L hence approximate standard enthalpy of formation of $H_2O(l)$ is -
- (a) -143 KJ (b) -286 KJ (c) Zero (d) $+286$ KJ

Answer: (b)

$$1 \text{ L } H_2(g) \text{ at STP} = \frac{1}{22.4} \text{ mol,}$$

$$\text{Heat released due to combustion of } \frac{1}{22.4} \text{ mol of } H_2(g)$$

$$= 12.78 \text{ KJ}$$

$$\text{Heat released due to combustion of } 1 \text{ mol of } H_2(g)$$

$$= 12.78 \times 22.4 = 286.27 \text{ KJ}$$

\therefore approximate standard enthalpy of formation of $\text{H}_2\text{O}(\text{l})$
 $= -286 \text{ KJ}$

3. The standard enthalpy and entropy changes for the reaction in equilibrium for the forward reaction are given below.



$$\Delta H^\circ_{300\text{K}} = -41.16 \text{ kJ mol}^{-1}$$

$$\Delta S^\circ_{300\text{K}} = -4.24 \times 10^{-2} \text{ kJ mol}^{-1}$$

$$\Delta H^\circ_{1200\text{K}} = -32.93 \text{ kJ mol}^{-1}$$

Then, the incorrect statement is -

- (a) The reaction proceeds in the forward direction at 300 K
- (b) At 1200 K, reaction proceeds in the reverse direction
- (c) At 1200 K, $K_p > 1$
- (d) At 300 K, the products will be favoured more than reactants at equilibrium

Answer: (c)

$$\text{At } 300 \text{ K, } \Delta G^\circ = -41.16$$

$$- (300 \times -4.24 \times 10^{-2}) < 0 \Rightarrow \text{spontaneous}$$

$$\text{At } 1200\text{K, } \Delta G^\circ = -32.93$$

$$- (1200 \times (-4.24 \times 10^{-2})) > 0$$

\Rightarrow Non-spontaneous

4. A certain engine which operates in a carnot cycle absorbs 3.347 kJ at 400°C , how much work is done by the engine per cycle? The temperature of sink is 100°C .

pressure of hydrogen at this temperature to reduce the magnitude of ΔG for the reaction to zero -

- (a) 103 atm (b) 103 atm (c) 855 atm (d) 10 atm

Answer: (a)



If $\Delta G = 0$, then $K_{\text{eq}} = Q$. (Q = reaction quotient)

Now from $\Delta G_{\text{f}} = -RT \ln K_{\text{eq}}$

$$\Rightarrow 9.67 = 8.314 \times 500 \ln K$$

$$\Rightarrow K_{\text{eq}} = 10$$

$$\Rightarrow 10 = \frac{(P_{\text{HI}})^2}{(P_{\text{H}_2})^{1/2} (P_{\text{I}_2})^{1/2}} = 10^3 \text{ atm}$$

7. For one mole of an ideal gas which of the following relation is correct?

(a) $\left(\frac{\partial P}{\partial T}\right)_V \left(\frac{\partial V}{\partial P}\right)_T = -1$

(b) $\left(\frac{\partial T}{\partial V}\right)_P \left(\frac{\partial V}{\partial P}\right)_T = 1$

(c) $\left(\frac{\partial P}{\partial T}\right)_V \left(\frac{\partial T}{\partial P}\right)_V = -1$

(d) None of these

Answer: (a)

For one mole of an ideal gas $PV = RT$

or $PdV + VdP = RdT$

Dividing the equation by dT and introducing the condition of constant volume, we get

$$V \left(\frac{\partial P}{\partial T}\right)_V$$

$$\text{or } \left(\frac{\partial P}{\partial T}\right)_V = R \frac{1}{V}$$

$$\left(\frac{\partial V}{\partial P}\right)_T = -\frac{V}{P} \quad \text{and} \quad \left(\frac{\partial V}{\partial T}\right)_P = \frac{V}{T}$$

$$\therefore \frac{P}{T} \left(\frac{\partial V}{\partial T}\right)_P - \left(\frac{\partial V}{\partial P}\right)_T = -R \times \frac{P}{R} \times \frac{V}{P} = -1$$

