



ChapterNotes

chapternotes.in · free NEET & JEE prep

CHAPTER NOTES

Thermodynamics Class 11 Notes I CBSE Chemistry Chapter 5

- ✓ Chapter-wise notes & important questions
- ✓ Free gamified practice app — streaks, levels, revision
- ✓ Free mock & prediction papers with answer keys

Source: <https://chapternotes.in/cbse/class-11-chemistry/thermodynamics-class-11-chemistry-notes/>



Scan to practice this
chapter free in the app

Thermodynamics Class 11 Notes | CBSE Chemistry

Chapter 5

Thermodynamics is Chapter 5 of CBSE Class 11 Chemistry — the chapter that tells you *which* reactions can happen, how much heat they release or absorb, and why. It does not worry about how fast a reaction goes (that is kinetics); instead it deals with energy bookkeeping: heat, work, internal energy, enthalpy, and the direction in which nature prefers to move. Master it and a whole family of numericals in JEE, NEET, and your board exam become almost mechanical.

By the end of these notes you will be able to define a system and its surroundings, apply the first law of thermodynamics, calculate enthalpy changes using Hess's law and bond enthalpies, and decide whether a reaction is spontaneous using entropy and Gibbs free energy. This is a high-weightage chapter carrying roughly 7–9 marks in boards, and the conceptual foundation for Chemical Equilibrium, Electrochemistry, and Thermochemistry in Class 12.

Table of Contents

- Key Concepts — System, state functions, first law, enthalpy, Hess's law, entropy, Gibbs energy
 - Weightage in Board & Entrance Exams
 - Important Definitions
 - Solved Examples
 - Important Questions for Board Exams
 - Quick Revision Points
-

Key Concepts

1. System, Surroundings and Boundary

A **system** is the part of the universe chosen for study — for example, the reactants inside a test tube. Everything else that can exchange energy or matter with it is the **surroundings**. The real or imaginary surface separating them is the **boundary**.

Types of System

- **Open system:** exchanges both matter and energy (an open beaker of hot water).
 - **Closed system:** exchanges energy but not matter (a sealed flask that can be heated).
 - **Isolated system:** exchanges neither matter nor energy (a perfect thermos flask).
-

2. State Functions and Path Functions

A **state function** depends only on the present state of the system, not on how that state was reached. Pressure (P), volume (V), temperature (T), internal energy (U), enthalpy (H), entropy (S) and Gibbs energy (G) are all state functions.

A **path function** depends on the route taken between two states. **Heat (q)** and **work (w)** are path functions — the same change in state can involve different amounts of heat and work depending on the path.

Key idea: For a state function, the change depends only on the initial and final states, e.g. $\Delta U = U_{\text{final}} - U_{\text{initial}}$.

3. Internal Energy (U)

Internal energy is the total energy stored in a system — the sum of the kinetic and potential energies of all its molecules (translational, rotational, vibrational, electronic and nuclear). It is a state function whose absolute value cannot be measured; only the change ΔU is measurable.

Internal energy of a system can be changed in two ways — by transferring **heat** or by doing **work**.

$$\Delta U = q + w$$

4. Work and Heat (Sign Conventions)

Work (w) in chemistry is usually pressure–volume work done during expansion or compression of gases. For an irreversible process against constant external pressure:

$$w = -P_{\text{ext}} \Delta V$$

For an isothermal reversible expansion of an ideal gas:

$$w = -2.303 nRT \log(V_2/V_1)$$

- **Heat (q):** energy transferred due to a temperature difference.
 - **q is +ve** when heat is absorbed by the system; **–ve** when released.
 - **w is +ve** when work is done *on* the system (compression); **–ve** when done *by* the system (expansion).
-

5. First Law of Thermodynamics

The first law is simply the **law of conservation of energy**: energy can neither be created nor destroyed, only converted from one form to another. The total energy of an isolated system is constant.

$$\Delta U = q + w$$

- If the system absorbs heat q and work w is done on it, the internal energy rises.
- For an isolated system ($q = 0$, $w = 0$), $\Delta U = 0$.

- For a cyclic process, $\Delta U = 0$, so $q = -w$.
-

6. Enthalpy (H)

Most reactions are carried out in open vessels at constant pressure, where the system can do work by expanding. To handle this conveniently we define **enthalpy**, the heat content of a system at constant pressure.

$$H = U + PV$$

At constant pressure the heat exchanged equals the enthalpy change:

$$q_p = \Delta H = \Delta U + P\Delta V$$

For reactions involving gases, $P\Delta V = \Delta n_g RT$, so:

$$\Delta H = \Delta U + \Delta n_g RT$$

where Δn_g = (moles of gaseous products – moles of gaseous reactants).

- **Exothermic reaction:** heat is released, ΔH is **negative** (combustion, neutralization).
 - **Endothermic reaction:** heat is absorbed, ΔH is **positive** (decomposition of CaCO_3).
-

7. Heat Capacity (C, C_p and C_v)

Heat capacity is the heat required to raise the temperature of a substance by 1 K: $q = C\Delta T$. **Molar heat capacity** is the heat needed per mole, and **specific heat** is the heat needed per gram.

- **C_v :** molar heat capacity at constant volume — here heat goes only into internal energy.
- **C_p :** molar heat capacity at constant pressure — extra heat is needed because the gas also does expansion work.

For an ideal gas the two are related by **Mayer's relation**:

$$C_p - C_v = R$$

Therefore $C_p > C_v$, because at constant pressure part of the supplied heat is used to do work against the surroundings.

