

Thermodynamics — Assignment 2

Class 11 Chemistry · Higher-difficulty challenge set + integrated problems · Answer key included · Narayan Gurukul Academy

Instructions: (1) Attempt all questions; start with the topic-wise challenge set, then attempt the integrated (multi-topic) problems.
(2) Give all answers with correct SI units. $g = 10 \text{ m/s}^2$ unless stated.
(3) A final answer key is provided at the end — try the full set before checking it.

Part A — Topic-wise challenge questions

Systems, Work and Heat

1. Distinguish between an open, a closed and an isolated system with examples.
2. Define work and heat in thermodynamics and state the sign convention used.
3. Explain why internal energy is a state function while work and heat are path functions.

First Law of Thermodynamics

1. State the first law of thermodynamics and explain each term with sign conventions.
2. Apply the first law to an isothermal expansion of an ideal gas.
3. What is the value of ΔU for a cyclic process and why?

Enthalpy and Calorimetry

1. Define enthalpy and show that the heat change at constant pressure equals ΔH .
2. Derive the relation $\Delta H = \Delta U + \Delta n_g RT$.
3. Explain the principle of calorimetry with the equation used to measure heat changes.

Hess's Law of Constant Heat Summation

1. State Hess's law and explain its basis in the first law of thermodynamics.
2. Use Hess's law to explain why ΔH for a reaction is independent of the route.
3. What happens to ΔH when a reaction is reversed and when it is multiplied by a factor?

Bond Enthalpy and Enthalpy of Combustion

1. Define bond enthalpy and the enthalpy of combustion, giving one example of each.
2. Calculate ΔH for a reaction from the bond enthalpies of the bonds broken and formed.
3. What is the standard enthalpy of formation of an element in its standard state?

Entropy and Spontaneity

1. Define entropy and state the second law of thermodynamics.
2. Define Gibbs energy and state the criterion for spontaneity at constant T and p .
3. Explain how the sign of ΔG predicts whether a process is spontaneous.

Part B — Integrated problems (multi-topic)

1. Calculate ΔU when 2 mol of an ideal gas absorbs 500 J of heat and does 200 J of work.
2. Calculate ΔH for a reaction at 298 K where $\Delta U = -100$ kJ and $\Delta n_g = -2$.
3. The bond enthalpies of H–H, Cl–Cl and H–Cl are 436, 243 and 431 kJ mol⁻¹. Calculate ΔH for $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g})$.
4. For a reaction $\Delta H = +50$ kJ mol⁻¹ and $\Delta S = +150$ J K⁻¹ mol⁻¹. Calculate ΔG at 298 K and state whether the reaction is spontaneous.
5. State the sign of ΔS for (a) melting of ice, (b) condensation of steam and (c) dissolution of ammonium nitrate in water.

Answer Key**Systems, Work and Heat**

Open: exchanges matter and energy; closed: energy only; isolated: neither. Work done on the system is positive; heat absorbed is positive. Internal energy U is a state function; w and q depend on the path.

First Law of Thermodynamics

$\Delta U = q + w$. Isothermal ideal gas: $\Delta U = 0$, so $q = -w$. Adiabatic: $q = 0$, so $\Delta U = w$. Cyclic: $\Delta U = 0$.

Enthalpy and Calorimetry

$H = U + pV$; at constant pressure $q_p = \Delta H$. $\Delta H = \Delta U + \Delta n_g RT$. Calorimetry uses $q = C \Delta T$ (or the bomb calorimeter at constant volume gives ΔU).

Hess's Law of Constant Heat Summation

The enthalpy change is independent of the path (a consequence of the first law). Reversing a reaction changes the sign of ΔH ; multiplying the reaction multiplies ΔH .

Bond Enthalpy and Enthalpy of Combustion

Bond enthalpy: energy to break one mole of bonds in the gaseous state. Enthalpy of combustion: heat on complete combustion of one mole. $\Delta H(\text{reaction}) = \sum \text{bond enthalpies broken} - \sum \text{bond enthalpies formed}$. $\Delta_f H^\circ$ of an element = 0.

Entropy and Spontaneity

S measures disorder; $\Delta S_{\text{universe}} > 0$ for a spontaneous change. $G = H - TS$; $\Delta G < 0$ spontaneous, $= 0$ equilibrium, > 0 non-spontaneous.

Integrated problems

1: $\Delta U = 500 + (-200) = 300$ J. 2: $\Delta H = \Delta U + \Delta n_g RT = -100 + (-2)(8.314 \times 10^{-3})(298) \approx -104.96$ kJ. 3: $\Delta H = (436 + 243) - (2 \times 431) = -183$ kJ mol⁻¹. 4: $\Delta G = 50 - 298(0.150) = +5.3$ kJ mol⁻¹ → non-spontaneous. 5: (a) positive, (b) negative, (c) positive.