8. The coefficient of thermal expansion, α , is nearly constant for a liquid $\alpha = \frac{1}{V} \left(\frac{\partial V}{\partial T}\right)_P$

At 293 K and 1 atm for water $\alpha = 2.1 \text{ K}^{-1}$

What is the approximate work done when 1 mole of liquid is heated from 288 K to 298 K at 1 atm? Molar volume of liquid = 18 ml

- (a) -78.25 kJ (b) 78.15 J (c) -37.8 J (d) -83.14 J

Answer: (c)

$$dV = \alpha V dT$$

$$\Delta W = - \int P dV = - \alpha V \times P \int_{T_1}^{T_2} dT$$

$$\text{or } \Delta W = - 2.1 \times 18 \times 10^{-6} \times 10^5 \times 10 \text{ J}$$

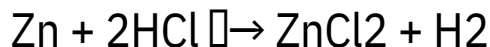
$$\text{or } \Delta W = - 2.1 \times 18 \text{ J}$$

$$\text{or } \Delta W = - 37.8 \text{ J}$$

9. Which of the following is/are correct?
- (a) When 1 mole of Zn is dissolved in excess HCl the work done is approximately equal to -2.46 kJ in open beaker at 300 K and 1 atm.
- (b) When 1 mole of Zn is dissolved in excess HCl work done is equal to zero in closed beaker.

- (c) Both (a) and (b) are correct
 (d) Neither (a) and nor (b) are correct.

Answer: (c)



The no. of moles of H₂ produced = 1 mole

$$\text{Vol. of H(g) produced} = 0.082 \times 300 = 24.6 \text{ L}$$

$$= 24.6 \times 10^{-3} \text{ m}^3$$

$$\therefore \text{ in open beaker, } \Delta W = - P_{\text{ext}} \times \Delta V$$

$$= - 105 \times 24.6 \times 10^{-3} = - 2460 \text{ J}$$

In a closed vessel, $\Delta V = 0$ and $\Delta W = 0$

10. Initial and final state of gas are represented respectively by (P₁, V₁, T₁) and (P₂, V₂, T₂). The pressure and volume of the gas during process is related by PVⁿ = constant. Therefore net reversible work done is

(b) $\frac{P_2 V_2 - P_1 V_1}{(1-n)}$ (c) $-\frac{P_2 V_2}{P_1 V_1}$ (d) $-P \Delta V$

Answer: (b)

PVⁿ = c where c is a constant

$$\therefore P = c/V^n$$

\therefore Work done,

$$\Delta W = - \int_{V_1}^{V_2} P dV = - \int_{V_1}^{V_2} \frac{c}{V^n} dV = - \frac{c}{(n-1)} \left[\frac{1}{V^{n-1}} \right]_{V_1}^{V_2}$$

$$\text{or } \Delta W = \frac{1}{(n-1)} (P_2 V_2 - P_1 V_1)$$

11. For a gas having molar mass M, specific heat at constant pressure per gm of gas is given by

$$\frac{\gamma R}{M(\gamma-1)} \quad (a) \quad \text{---} \quad (b) \quad \frac{\gamma}{R M} \quad (c) \quad \frac{M}{R(\gamma-1)} \quad (d) \quad \frac{\gamma R M}{(\gamma+1)}$$

