Ashoka Scientific Forum

Developing Scientific Temper Among Students

IIT-JAM | TIFR | DU | BHU | JNU | ISM | PU & All Other University M.Sc. Entrance Exam
CSIR-NET | GATE | TIFR

THERMODYNAMICS-3

1. Heat product in calories by the combustion of 1g of carbon is called?
   (a) Heat of combustion of carbon (b) Heat of formation of carbon
   (c) Calorific Value of carbon (d) Heat of product of carbon
2. The temperature of the system decreases in an:
   (a) Adiabatic Compression (b) Isothermal Compression
   (c) Isothermal Expansion (d) Adiabatic Expansion
3. For the isothermal expansion of an ideal gas:
   (a) E and H increases (b) E increases but H decreases
   (c) H increases and E decreases (d) E and H are unaltered
4. In an isochoric process, the increased internal energy is:
   (a) Equal to the heat absorbed (b) Equal to heat evolved
   (c) Equal to the work done (d) Equal to sum of heat absorbed and work done
5. Internal energy is an example of:
   (a) Path Function (b) State Function (c) Both a and b (d) None of these
6. The process in which no heat enters or leaves the system is termed as:
   (a) Isochoric (b) Isobaric (c) Isothermal (d) Adiabatic
7. If in a container, neither mass nor heat exchange occurs, then it constitutes:
   (a) Closed System (b) Open System (c) Isolated System (d) Imaginary System
8. Which of the following is true for an adiabatic process?
   (a) $\Delta H = 0$ (b) $\Delta W = 0$ (c) $\Delta Q = 0$ (d) $\Delta V = 0$
9. Among the following, intensive property is:
   (a) Mass (b) Volume (c) Surface Tension (d) Enthalpy
10. For the reaction of one mole of zinc dust with one mole of sulphuric acid in a bomb calorimeter, ΔV and W correspond to:
    (a) $\Delta U < 0, W = 0$ (b) $\Delta U = 0, W < 0$ (c) $\Delta U > 0, W = 0$ (d) $\Delta U = 0, W > 0$
11. Which of the following expressions represent the first law of thermodynamics?
    (a) $\Delta E = -q + W$ (b) $\Delta E = q - W$ (c) $\Delta E = q + W$ (d) $\Delta E = -q - W$
12. At 27°C one mole of an ideal gas is compressed isothermally and reversibly from a pressure of 2 atm to 10 atm. The value of ΔE and q are (R = 2):
    (a) 0, - 965.84 cal (b) – 965.84 cal, - 865.58 cal
    (c) 865.58 cal, - 865.58 cal (d) – 865.58 cal, - 865.58 cal
13. The heat required to raise the temperature of a body by 1K is called:
    (a) Specific Heat (b) Thermal Capacity (c) Water Equivalent (d) None of these
14. Which of the following is true for the reaction:
    H_2O (l) $\rightleftharpoons$ H_2O (g) at 100°C at one atmosphere?
    (a) $\Delta E = 0$ (b) $\Delta H = 0$ (c) $\Delta H = \Delta E$ (d) $\Delta H = T\Delta S$
15. Identify the correct statement regarding entropy:

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(a) At 0°C, the entropy of a perfectly crystalline substance is taken to be zero

(b) At absolute zero of the temperature, the entropy of all perfectly crystalline substance is positive

(c) At absolute zero of the temperature the entropy of all perfectly crystalline substance is taken to be zero

(d) At absolute zero temperature, the entropy of a perfectly crystalline substance is taken to be zero

16. Maximum entropy will be in which of the following?
(a) Ice (b) Liquid Water (c) Snow (d) Water Vapours

17. If enthalpies of formation C_2H_4 (g), CO_2 (g) and H_2O (l) at 250°C and 1 atm. pressure be 52, -394 and -286 KJ mol^{-1} respectively. The enthalpy of combustion of C_2H_4 (g) will be:
(a) + 1412 KJ mol^{-1} (b) – 1412 KJ mol^{-1} (c) + 141.2 KJ mol^{-1} (d) – 141.2 KJ mol^{-1}

18. Heat of neutralization of strong acid and weak base is:
(a) 57.1 KJ mol^{-1} (b) 13.7 KJ mol^{-1} (c) Less than 13.7 Kcal mol^{-1} (d) More than 13.7 Kcal mol^{-1}

19. The heat evolved in the combustion of methane is given by the following equations:
CH_4 (g) + 2 O_2 (g) → CO_2 (g) + H_2O (l) $\Delta H = -890.3$ KJ
How many grams of methane would be required to produce 444.15 KJ of heat of combustion?
(a) 4 g (b) 8 g (c) 12 g (d) 16 g

