

NEET ANSWER KEY & SOLUTIONS

SUBJECT :- CHEMISTRY

CLASS :- 11th

PAPER CODE :- CWT-5

CHAPTER :- THERMODYNAMICS & THERMOCHEMISTRY

ANSWER KEY

1. (C)	2. (D)	3. (A)	4. (D)	5. (B)	6. (A)	7. (B)
8. (A)	9. (A)	10. (D)	11. (C)	12. (D)	13. (C)	14. (D)
15. (B)	16. (C)	17. (A)	18. (A)	19. (D)	20. (C)	21. (D)
22. (C)	23. (B)	24. (C)	25. (B)	26. (C)	27. (D)	28. (D)
29. (B)	30. (C)	31. (C)	32. (B)	33. (B)	34. (C)	35. (B)
36. (D)	37. (A)	38. (C)	39. (A)	40. (C)	41. (A)	42. (C)
43. (C)	44. (A)	45. (D)	46. (A)	47. (C)	48. (B)	49. (D)
50. (A)						

SOLUTIONS

SECTION-A

1. (C)
Sol. Test tube is open to atmosphere.
2. (D)
Sol. The properties of the system value of which is independent of the amount of substance present in the system are called intensive properties e.g., viscosity, surface tension temperature, pressure etc.
3. (A)
Sol. The fixed quantity of any substance is associated with a definite amount of energy which depends upon chemical nature of the substance and its state of energy. This energy is called Internal energy or intrinsic energy (μ). μ is a state function and increases with increase in temperature .
4. (D)
Sol. $W = +ve$ because work done on system.
 $q = 0$ because thermally insulated container.
5. (B)
Sol. case (i) $\Delta V = 0, W = 0$
 case (ii) $P = \text{constant}, W = -P(2V_1 - V_1) = -PV_1$
6. (A)
Sol. $A \rightarrow B$, volume is not changing (Isochoric)
 $B \rightarrow C$ Isobaric
 $C \rightarrow A$ Temperature is constant (Isothermal)

7. (B)
Sol. According to first law of thermodynamics "Energy can neither be created nor be destroyed, it can only be converted from one form to another"
 Mathematically
 $\Delta E \text{ or } \Delta U = q + W$
 $\Delta U = \text{change in internal energy}$
 $q = \text{heat absorbed}$
 $W = \text{work done}$
 NOTE: Heat absorbed by the system $q = +ve$
 Heat given out by the system $q = -ve$
 Work done on the system $W = +ve$
 Work done by the system $W = -ve$
8. (A)
Sol. We know that $\Delta U = Q + W = 600 + (-300) = 300 \text{ J}$
 $W = -300$, because the work done by the system.
9. (A)
Sol. For isochoric process $\Delta V = 0$ so $q_v = \Delta U$
 i.e. heat given to a system under constant volume is used up in increasing ΔU .
10. (D)
Sol. $\Delta E = 0$, in a cyclic process.
11. (C)
Sol. The heat exchanged in a chemical reaction at the constant temperature and pressure is known as enthalpy of the reaction.
12. (D)
Sol. In isothermal reversible process ideal gas has constant temperature and so $\Delta U = 0$ and $\Delta H = \Delta U = 0$.

13. (C)
Sol. From 1st law of thermodynamics
 $\Delta U = q + W$
 $Q = + 800 \text{ J}$
 $W = -P (V_2 - V_1)$
 $= -1 (20 - 10) = -10 \text{ dm}^3 \text{ atm}$
 $= - \frac{10 \text{ dm}^3 \times 8.314 \text{ J/K mol}}{0.08206 \text{ dm}^3 / \text{K mol}}$
 $W = -1013 \text{ J}$
 $\Delta U = 800 \text{ J} + (-1013 \text{ J}) = -213 \text{ J}$

14. (D)
Sol. $\Delta H = \text{heat gives so process is isobaric.}$
 $w = -P_{\text{ext}} (\Delta V) = -1 \times 2 \text{ litre.atm} = -202.6 \text{ J}$
Hence $q = 202.6 \text{ J} = \Delta H$. and $\Delta E = q + w = 0$.

15. (B)
Sol. From I law of thermodynamics $\Delta E = Q + W$
where $Q = 0$ for adiabatic process.