Answer: (a)

$$C_p - C_v = R \quad \text{and} \quad \frac{C_p}{C_v} = \gamma$$

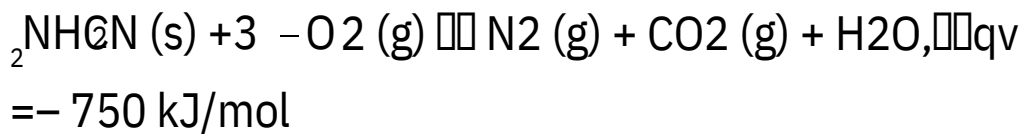
$$\therefore C_p - C_p \frac{1}{\gamma} = R$$

$$\text{or } C_p \left[1 - \frac{1}{\gamma} \right] = R$$

$$\text{or } C_p = \frac{R \gamma}{\gamma - 1} = M C_p$$

$$\text{or } C_p = \frac{\gamma R}{M(\gamma-1)}$$

12. The following reaction occurs in bomb calorimeter at 300K.



For this reaction, find the value of ΔH for 3 moles of $\text{NH}_2\text{CN (s)}$ -

- (a) - 2224.53.50 kJ/mol (b) - 2224.38 kJ/mol
 (c) - 1422.78.46 kJ/mol (d) -2242.0 kJ/mol

Answer: (b)

From first law of thermodynamics,

$$\Delta H = \Delta U + \Delta n g R T$$

Here,

$\Delta n g$ = no of mole of gaseous products - no of mole of gaseous reactant

The combustion reaction of NH_2CN (s) is



$$\Delta n = 0.5$$

$$\text{Now, } \Delta H = \Delta U + \Delta n gRT$$

$$= -750 \text{ KJ} + (0.5 \text{ mol} \times 8.314 \times 1000 \text{ KJ/mol/K} \times 300 \text{ K}) = -742 \text{ KJ} + 1.238786 \text{ KJ} = -741.46 \text{ KJ}$$

The enthalpy change for the combustion reaction of 1 mole of NH_2CN is -741.46 KJ .

Thus, The enthalpy change for the combustion reaction of 3 moles of NH_2CN is -741.46×3

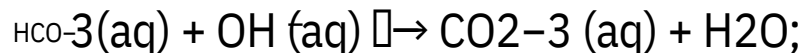
$$= -2224.38 \text{ KJ}.$$

13. The enthalpy of neutralisation of HCO_3^- (aq) with strong alkali is -42 kJ/mole and enthalpy of neutralisation of strong acid with strong alkali is -56 kJ/equiv . Therefore, enthalpy of dissociation of HCO_3^- (aq) is-

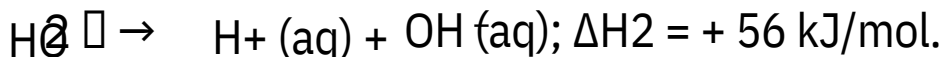
(a) -98 kJ/mol (b) 98 kJ/mol

(c) 14 kJ/mol (d) 24 kJ/mol

Answer: (c)



$$\Delta H_1 = -42 \text{ kJ/mol}.$$

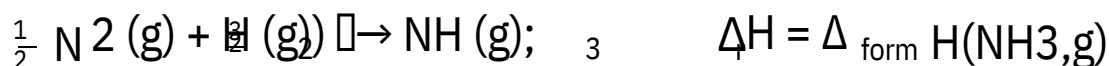




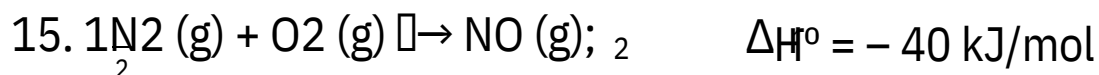
14. If bond dissociation energies of $\text{N} \equiv \text{N}$, $\text{H}-\text{H}$ and $\text{N}-\text{H}$ are x_1 , x_2 and x_3 respectively, hence enthalpy of formation of NH_3 (g) is=

- (a) $x_1 + 3x_2 - 6x_3$ (b) $3x_3 - \frac{1}{2}x_1 - \frac{3}{2}x_2$
 (c) $\frac{1}{2}x_1 + \frac{3}{2}x_2 - 3x_3$ (d) $6x_3 - x_1 - 3x_2$

