

Thermodynamics — CBSE Pattern Test 2

Class 11 Chemistry · 90 minutes · 40 marks · Sections A–D with answer key · Narayan Gurukul Academy

Test 2 — CBSE Pattern (90 minutes, 40 marks)

Class 11 Chemistry · Thermodynamics · Narayan Gurukul Academy (ClassApna)

General Instructions: (1) This paper has four sections — Section A (1×10), Section B (2×4), Section C (3×4) and Section D (5×2) — total 40 marks. Time allowed: 90 minutes.

(2) All questions are compulsory. No overall choice is given; internal choice, where allowed, is clearly marked.

(3) For MCQ questions choose the single correct option; for VSA answers write a single line.

(4) Take $g = 10 \text{ m/s}^2$. All symbols carry their usual meaning. Use of a calculator is not permitted.

Section A — MCQ & VSA ($10 \times 1 = 10$ marks)

- Work done on the system is taken as
 - positive
 - negative
 - zero
 - infinite
- In an adiabatic process, there is no exchange of
 - heat
 - work
 - matter
 - energy
- For an isothermal expansion of an ideal gas, ΔU is
 - zero
 - positive
 - negative
 - infinite
- For a cyclic process, the change in internal energy is
 - zero
 - positive
 - negative
 - infinite
- For an exothermic reaction, ΔH is
 - negative
 - positive
 - zero
 - infinite
- The relation between ΔH and ΔU is
 - $\Delta H = \Delta U + \Delta n_g RT$
 - $\Delta H = \Delta U - \Delta n_g RT$
 - $\Delta H = \Delta U$
 - $\Delta H = \Delta n_g RT$
- The standard enthalpy of formation of an element in its standard state is
 - zero

- (B) positive
 - (C) negative
 - (D) infinite
8. A process is spontaneous at constant T and p when
- (A) ΔG is negative
 - (B) ΔG is positive
 - (C) ΔG is zero
 - (D) ΔG is infinite
9. Define enthalpy.
10. State the second law of thermodynamics.

Section B — Short Answer I ($4 \times 2 = 8$ marks)

11. What is the sign of q when heat is absorbed by the system?
12. In an adiabatic process, what does the first law reduce to?
13. What is the effect on ΔH when a reaction is reversed?
14. Define the enthalpy of combustion.

Section C — Short Answer II ($4 \times 3 = 12$ marks)

15. Define a thermodynamic system and classify systems into open, closed and isolated with examples.
16. Define enthalpy and explain how calorimetry is used to measure heat changes.
17. State Hess's law and describe one application of the law.
18. Define bond enthalpy, the enthalpy of combustion and the enthalpy of formation, and explain how bond enthalpies estimate the enthalpy of a reaction.

Section D — Long Answers ($2 \times 5 = 10$ marks)

19. State the first law of thermodynamics. Calculate ΔU when 2 mol of an ideal gas absorbs 500 J of heat and does 200 J of work, and explain why internal energy is a state function.
20. Define entropy and Gibbs energy. For a reaction with $\Delta H = +50 \text{ kJ mol}^{-1}$ and $\Delta S = +150 \text{ J K}^{-1} \text{ mol}^{-1}$, calculate ΔG at 298 K and explain how the sign of ΔG predicts spontaneity.

Answer Key (Marking Scheme)

- Q1 — (A) positive. Q2 — (A) heat. Q3 — (A) zero. Q4 — (A) zero.
- Q5 — (A) negative. Q6 — (A) $\Delta H = \Delta U + \Delta n_g RT$. Q7 — (A) zero. Q8 — (A) ΔG is negative.
- Q9 — $H = U + pV$.
- Q10 — The entropy of the universe increases in a spontaneous process.
- Q11 — Positive.
- Q12 — $\Delta U = w$ (since $q = 0$).
- Q13 — The sign of ΔH is reversed.
- Q14 — Heat change when one mole of a substance is completely burnt in oxygen.
- Q15 — Open, closed, isolated systems as in notes.
- Q16 — $H = U + pV$; $q = C \Delta T$ in calorimetry.

Q17 — Independent of path; used to find ΔH of reactions that cannot be measured directly.

Q18 — Definitions as in notes; $\Delta H = \Sigma \text{ bonds broken} - \Sigma \text{ bonds formed}$.

Q19 — $\Delta U = 500 + (-200) = 300 \text{ J}$; U is a state function because it depends only on the state.

Q20 — $\Delta G = 50 - 298(0.150) = +5.3 \text{ kJ mol}^{-1}$, non-spontaneous.