20. In a calorimeter, the temperature of the calorimeter increases by 6.12K, the heat capacity of the system is 1.23 KJ/g/deg. What is the molar heat of decomposition of NH_4NO_3?
(a) – 7.53 KJ mol^{-1} (b) – 398.1 KJ mol^{-1} (c) – 16.1 KJ mol^{-1} (d) – 602 KJ mol^{-1}

21. Enthalpy change for a reaction does not depend upon:
(a) The physical states of reactants and products
(b) Use of different reactants for the same product
(c) The nature of the intermediate reaction steps
(d) The differences in initial or final temperature of involved substances

22. The heat of combustion of carbon to CO_2 is – 393 KJ/mol. The heat released upon formation of 35.2 g of CO_2 from carbon and oxygen gas is:
(a) + 315 KJ (b) – 31.5 KJ (c) – 315 KJ (d) 31.5 KJ

23. Consider the reaction:
N_2 (g) + 3 H_2 (g) ⇌ 2 NH_3 (g)
Carried out at constant temperature and pressure. If ΔH and ΔU are enthalpy and internal energy changes for the reaction, which of the following expressions is true?
(a) $\Delta H = 0$ (b) $\Delta H = \Delta U$ (c) $\Delta H < \Delta U$ (d) $\Delta H > \Delta U$

24. The work done in ergs for the reversible expansion of one mole of an ideal gas from a volume of 10L to 20L at 25°C is:
(a) $2.303 \times 298 \times 0.82 \log 2$ (b) $298 \times 10^7 \times 8.314 \times 2.303 \log 2$
(c) $2.303 \times 298 \times 0.082 \log 0.5$ (d) $2.303 \times 298 \times 2 \log 2$

25. Work done during isothermal expansion of 1 mole of an ideal gas from 10 atm to 1 atm at 300K is:
(a) 4938.8 J (b) 4138.8 J (c) 5744 J (d) 6257.2 J

26. ΔE^{o} of combustion of isobutylene is – x KJ mol^{-1}. The value of ΔH^{o} is:
(a) $= \Delta E^{o}$ (b) $> \Delta E^{o}$ (c) = 0 (d) $< \Delta E^{o}$

27. The free energy change for a reversible reaction at equilibrium is:
(a) Large positive (b) Small negative (c) Small Positive (d) Zero

28. The heat of neutralization of a strong acid and a strong alkali is 57.0 KJ mol^{-1}. The heat released when 0.5 mole of HNO_3 solution is mixed with 0.2 mole of KOH is:
(a) 57.0 KJ (b) 11.4 KJ (c) 28.5 KJ (d) 34.9 KJ

29. Heat of formation of H_2O is – 188 KJ per mole and H_2O_2 is – 286 KJ mol^{-1}. The enthalpy change for the reaction $2\ H_2O_2 \rightarrow 2\ H_2O + O_2$ is:
(a) 196 KJ (b) – 196 KJ (c) 984 KJ (d) – 984 KJ

30. The values of ΔH for the combustion of ethene and ethyne are – 341.1 and – 310.0 K cal respectively. Which of the following is a better fuel?
(a) C_2H_2 (b) C_2H_4 (c) Both a and b (d) None of these

31. Given that bond energies of H-H and Cl-Cl are 430 and 240 KJ mol^{-1} respectively, and ΔH_f for HCl is – 90 KJ mol^{-1}. Bond enthalpy of HCl is:
(a) 290 KJ mol^{-1} (b) 380 KJ mol^{-1} (c) 425 KJ mol^{-1} (d) 245 KJ mol^{-1}

32. Hess's Law deals with:
(a) A change in heat of reaction (b) Rate of reaction
(c) Equilibrium constant (d) Influence of pressure on volume of gas

33. One mole of methanol when burnt in O_2 gives out 723 KJ mol^{-1} heat. If one mole of O_2 is used, what will be the amount of heat observed?
(a) 723 KJ (b) 924 KJ (c) 482 KJ (d) 241 KJ

34. The enthalpy and entropy change for the reaction:
$Br_2(l) + Cl_2(g) \rightarrow 2\ BrCl(g)$
Are 30 KJ mol^{-1} and 105 $JK^{-1}mol^{-1}$ respectively. The temperature at which the reaction will be at equilibrium is:
(a) 285.7 K (b) 273 K (c) 450 K (d) 300 K

35. The change in the entropy for the fusion of 1 mole of ice is:
(Melting point of ice = 273 K, molar enthalpy of fusion of ice = 6.0 KJ mol^{-1})
(a) 11.73 $JK^{-1}mol^{-1}$ (b) 18.84 $JK^{-1}mol^{-1}$ (c) 21.97 $JK^{-1}mol^{-1}$ (d) [illegible]