Answer: (c)



$$\begin{aligned} \Delta_{\text{form}} H (\text{NH}_3, \text{g}) &= \frac{1}{2} D_{\text{N} \equiv \text{N}} + \frac{3}{2} D_{\text{H}-\text{H}} - 3 D_{\text{N}-\text{H}} \\ &= \frac{1}{2} x_1 + 3 \frac{x_2}{2} - 3x_3 \end{aligned}$$



Given: $C_{p,m} (\text{NO}_2, \text{g}) = 40 \text{ J/mol/K}$;

$C_{p,m} (\text{O}_2, \text{g}) = 30 \text{ JK}^{-1} \text{ mol}^{-1}$

$C_{p,m} (\text{N}_2, \text{g}) = 30 \text{ JK}^{-1} \text{ mol}^{-1}$

What is the enthalpy of formation of NO_2 (g) at 1298 K?

- (a) -40 kJ/mol (b) -50 kJ/mol
 (c) -45 kJ/mol (d) -6 kJ/mol

Answer: (c)

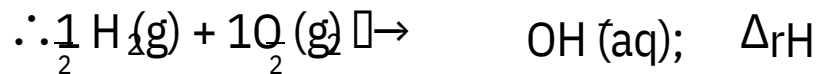
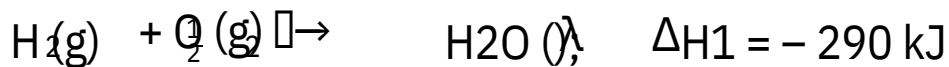
$$\Delta_r H_T = \Delta_r H^\circ + \int_{298}^{1298} \Delta C_p dT$$

$$\text{At } 1298 \text{ K} \quad \Delta_r H = -40 \text{ kJ} - 5$$

$$\Delta T = -40 \text{ kJ} - 5 \times 1000 \times 10^{-3} \text{ kJ} = -45 \text{ kJ/mol}$$

16. The enthalpy of formation of $\text{H}_2\text{O}(\lambda)$ is -290 kJ/mol and enthalpy of neutralisation of strong acid with strong alkali is -56 kJ/equiv . What is the enthalpy of formation of $\text{OH}^-(\text{aq})$? Given $\Delta_{\text{form}} H(\text{H}^+(\text{aq})) = 0$
- (a) -334 kJ/mol (b) -234 kJ/mol
 (c) -346 kJ/mol (d) -178 kJ/mol

Answer: (b)



$$\Delta_r H = -290 + 56 = -234 \text{ kJ/mol}$$

17. For the water gas reaction



At 1000 K the standard Gibbs free energy change of the reaction is -8.314 kJ/mol . Therefore at 1000 K the equilibrium constant of the above water gas reaction is-

- (a) 1 (b) 10 (c) $\frac{1}{e}$ (d) 2.718

Answer: (d)

$$\Delta_r G^\circ = -8.314 \times 10^3$$

$$\therefore -\Delta_r G^\circ = 8.314 \times 10^3 = RT \ln K_{\text{eq}}$$

$$\text{or } 8.314 \times 10^3 = 8.314 \times 10^3 \ln K_{eq}$$

$$\text{or } \ln K_{eq} = 1 = \ln e$$

$$\text{or } K_{eq} = e = 2.718$$

18. At 500 K for an isobaric process.

$$\Delta S_{\text{system}} = -10 \frac{\text{kJ}}{\text{molK}} \text{ and } \Delta S_{\text{surr}} = 12 \frac{\text{kJ}}{\text{molK}}$$

Therefore, ΔG for the entire process is-

(a) -500 kJ/mol

(b) -1000 kJ/mol

(c) -600 kJ/mol

(d) -1100 kJ/mol

Answer: (b)

$$\Delta S_{\text{total}} = \Delta S_{\text{system}} - \frac{\Delta H}{T} = 2 \text{ kJ/mol.}$$

$$\text{or } \frac{\Delta H}{T} - T\Delta S = -2 \text{ kJ/mol.}$$

$$\text{or } \Delta H - T\Delta S = -2T \text{ kJ/mol.} = -2 \times 500 \text{ kJ/mol}$$

$$= -1000 \frac{\text{kJ}}{\text{mol}}$$

19. For the oxidation of glucose, $\Delta_r H^\circ = -2808 \text{ kJ/mol}$ and $\Delta_r G^\circ = -3000 \text{ kJ/mol}$, 25% of energy of the is oxidised for muscle work. Therefore, in order to climb a hill of height 500 metres, how many gm of glucose is required for a man of mass 100 kg? ($g = 10 \text{ m/s}^2$)

