Thermodynamics

Class 11 Chemistry · 24 NEET MCQs with answers & explanations

Thermodynamics covers the laws of thermodynamics, enthalpy, Hess's law, entropy, Gibbs free energy and spontaneity. NEET questions involve enthalpy calculations (bond enthalpy, formation enthalpy), predicting spontaneity using $\Delta G = \Delta H - T\Delta S$, and applying Hess's law for indirect enthalpy determination.

Key Thermodynamic Relations

Concept	Key Formula / Fact
First law	$\Delta U = q + w$; for ideal gas isothermal expansion: $\Delta U = 0$
Enthalpy	$\Delta H = \Delta U + \Delta ngRT$; at constant pressure: $q_p = \Delta H$
Hess's law	ΔH is path-independent; sum of steps = overall enthalpy change
Entropy	$\Delta S = q_{rev}/T$; increases with disorder; S(gas) > S(liquid) > S(solid)
Gibbs free energy	$\Delta G = \Delta H - T\Delta S$; spontaneous if $\Delta G < 0$
Spontaneity conditions	$\Delta H < 0, \Delta S > 0$ → always spontaneous; $\Delta H > 0, \Delta S < 0$ → never spontaneous

Questions

Q1. Using combustion enthalpies $\Delta H°c[C(s)] = -94$ kcal mol^{-1}, $\Delta H°c[H_2(g)] = -68$ kcal mol^{-1} and $\Delta H°c[CH_4(g)] = -213$ kcal mol^{-1}, the standard enthalpy of formation of $CH_4(g)$ is:

(A) -111 kcal mol^{-1}
(B) -17 kcal mol^{-1}
(C) -85 kcal mol^{-1}
(D) -170 kcal mol^{-1}

Q2. A closed, insulated container holds a liquid stirred by a paddle until its temperature rises. Using $\Delta U = q + w$, which statement is correct?

(A) $\Delta U = 0$ and $w = q \neq 0$
(B) $w = 0$ and $\Delta U = q \neq 0$
(C) $\Delta U = w \neq 0$ and $q = 0$
(D) $\Delta U = w = q \neq 0$

Q3. For a reversible isothermal expansion of 2 mol of an ideal gas at 300 K from 2 L to 20 L, the entropy change is approximately ($R = 8.314$ J mol^{-1} K^{-1}):

(A) 38.3 J K^{-1}
(B) 19.1 J K^{-1}
(C) 76.6 J K^{-1}
(D) 9.57 J K^{-1}

Q4. For a cyclic process at constant temperature, which statement is necessarily correct?

(A) $\Delta U = q = 0$
(B) $\Delta U > 0, q > 0$
(C) $\Delta U = 0, q = -w$
(D) $\Delta S = 0, w = 0$

Q5. Find the heat released when 35.0 g of CO_2 is formed from carbon by combustion. Given $\Delta Hc(C) = -390$ kJ mol^{-1}.

(A) 245 kJ
(B) 390 kJ
(C) 310 kJ
(D) 700 kJ

Q6. At constant volume, 1 mol of gas is heated from 298 K to 308 K and 500 J of heat is supplied. Which relation is correct?

(A) $q=0$, $w=500$ J, $\Delta U=500$ J
(B) $q=500$ J, $w=0$, $\Delta U=500$ J
(C) $q=500$ J, $w=500$ J, $\Delta U=0$
(D) $q=500$ J, $w=0$, $\Delta U=0$

Q7. When heat is absorbed and work is done BY the system, the correct IUPAC signs are:

(A) q positive, w positive
(B) q positive, w negative
(C) q negative, w positive
(D) q negative, w negative

Q8. At constant volume, which of the following is correct?

(A) $\Delta U = \Delta H$
(B) $\Delta U = q - w$
(C) $\Delta U = 0$
(D) $\Delta U = qv$

Q9. Which of the following correctly defines bond enthalpy?

(A) Energy to break one mole of bonds in gaseous state
(B) Energy released during bond formation in liquid state
(C) Energy absorbed during bond formation in any state
(D) Energy of activation for bond breaking in solution

Q10. Which of the following thermodynamic quantities is always positive?

(A) Enthalpy of formation of all compounds
(B) Enthalpy of combustion
(C) Bond enthalpy
(D) Enthalpy of neutralization of strong acid-base

Q11. Which of the following is the most accurate definition of entropy?

(A) Measure of randomness or disorder in a system
(B) Total heat content of a system
(C) Work done by system on surroundings
(D) Measure of internal energy change

Q12. The correct mathematical expression for entropy change in a reversible process is:

(A) $\Delta S = q_{rev} \times T$
(B) $\Delta S = \Delta H / T$
(C) $\Delta S = q_{rev} / T$
(D) $\Delta S = \Delta G / T$

Q13. Which of the following is the correct definition of Gibbs free energy (G)?