36. For a spontaneous process, the correct statement is:
(a) Entropy of the system always increases (b) Free energy of the system always increases
(c) Total entropy change is always negative (d) Total entropy change is always positive

37. The absolute enthalpy of neutralization of the reaction:
$MgO(s) + 2\ HCl(aq.) \rightarrow MgCl_2(aq.) + H_2O(l)$ will be:

(a) Less than – 57.33 KJ mol^{-1} (b) – 57.33 KJ mol^{-1}

(c) Greater than – 57.33 KJ mol^{-1} (d) 57.33 KJ mol^{-1}

38. For the reaction of one mole of zinc dust with one mole of sulphuric acid in a bomb calorimeter, ΔU and W corresponds to:

(a) $\Delta U < 0, W = 0$ (b) $\Delta U < 0, W < 0$ (c) $\Delta U > 0, W = 0$ (d) $\Delta U > 0, W > 0$

39. Standard enthalpy and standard entropy changes for the oxidation of NH_3 at 298 K are – 382.64 KJ mol^{-1} and – 145.6 $JK^{-1}mol^{-1}$, respectively. Standard Gibb's energy change for the same reaction at 298 K is:

(a) – 2221.1 $KJmol^{-1}$ (b) – 339.3 $KJmol^{-1}$ (c) – 439.3 $KJmol^{-1}$ (d) – 523.2 $KJmol^{-1}$

40. $\Delta G = \Delta H - T\Delta S$ was given by:

(a) Faraday (b) Kirchhoff (c) Einstein (d) Gibbs-Helmholtz

ANSWER KEY

1. c	2. d	3. d	4. a	5. b	6. d	7. c	8. c	9. c	10. a
11. c	12. a	13. b	14. d	15. d	16. d	17. b	18. c	19. b	20. d
21. c	22. c	23. b	24. b	25. c	26. b	27. d	28. b	29. a	30. a
31. b	32. a	33. a	34. a	35. c	36. d	37. a	38. a	39. b	40. d

HINTS AND EXPLANATIONS

1) Heat produced in calories by the combustion of 1 g of carbon is called calorific value of carbon
2) When a real gas is forced to through a porous plug into a region of low pressure, it is found that due to the expansion of the gas on the side of low pressure gets cooled
3) In isothermal reversible process, ideal gas has constant volume and so, $\Delta E = 0$ and $\Delta H = \Delta E = 0$
4) For an isochoric process, $\Delta V = 0$, so $q_v = \Delta E$ i.e. heat given to a system under constant volume is used up in increasing ΔE.
5) The functions whose value depends only on the state of a system are known as state functions
6) For adiabatic process, $q = 0$
7) An isolated system neither shows exchange of heat nor matter with surroundings
8) For adiabatic process, $\Delta Q = 0$
9) Surface tension is an intensive property which do not depend upon the quantity of matter present in the system
10) In bomb calorimeter, the process is run in a sealed container and no expansion or compression is allowed, so, $W = 0$ and $\Delta U < 0$
11) $\Delta E = q + w$
12) $W = 2.303\ nRT \log p_2/p_1 = 2.303 \times 2 \times 300 \log 10/2 = 965.84$
    At constant temperature, $\Delta E = 0$
    $\Delta E = q + W$; $q = -W = -965.84$ cal
13) Heat required to raise the temperature of a body by 1 K is called thermal capacity of the body
14) At equilibrium, $\Delta G = 0$ hence, $0 = \Delta H - T\Delta S$ or $\Delta H = T\Delta S$
15) This is the statement of third law of thermodynamics
16) Entropy is the measure of randomness in the molecules. Randomness is maximum in case of gases (water vapours)
17) $C_2H_4 + 3\ O_2 \rightarrow 2\ CO_2 + 2\ H_2O$
    $\Delta H_{reaction} = [2 \times \Delta H_f^o(CO_2) + 2 \times \Delta H_f^o(H_2O)] - [\Delta H_f^o(C_2H_4) + 3 \times \Delta H_f^o(O_2)]$
    $= [2(-394) + 2(-286)] - [52 + 0] = -1412$ KJ
18) Heat of neutralization of strong acid and weak base is less than 13.7 Kcal/mol
19) CH_4 required = $\frac{445.15 \times 16}{890.3} = 8$ g
20) Molecular weight of $NH_4NO_3 = 80$
    Heat evolved = 1.23×6.12
    Molar heat of decomposition = $1.23 \times 6.12 \times 80 = 602$ KJmol^{-1}
21) According to Hess's Law, the enthalpy change for a reaction does not depend upon the nature of intermediate reaction steps
22) $C + O_2 \rightarrow CO_2$; $\Delta H = -393.5$ KJ mol^{-1}
    44 g CO_2 formed by which heat released = -393.5 KJ
    1 g CO_2 formed by which heat released = -393.5/44