(a) 100 gm

(b) 180 gm

(c) 200 gm

(d) 120 gm

Answer: (d)

$$\text{PE required} = mgh = 100 \times 10 \times 500 = 5 \times 10^5 \text{ J}$$

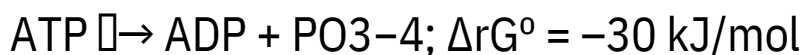
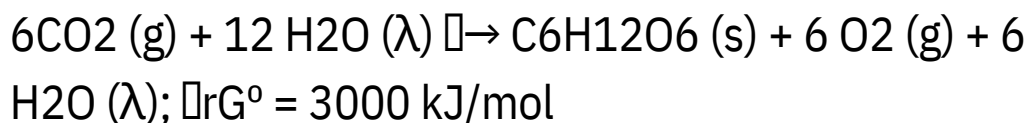
Let W gm of glucose required

$$\therefore \frac{W}{180} \times 3000 \times \frac{103}{4} = 5 \times 105$$

$$\text{or } \frac{W}{180} \times 3 \times \frac{106}{4} = 5 \times 105$$

$$\text{or } W = \frac{2}{3} \times 180 = 120 \text{ gm}$$

20. For the reaction in the plant cells



Glucose is stored in the plant cell as starch, $(\text{C}_6\text{H}_{10}\text{O}_5)_n$. To produce 162 gm of starch how many moles of ATP are minimum required?

- (a) 100 mol (b) 10 mol (c) 20 mol (d) 200 mol

Answer: (a)

$$\text{Mol wt of starch} = 162 n$$

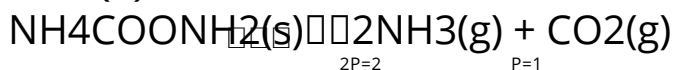
$$\therefore \text{moles of } \text{C}_6\text{H}_{12}\text{O}_6 \text{ to be produced} = 1 \text{ mol}$$

$$\therefore \text{moles of ATP required} = \frac{3000}{30} = 100 \text{ mol}$$

21. $\text{NH}_4\text{COONH}_2(\text{s}) \rightleftharpoons 2\text{NH}_3(\text{g}) + \text{CO}_2(\text{g})$. If equilibrium pressure is 3 atm for the above reaction; K_p will be -

- (a) 4 (b) 27 (c) 4/27 (d) 1/27

Answer: (a)



$$3P = 3 \Rightarrow P = 1$$

$$K_p = p_{\text{NH}_3}^2 \times p_{\text{CO}_2} = 2^2 \times 1 = 4$$

22. Pre exponential factor for free radical combination of chlorine is $2 \times 10^{13} \text{ sec}^{-1}$. the rate constant (in hr^{-1}) of this reaction at STP -

- (a) $7.2 \times 10^{16} \text{ hr}^{-1}$ (b) $3.6 \times 10^{16} \text{ hr}^{-1}$
 (c) $14.4 \times 10^{16} \text{ hr}^{-1}$ 1 (d) $7.2 \times 10^{15} \text{ hr}^{-1}$

Answer: (a)

For free radical combination reaction $E_a = 0$

$$K = A e^{-E_a/RT} \Rightarrow \text{Putting } E_a = 0$$

$$K = 2 \times 10^{13} \text{ sec}^{-1}$$

$$K = 2 \times 10^{13} \times 3600 \text{ hr}^{-1}$$

$$K = 7.2 \times 10^{16} \text{ hr}^{-1}$$

23. The work done during expansion of a gas is independent of pressure if the expansion takes place

- (a) isothermally (b) adiabatically
 (c) in both the above cases (d) in none of the above cases

Answer: (d)

During isothermal expansion

Work done, $W = nRT \log \frac{V_f}{V_i}$ & we also know for

$$\text{isothermal change } \frac{V_f}{V_i} = \frac{P_i}{P_f}$$

While during adiabatic expansion

$$\text{Work done, } W = \frac{P_f V_f - P_i V_i}{\gamma - 1}$$

We see in both these cases work done depends on pressure.