(A) $G = H - TS$
(B) $G = H + TS$
(C) $G = U - TS$
(D) $G = TS - H$

Q14. Which of the following correctly identifies the condition for a spontaneous process at constant T and P?

(A) $\Delta G > 0$
(B) $\Delta G < 0$
(C) $\Delta G = 0$
(D) $\Delta G = \Delta H$

Q15. Which of the following correctly defines an isolated system?

(A) System exchanging both matter and energy
(B) System exchanging only energy
(C) System exchanging only matter
(D) System exchanging neither matter nor energy

Q16. Which type of system allows exchange of energy but not matter?

(A) Open system
(B) Closed system
(C) Isolated system
(D) Adiabatic system

Q17. Which of the following correctly defines enthalpy (H)?

(A) $H = U + PV$
(B) $H = U - PV$
(C) $H = q + w$
(D) $H = U + RT$

Q18. Under which condition is heat exchanged by a system equal to its enthalpy change?

(A) Constant volume
(B) Constant pressure
(C) Adiabatic condition
(D) Constant temperature

Q19. Which of the following correctly defines heat in thermodynamics?

(A) Energy stored within the system
(B) Energy transferred due to temperature difference
(C) Energy due to molecular motion
(D) Energy stored as pressure

Q20. Which of the following correctly represents work done BY a gas during expansion against constant external pressure?

(A) $w = P_{ext}\Delta V$
(B) $w = \Delta V / P_{ext}$
(C) $w = q\Delta V$
(D) $w = -P_{ext}\Delta V$

Q21. Which of the following is a thermochemical equation?

(A) $C + O_2 \rightarrow CO_2$ only
(B) $C(s) + O_2(g) \rightarrow CO_2(g)$, $\Delta H = -393.5$ kJ mol^{-1}
(C) $H_2 + Cl_2 \rightarrow 2HCl$ only
(D) $N_2 + H_2 \rightarrow NH_3$ only

Q22. Which of the following must be specified in a thermochemical equation?

(A) Colour of reactants
(B) Physical states of all substances
(C) Time taken for reaction
(D) Catalyst used

Q23. Which of the following is the most precise statement of the Second Law of Thermodynamics?

(A) Energy of universe is constant
(B) Entropy of universe increases in every spontaneous process
(C) Entropy of system always increases
(D) Heat always converts to work completely

Q24. Which of the following correctly states the change in entropy of the universe for a spontaneous process?

(A) $\Delta S_{universe} = 0$
(B) $\Delta S_{universe} < 0$
(C) $\Delta S_{universe} > 0$
(D) $\Delta S_{universe} = \Delta S_{system}$ only

Answer key

1-B	**2-C**	**3-A**	**4-C**	**5-C**	**6-B**	**7-B**	**8-D**
9-A	**10-C**	**11-A**	**12-C**	**13-A**	**14-B**	**15-D**	**16-B**
17-A	**18-B**	**19-B**	**20-D**	**21-B**	**22-B**	**23-B**	**24-C**

Answers with explanations

Q1 — Answer: (B) −17 kcal mol^{-1}
By Hess's law: $\Delta H°f(CH_4) = \Delta H°c(C) + 2\Delta H°c(H_2) - \Delta H°c(CH_4) = -94 + 2(-68) - (-213) = -94 - 136 + 213 = -17$ kcal mol^{-1}. Option B is correct.

Q2 — Answer: (C) ΔU = w ≠ 0 and q = 0
The container is insulated, so no heat exchange occurs: q = 0. The paddle does work on the system, so w > 0. By the first law, ΔU = q + w = w > 0. Options A and D require q ≠ 0, which contradicts insulation. Option B requires w = 0, which contradicts paddle action.

Q3 — Answer: (A) 38.3 J K^{-1}
$\Delta S = nR \ln(V_2/V_1) = 2 \times 8.314 \times \ln(10) = 2 \times 8.314 \times 2.3026 \approx 38.3$ J K^{-1}. Option B uses n=1; Option C doubles incorrectly; Option D uses n=0.5.

Q4 — Answer: (C) ΔU = 0, q = −w
In a cyclic process the system returns to its initial state, so all state functions including U return to their original values: ΔU = 0. From the first law, 0 = q + w, therefore q = −w. Options A, B and D incorrectly assign zero values to q or w or non-zero values to ΔU.

Q5 — Answer: (C) 310 kJ
Molar mass of CO_2=44 g mol^{-1}. Moles=35.0/44=0.7955 mol. Heat released=0.7955×390≈310 kJ. Option B is the value for exactly 1 mol. Options A and D are numerically incorrect.

Q6 — Answer: (B) q=500 J, w=0, ΔU=500 J
At constant volume ΔV=0, so w=−PextΔV=0. By the first law ΔU=q+w=500+0=500 J. Option A confuses adiabatic with constant-volume conditions. Option C assigns non-zero work. Option D incorrectly sets ΔU=0.