35.2 g CO_2 formed by which heat released = $\frac{393.5 \times 35.2}{44}$ = - 315 KJ

23) At constant p or T,
$\Delta H = \Delta U + \Delta nRT$
$\Delta n = n_p - n_R = 2 - 4 = -2$
$\Delta H < \Delta U$

24) $W = 2.303\ nRT \log V_2/V_1 = 298 \times 10^7 \times 8.314 \times 2.303 \log 2$

25) $W = 2.303\ nRT \log P_2/P_1 = 2.303 \times 1 \times 8.314 \times 300 \log 1/10 = 5744.1$ J

26) $CH_3\text{-}C(CH_3)=CH_2\ (g) + 6\ O_2\ (g) \rightarrow 4\ CO_2\ (g) + 4\ H_2O\ (l)$
$\Delta n_g = 4 - 7 = -3$
$\Delta H = \Delta E + \Delta n_g RT$
$\Delta H = \Delta E - \Delta n_g RT$
$\Delta H < \Delta E$

27) At equilibrium, $\Delta G = 0$

28) 0.2 mole will neutralize 0.2 mole of HNO_3, heat evolved = $51 \times 0.2 = 11.4$ KJ

29) $H_2 + \frac{1}{2} O_2 \rightarrow H_2O$; $\Delta H = -188\ KJmol^{-1}$ ..... (i)
$H_2 + O_2 \rightarrow H_2O_2$; $\Delta H = -286\ KJmol^{-1}$ ..... (ii)
$2\ H_2 + O_2 \rightarrow 2\ H_2O$; $\Delta H = -376\ KJ\ mol^{-1}$ ..... (iii)
$2\ H_2 + 2\ O_2 \rightarrow 2\ H_2O_2$; $\Delta H = -572\ KJ\ mol^{-1}$ ..... (iv)
By (iii) – (iv)
$2\ H_2O_2 \rightarrow 2\ H_2O + O_2$; $\Delta H = 196$ KJ

30) ΔH for $C_2H_4 = -341.1$ Kcal
Its calorific value = - 341.1/30 = -11.37 Kcal/g
ΔH for $C_2H_2 = -310.0$ Kcal
Its calorific value = - 310.0/26 = - 11.92
Hence, C_2H_2 is a better fuel

31) $\Delta H_{reaction} = \Delta_{H\text{-}H} + \Delta H_{Cl\text{-}Cl} - 2\Delta H_{HCl} = -90$ KJ
$\Delta H_{H\text{-}Cl} = \frac{430 + 240 - (-90)}{2} = 760/2 = 380\ KJ\ mol^{-1}$

32) It is also known as constant heat summation law. It states that the total amount of heat evolved or absorbed in a reaction is same whether the reaction takes place in one step or in a number of steps

33) $CH_3OH\ (l) + 3/2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (l)$
$\Delta H = -723$ KJ
By burning with 3/2 O_2 (g), heat evolved = - 723 KJ
By burning one mole of O_2 (g) = $(-723 \times 2)/3 = -482$ KJ

34) At equilibrium $\Delta G^0 = 0$ ($\Delta G^0 = \Delta H^0 - T\Delta S$)

$0 = 30 \times 10^3 - T \times 105$

$T = (30 \times 10^3)/105$ K = 285.71 K]

35) Entropy change of fusion $\Delta S_f^0 = \Delta H_f^0/T$

$\Delta S_f^0 = 60 \times 10^3/273 = 21.97$ $JK^{-1}mol^{-1}$

36) For a spontaneous process, total entropy change is always positive

37) Heat of neutralization of strong acid and strong base is 57.33 KJ. MgO is a weak base while HCl is strong acid, so the heat of neutralization of MgO and HCl is lower than – 57.33 KJ because MgO requires some heat in ionization, then net released amount of heat is decreased

38) In bomb calorimeter, heat of combustion is determined at constant volume, hence heat of reaction corresponds to ΔU

$\Delta U < 0$, $W = 0$

39) $\Delta G = \Delta H - T\Delta S = 382.64 - (298 \times -145.6 \times 10^{-3}) = -339.3$ KJ mol^{-1}

40) Gibbs Helmholtz equation $\Delta G = \Delta H - T\Delta S$