24. To raise the temperature of 2 mole diatomic gas by 1°C at constant pressure, the required amount of heat is 120 cal. The amount of heat which goes as the gas's internal energy is almost.
- (a) 82.5 Cal (b) 80 Cal (c) 85.2 Cal (d) 58.8 Cal

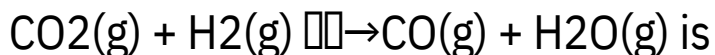
Answer: (c)

C_p and C_v for a diatomic gas is $\frac{7}{2} R$ and $\frac{5}{2} R$ respectively.

The gas's internal energy is increase by only 0.71 of the supplied energy. The rest is used to do work against external pressure.

$$0.71 \times 120 = 85.2 \text{ Cal}$$

25. The ΔH_{of} for $\text{CO}_2(\text{g})$, $\text{CO}(\text{g})$ and $\text{H}_2\text{O}(\text{g})$ are -393.5 , -110.5 and $-241.8 \text{ kJ mol}^{-1}$ respectively. The standard enthalpy change (in kJ) for the reaction,



- (a) 524.1 (b) 41.2 (c) -262.5 (d) -41.2

Answer: (b)

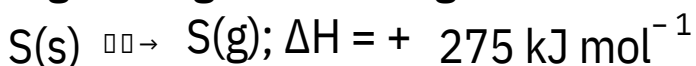
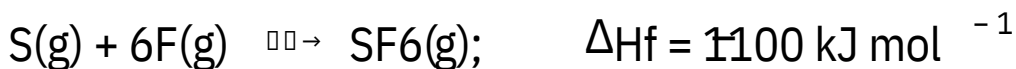
$$\Delta H_{\text{oreaction}} = \Delta H_f(\text{H}_2\text{O}, \text{g}) - \Delta H_f(\text{CO}_2, \text{g})$$

present in the system e.g. mass, volume, energy, entropy, Gibbs potential, enthalpy. Whereas intensive properties are those properties of a system, which are characteristic of the substance or substances present, and are independent of its mass. Temperature, pressure, density, heat capacity, boiling point etc. are intensive property.

28. The standard heat of formation values of SF₆(g), S(g) and F(g) are: -1100, 275 and 80 kJ mol⁻¹ respectively. Then the average S – F bond energy in SF₆.

- (a) 301 kJ mol⁻¹ (b) 320 kJ mol⁻¹
 (c) 309 kJ mol⁻¹ (d) 280 kJ mol⁻¹

Answer: (c)



Therefore heat of formation = Bond energy of reactant – Bond energy of product

$$-1100 = [275 + 6 \times 80] - [6 \times S - F]$$

Thus bond energy of S – F = 309 kJ mol⁻¹

29. At a particular temperature H⁺(aq) + OH⁻(aq) → H₂O(l);
 $\Delta H = -57.1 \text{ kJ}$

31. How many gram of butane must be burnt to heat 10 kg water from 30°C to 100°C. Heat of combustion of butane is -700 kcal/mol and specific heat of water is 1 cal gm⁻¹ K⁻¹.
 (a) 58 gm (b) 56 gm (c) 54 gm (d) 52 gm