16. (C)
17. (A)
Sol. $W_{\text{max, iso}} = -2.303 nRT \log \frac{V_2}{V_1}$
 $= -2.303 \times 1 \times 8.31 \times 310 \times \log \frac{25}{15}$
 $= -1303.73 \text{ J}$

(-ve sign shows work of expansion i.e. work done by the system)

18. (A)
Sol. Since C to D is a reversible isothermal process

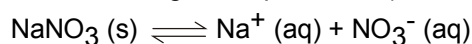
$$W = -2 RT \ln \frac{P_1}{P_2} = -800 R \ln 2$$

19. (D)
Sol. When a real gas is forced through a porous plug into a region of low pressure, it is found that due to expansion, the gas on the side of low pressure gets cooled. The phenomenon of producing lowering of temperature when a gas is made to expand adiabatically from a region of high pressure into a region of low pressure is known as Joule-Thomson effect.

20. (C)
Sol. On occurrence of
 $\text{NH}_4\text{Cl (s)} \longrightarrow \text{NH}_3(\text{g}) + \text{HCl (g)}$
In above reaction, the disorder increases with the formation of NH_3 and HCl . Therefore entropy increases.

21. (D)
Sol. $\Delta n_g < 0$ for $3\text{C}_2\text{H}_2(\text{g}) \longrightarrow \text{C}_6\text{H}_6(\ell)$

22. (C)
Sol. An entropy (ΔS) of a reaction is positive if the products are in more random state as compared to reactants (order of randomness is gas > liquid > solid).



This process is accompanied by increase in randomness because Na^+ and NO_3^- ions are free to move in solution whereas they are not free to move in solid NaNO_3 .

In (A) product (solid) has lesser entropy as compared to reactants.

In (B) NaCl (solid) has lesser entropy as compared to NaCl (aq) (Na^+ and Cl^-).

In (D) both reactants and products are in gaseous state but number of moles of products are decreasing.

23. (B)
Sol. Reaction is endothermic and no. of moles increases on decomposition.

24. (C)
Sol. $\Delta S = nC_{p,m} \ln \frac{T_2}{T_1} = 2 \times \frac{5}{2} R \ln \frac{600}{300} = 5 R \ln 2$

25. (B)
Sol. An irreversible process
 \Rightarrow spontaneous process
 $\Rightarrow (dS)_{V,E} \text{ (change in entropy)} = +ve > 0$
 $\Rightarrow (dG)_{T,P} \text{ (change in Gibbs free energy)}$
 $-ve \Rightarrow < 0$

26. (C)
Sol. $\text{ZnO} \longrightarrow \text{Zn (s)} + \frac{1}{2} \text{O}_2(\text{g})$;
 $\text{HgO} \longrightarrow \text{Hg}(\ell) + \frac{1}{2} \text{O}_2(\text{g})$
 $\Delta H^\circ = +348.3 \text{ kJ} \ \& \ \Delta G^\circ = 318.4$
 $\Delta H^\circ = 90.8 \text{ kJ} \ \& \ \Delta G^\circ = 58.5$
on comparing ΔG° values, decomposition of HgO is more favorable.

27. (D)
Sol. $\Delta G = (\Delta H) - T(\Delta S)$
 $\downarrow \quad \downarrow$
 $-ve \quad -ve$
 since both are $-ve$, the reaction would have a $-ve \Delta G$ below a temperature of K ($= 569K$)

28. (D)
Sol. using $\Delta H = \Delta E + \Delta n_g RT$

29. (B)
Sol. $2C + O_2 \rightarrow 2CO$; $\Delta H = -220$ KJ
 This reaction does not represent complete combustion of carbon, hence heat of combustion of carbon will not be equal to 110 kJ . The negative sign of ΔH indicates that this reaction is exothermic. Also, despite being spontaneous reaction, it requires initiation.

30. (C)
Sol. Enthalpy of reaction (ΔH) is defined as heat exchanged during any chemical reaction

$$\Delta H = H_P - H_R$$
 for exothermic reaction $H_P < H_R$
 $\therefore \Delta H$ is $-ve$.