Q7 — Answer: (B) q positive, w negative
IUPAC: heat absorbed by system → q positive; work done BY system on surroundings → w negative (system loses energy as work). Both signs reflect the system's perspective.

Q8 — Answer: (D) ΔU = qv
At constant volume, no PV work is done (w = 0), so first law gives ΔU = q. This heat at constant volume is denoted qv, measured directly by bomb calorimetry. Option A is only true when Δng = 0.

Q9 — Answer: (A) Energy to break one mole of bonds in gaseous state
Bond enthalpy is the energy required to break one mole of a specific type of bond in gaseous-phase molecules. It is always defined for the gaseous state, not liquid or solution. Option C is wrong because bond formation releases, not absorbs, energy.

Q10 — Answer: (C) Bond enthalpy
Bond breaking always requires energy input, making bond enthalpy always endothermic and hence always positive. Enthalpy of formation can be negative (e.g., H_2O), combustion is always negative, and neutralization of strong acid-base is always negative (−57 kJ/mol).

Q11 — Answer: (A) Measure of randomness or disorder in a system
Entropy (S) quantifies the degree of randomness or disorder; it is a thermodynamic state function. Heat content is enthalpy (H), not entropy. Internal energy is a separate state function.

Q12 — Answer: (C) ΔS = qrev / T
For a reversible process, ΔS = qrev/T, where qrev is heat exchanged reversibly and T is absolute temperature. Option A inverts the relationship. Options B and D use ΔH and ΔG, which apply to different expressions.

Q13 — Answer: (A) G = H − TS
Gibbs free energy is defined as G = H − TS, where H is enthalpy, T is absolute temperature, and S is entropy. Option C (G = U − TS) defines Helmholtz free energy (A), a common NEET-level confusion. Option D reverses the sign, giving a physically meaningless quantity.

Q14 — Answer: (B) ΔG < 0
At constant T and P, a process is spontaneous when ΔG < 0 — the system loses free energy. ΔG = 0 represents equilibrium, not spontaneity. ΔG > 0 means the reverse process is spontaneous. ΔG = ΔH would only hold if TΔS = 0.

Q15 — Answer: (D) System exchanging neither matter nor energy
An isolated system has no exchange of matter or energy with surroundings. Option B describes a closed system — the most common wrong choice here.

Q16 — Answer: (B) Closed system
A closed system permits energy exchange (heat/work) but not matter. Open systems allow both; isolated systems allow neither. Adiabatic is a process property, not a system type.

Q17 — Answer: (A) H = U + PV
Enthalpy is defined as H = U + PV, where U is internal energy and PV is the pressure-volume term. This definition accounts for the energy stored in the system plus the work done by the system against external pressure. H = q + w is the first law expression for internal energy change, not enthalpy.

Q18 — Answer: (B) Constant pressure
At constant pressure (isobaric process), no PV work other than expansion work occurs, so qp = ΔH. At constant volume, qv = ΔU. Adiabatic means q = 0, so ΔH ≠ q. This is the thermodynamic basis of coffee-cup calorimetry.

Q19 — Answer: (B) Energy transferred due to temperature difference
Heat is energy in transit between system and surroundings driven by a temperature gradient — it is not stored. Internal energy is stored; temperature measures average kinetic energy. All three are distinct concepts.

Q20 — Answer: (D) w = −PextΔV
IUPAC: w = −PextΔV for work done BY the system. During expansion ΔV > 0, making w negative — the system loses energy as work. Option A gives the magnitude without sign, causing errors in ΔU calculations.

Q21 — Answer: (B) $C(s) + O_2(g) \rightarrow CO_2(g)$, $\Delta H = -393.5\ kJ\ mol^{-1}$
A thermochemical equation must include: (1) a balanced equation, (2) physical state symbols for all species, and (3) the enthalpy change (ΔH) value. Options A, C, and D lack ΔH values and state symbols — they are ordinary chemical equations, not thermochemical equations.

Q22 — Answer: (B) Physical states of all substances
Physical states must be specified because enthalpy depends on state — $H_2O(l)$ and $H_2O(g)$ have different enthalpies differing by the enthalpy of vaporisation (≈44 kJ mol^{-1}). Changing a state symbol changes the ΔH value of the equation. Colour, time, and catalyst do not affect ΔH.

Q23 — Answer: (B) Entropy of universe increases in every spontaneous process
The Second Law states that for any spontaneous process, the total entropy of the universe (system + surroundings) increases. Option C is incorrect — entropy of the system alone can decrease (e.g., freezing of water) as long as entropy of surroundings increases more.

Q24 — Answer: (C) ΔS_universe > 0
For spontaneous processes, ΔS_universe = ΔS_system + ΔS_surroundings > 0. At equilibrium, ΔS_universe = 0. ΔS_universe < 0 is impossible for any real process (it would violate the Second Law). Option D is a trap — surroundings entropy must also be included.

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