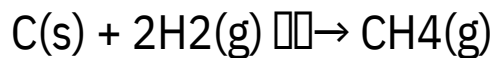
Answer: (a)

$$\text{Heat} = m \times s \times t = 10,000 \times 1 \times 70 = 700,000 \text{ cal} = 700 \text{ kcal}$$

$$\Delta H_{\text{comb}}(\text{C}_4\text{H}_{10}) = 700 \text{ kcal/mol}$$

$$\therefore \text{Mass of C}_4\text{H}_{10} \text{ required} = 58 \text{ gm}$$

32. The enthalpy of atomization of carbon and hydrogen are q_1 and q_2 respectively. If q_3 is the B.E. of C-H bond, then ΔH of the following reaction is



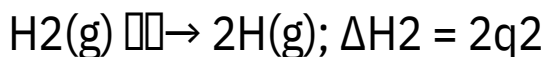
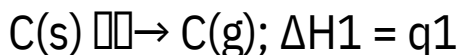
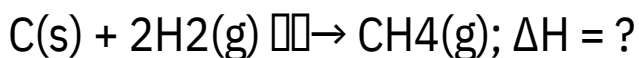
(a) $4q_3 - q_1 - 2q_2$

(b) $q_1 + q_2 - q_3$

(c) $q_1 + 2q_2 - 4q_3$

(d) $q_1 + 4q_2 - 4q_3$

Answer: (d)



$$\Delta H = q_1 + 4q_2 - 4q_3$$

33. An ideal gas is allowed to expand both reversibly and irreversibly in an isolated system. If T_i is the initial temperature and T_f is the final temperature, which of the following statements is correct?

(a) $T_f > T_i$ for reversible process but $T_f = T_i$ for irreversible process.

(b) $(T_f)_{\text{rev}} = (T_f)_{\text{irrev}}$

(c) $T_f = T_i$ for both reversible and irreversible processes

(d) $(T_f)_{\text{irrev}} > (T_f)_{\text{rev}}$

Answer: (d)

In isolated system the expansion of gas is carried out adiabatically. Since heat exchange between system and surrounding is not possible i.e. $q = 0$ and secondary work is always greater than work for reversible process there must be comparatively higher decreases in internal energy i.e. ΔU for reversible process will be more negative. Hence final temperature in reversible process will be smaller than irreversible process.

$\therefore (T_f)_{\text{irrev}} > (T_f)_{\text{rev}}$

34. $(\Delta H - \Delta U)$ for the formation of carbon monoxide (CO) from its elements at 298 K is ($R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$)

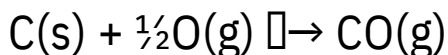
(a) $1238.78 \text{ J mol}^{-1}$

(b) $-2477.57 \text{ J mol}^{-1}$

(c) $2477.57 \text{ J mol}^{-1}$

(d) $-1238.78 \text{ J mol}^{-1}$

Answer: (d)



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

$$(\Delta H - \Delta U) = \Delta n_g RT = \frac{1}{2} \times 8.314 \times 298 = 1238.78 \text{ J mol}^{-1}$$

35. If ΔH is the enthalpy change and ΔE the change in internal energy accompanying a gaseous reaction then
- (A) ΔH is always less than ΔE .
 - (B) ΔH is always greater than ΔE .
 - (C) ΔH is less than ΔE if the number of moles of gaseous products is greater than the number of moles of gaseous reactants.
 - (D) ΔH is less than ΔE if the number of moles of gaseous products is less than the number of moles of gaseous reactants.

Answer: (d)

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

ΔH value is less than or greater than ΔE depending on the value of Δn_g which is the change in number of moles of the gaseous components.

As result (a) and (b) cannot be true.

$\Delta H < \Delta E$ when $\Delta n_g < 0$, hence (c) is also false only (d) is correct.