31. (C)
Sol. $C_2H_2(g) + H_2(g) \rightarrow C_2H_4(g)$
 $\Delta H = \Delta V + \Delta n_g RT$
 $= \Delta U + P\Delta V$
 $-0.5 = \Delta U + 1.5(-5) \times \frac{100}{1000}$
 $\Delta U = -0.5 + 0.75$
 $\Delta U = 0.25$ kJ

32. (B)
Sol. $\Delta H_f^\circ (Cl_2, g) = 0$, As ΔH_f° of elements in their standard state is taken to be zero.

33. (B)
Sol. Black phosphorus has a highly polymeric layer type structure. Hence, thermodynamically, it is the most stable form of phosphorus.

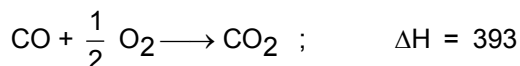
34. (C)
Sol. $C + \frac{1}{2} O_2(g) \rightarrow CO$; $\Delta H = -$

$$110 \text{ kJ mol}^{-1} \dots (i)$$



$$393 \text{ kJ mol}^{-1} \dots (ii)$$

On subtracting eq. (i) from eq. (ii), we get



$$-(-110) = -283 \text{ kJ mol}^{-1}$$

35. (B)
Sol. Combustion reaction of solid boron

$$B(s) + \frac{3}{4} O_2(g) \rightarrow \frac{1}{2} B_2O_3(s)$$

$$\Delta H_r^\circ = \Delta H_c^\circ = \frac{1}{2} \Delta H_f^\circ (B_2O_3, s) - \Delta H_f^\circ (B, s) - \frac{3}{4} \Delta H_f^\circ (O_2, g)$$

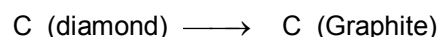
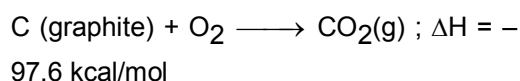
 ΔH_f° of element in stable state of aggregation is assumed to be zero.

$$\Delta H_c^\circ = \frac{1}{2} \Delta H_f^\circ (B_2O_3, s)$$

SECTION-B

36. (D)
Sol. $C(\text{diamond}) + O_2 \rightarrow CO_2(g)$; $\Delta H = -$

$$94.3 \text{ kcal/mol}$$



$$\Delta H = 3.3 \text{ kcal}$$

Heat required to convert 12 gram diamond to graphite = 3.3

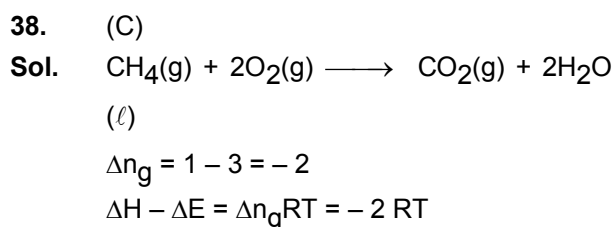
$$\therefore \text{Heat required to convert 1 g diamond to graphite} = \frac{3.3}{12} = 0.275$$

37. (A)
Sol. Calorific value = $\frac{\Delta H}{\text{mol. wt}}$

$$\text{for ethane } (C_2H_6) = - \frac{341}{30} = -11.37$$

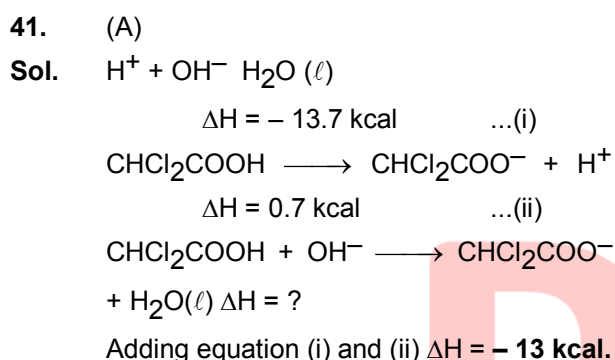
$$\text{for ethyne } (C_2H_2) = - \frac{310}{26} = -11.9$$

So, better calorific value is better fuel hence, C_2H_2 .

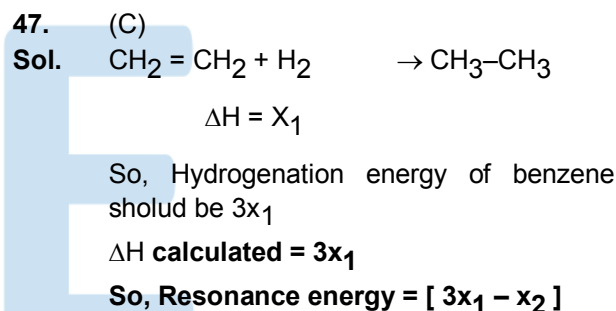
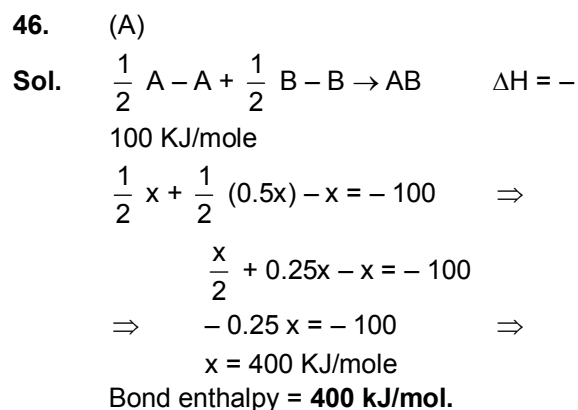
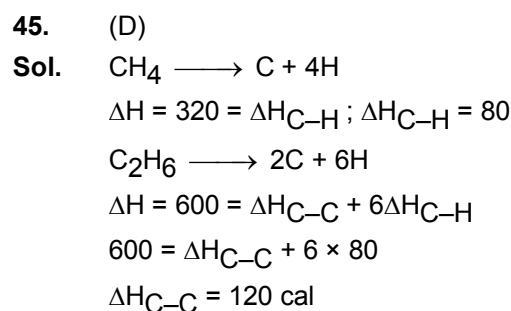
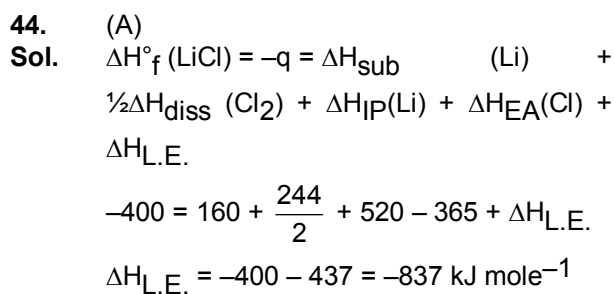
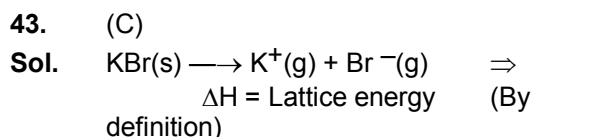


39. (A)
Sol. Since process is exothermic then heat is evolved due to this temperature of water increase.

40. (C)
Sol. Strong acid and strong base neutralization heat is equal to -13.7 kcal.



42. (C)
Sol. Enthalpy of neutralization is defined as amount of heat liberated when one mole of a strong acid is completely neutralized by one mole of a strong base. Its value is less in case of weak acid or weak base because small amount of heat is utilized in ionising the weak acid/base.
 ΔH for ionisation of CH_3COOH = Heat of neutralization for CH_3COOH - Heat of neutralization of strong acid
 $= -50.6 - (-55.9)$ kJ / mol
 $= +5.3$ kJ / mol



48. (B)
Sol. (B) In an isothermal process change in internal energy (ΔE) is zero (as it is a function of temperature). \therefore According to first law of thermodynamics
 $\therefore q + w = \Delta E$, hence $Q = -W$ (if $\Delta E = 0$)
 If a system undergoes a change in which internal energy of the system remains constant (i.e., $\Delta E = 0$), then $-W = Q$.
 This means that work done by the system equals the heat absorbed by the system.

49. (D)
Sol. (D) If a refrigerator's door is kept open room gets heated as heat energy is released to the surroundings.

50. (A)
Sol. (A) The properties whose magnitude depends upon the quantity of matter present in the system are called extensive properties.