36. The molar heat capacity for a gas at constant T and P is
- (a) $\frac{3}{2}R$
 - (b) $\frac{5}{2}R$
 - (c) dependent on the atomicity of the gas
 - (d) infinity

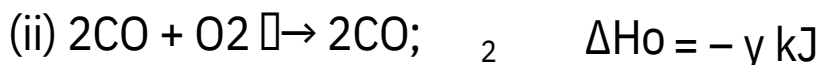
Answer: (d)

$$C_p = \left(\frac{\partial H}{\partial T} \right)_P$$

Since at constant T, $dT = 0$

$$\therefore dP = \infty$$

37. Given



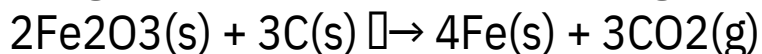
The enthalpy of formation of CO_2 will be

(a) $y - 2x$ (b) $2x - \frac{y}{2}$ (c) $\frac{y-2x}{2}$ (d) $2x - y$

Answer: (c)

The enthalpy of formation of CO_2 will be determined by the Hess's law of constant heat summation.

38. Smelting of iron ore takes place through this reaction



ΔH_o of Fe_2O_3 and CO are $-8242 \text{ kJ mole}^{-1}$ and $-393.7 \text{ kJ mole}^{-1}$ respectively.

The reaction is

(a) endothermic

(b) exothermic

(c) $\Delta H = 0$

(d) none of the above

Answer: (a)

$$\begin{aligned} \Delta H &= (\Delta H_o)_{\text{products}} - (\Delta H_o)_{\text{reactants}} \\ &= 3 \times (-393.7) - 2 \times (-824.2) \\ &= 476.3 \text{ kJ/mole} \end{aligned}$$

39. 5 mole of an ideal gas expand reversibly from a volume of 8 dm³ to 80 dm³ at a temperature of 27°C. Calculate the change in entropy.

(a) 70.26 JK⁻¹

(b) 82.55 JK⁻¹

(c) 95.73 JK⁻¹

(d) 107.11 JK⁻¹

Answer: (c)

$$\begin{aligned}\text{Entropy change } (\Delta S) &= 2.303 nR \log \frac{V_2}{V_1} = 2.303 \times 5 \times 8.314 \\ &\times \log 80 \\ &= 95.73 \text{ JK}^{-1}\end{aligned}$$

40. The standard heat of combustion of a hydrocarbon compound is an/a-

(a) Extensive property

(b) Colligative property

(c) Intensive property

(d) Constitutive property

Answer: (c)

The standard heat of combustion is defined as the heat evolved on burning one mole of the compound in presence of O₂ to give CO₂ and water. Since the amount of the substance is specified, it becomes an intensive property.

41. In which case of mixing of a strong acid and a base, each of 1(N) concentration, temperature-increase is the highest?

(a) 20 ml acid and 30 ml alkali

(b) 10 ml acid and 40 ml alkali

(c) 25 ml acid and 25 ml alkali

(d) 35 ml acid and 15 ml alkali

Answer: (c)

When 25 ml acid is reacting with 25 ml of alkali, maximum equivalent of acid is reacting with alkali since the volume of the resulting solution is the same in every cases, therefore, increase in temperature is the highest where maximum equivalents of acids and alkalies are reacting.

42. If a process is both endothermic and spontaneous, then:

(a) $\Delta S > 0$ (b) $\Delta S < 0$ (c) $\Delta H < 0$ (d) $\Delta G > 0$

Answer: (a)

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

For spontaneous process, $\Delta G =$ negative

For endothermic process, $\Delta H =$ positive

Therefore, $\Delta S > 0$

43. The product of pressure and volume decreases in an

(a) adiabatic expansion (b) adiabatic compression
(c) Both a and b (d) None of these

Answer: (b)

In an adiabatic expansion, energy is transferred only as work.

Thus, internal energy decreases and hence temperature decreases. from equation of state of ideal gas,

To predict the direction of process, entropy and free energy change values also used.

The standard enthalpy of graphite is *zero* at 298K and 1 atm pressure and the standard enthalpy of a diamond is 1.8kJ/mol at 298K and 1 atm pressure.

In the case of an exothermic reaction, the overall enthalpy of the products is lesser than that of reactants.

An additional data of heat of vapourisation of $\text{Br}_2(\text{l})$ is also necessary for the calculation of ΔH .