8. Standard Enthalpy of Reaction and Standard State

The **standard state** of a substance is its pure, most stable form at 1 bar pressure and the specified temperature (usually 298 K). Enthalpy changes measured under these conditions are written with the symbol ΔH° (standard enthalpy change).

The **standard enthalpy of a reaction ($\Delta_r H^\circ$)** is the enthalpy change when reactants in their standard states convert to products in their standard states.

9. Types of Enthalpy Change

The same idea — heat exchanged at constant pressure — is given different names depending on the process.

Type	Symbol	Definition
Enthalpy of formation	$\Delta_f H^\circ$	Heat change when 1 mole of a compound forms from its elements in their standard states
Enthalpy of combustion	$\Delta_c H^\circ$	Heat released when 1 mole of a substance burns completely in oxygen (always –ve)
Enthalpy of neutralization	$\Delta_n H^\circ$	Heat released when 1 mol H^+ neutralizes 1 mol OH^- ; ≈ -57.1 kJ/mol for strong acid–strong base
Enthalpy of solution	$\Delta_{sol} H^\circ$	Heat change when 1 mole of solute dissolves in a large amount of solvent
Enthalpy of atomization	$\Delta_a H^\circ$	Heat needed to break 1 mole of a substance into gaseous atoms
Bond enthalpy	$\Delta_{bond} H^\circ$	Energy required to break 1 mole of a particular bond in the gas phase

Key point: The standard enthalpy of formation of any element in its most stable form is taken as **zero** (e.g. $\Delta_f H^\circ$ of $O_2(g)$, $H_2(g)$, $C(\text{graphite}) = 0$).

Using bond enthalpies, $\Delta_r H^\circ = \Sigma (\text{bond enthalpies of bonds broken}) - \Sigma (\text{bond enthalpies of bonds formed})$.

10. Hess's Law of Constant Heat Summation

Hess's law states that the total enthalpy change of a reaction is the same whether it takes place in one step or in several steps. This follows directly from enthalpy being a state function.

[DIAGRAM: An enthalpy cycle — reactants converting to products directly (ΔH) versus through an intermediate in two steps ($\Delta H_1 + \Delta H_2$), with $\Delta H = \Delta H_1 + \Delta H_2$.]

Hess's law lets us calculate enthalpy changes that are hard to measure directly (like $\Delta_f H^\circ$ of CO) by adding or subtracting known thermochemical equations:

$$\Delta_r H^\circ = \Delta_f H^\circ(\text{products}) - \Delta_f H^\circ(\text{reactants})$$

11. Spontaneity and Entropy (S)

A **spontaneous process** is one that occurs on its own without any continuous external help (water flowing downhill, iron rusting). Exothermicity alone does not decide spontaneity — some endothermic processes (melting of ice, dissolving of NH_4Cl) are also spontaneous.

The missing factor is **entropy (S)**, a measure of the disorder or randomness of a system. The more ways energy and particles can be arranged, the higher the entropy.

$$\Delta S = q_{\text{rev}} / T$$

- Entropy increases when solids melt, liquids vaporize, or gases are produced.
 - **Second law of thermodynamics:** the total entropy of the universe always increases in a spontaneous process:
 $\Delta S_{\text{total}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}} > 0$.
 - **Third law:** the entropy of a perfectly crystalline substance is zero at absolute zero (0 K).
-

12. Gibbs Free Energy (G) and Spontaneity

To judge spontaneity from the system alone, we combine enthalpy and entropy into the **Gibbs free energy**, the energy available to do useful work.

$$G = H - TS, \text{ and at constant T and P: } \Delta G = \Delta H - T\Delta S$$

ΔG	Nature of process
$\Delta G < 0$ (negative)	Spontaneous (feasible)
$\Delta G = 0$	System at equilibrium
$\Delta G > 0$ (positive)	Non-spontaneous (reverse is spontaneous)

Important: Both ΔH and $T\Delta S$ decide the sign of ΔG . A reaction that is exothermic ($-\Delta H$) and increases disorder ($+\Delta S$) is spontaneous at all temperatures.

ΔG and Equilibrium

The standard Gibbs energy change is linked to the equilibrium constant K by:

$$\Delta G^\circ = -2.303 RT \log K$$

So a large negative ΔG° means a large K (products favoured), while a positive ΔG° means $K < 1$ (reactants favoured).

Weightage in Board & Entrance Exams

Exam	Typical Weightage	Most-Tested Areas
CBSE Board (Class 11)	7–9 marks	First law, ΔH vs ΔU , Hess's law, $\Delta G = \Delta H - T\Delta S$
JEE Main / Advanced	2–3 questions	Work in reversible/irreversible processes, bond enthalpy, ΔG –K relation
NEET	2–3 questions	First law, enthalpy of reaction, spontaneity, entropy

[TABLE: Question-type split — VSA (1 mark): definitions, sign conventions, state vs path functions; SA (2–3 marks): $\Delta H = \Delta U + \Delta n_g RT$ numericals, Hess's law, $C_p - C_v = R$; LA (5 marks): bond-enthalpy calculations, Gibbs energy and spontaneity at different temperatures.]

Important Definitions

Term	Definition
System	The part of the universe under study; can be open, closed or isolated
State function	A property depending only on the state, not the path: U, H, S, G, P, V, T
Internal energy (U)	Total energy of a system; only ΔU is measurable: $\Delta U = q + w$
First law of thermodynamics	Energy is conserved: $\Delta U = q + w$
Enthalpy (H)	Heat content at constant pressure: $H = U + PV$; $q_p = \Delta H$
Heat capacity (C_p, C_v)	Heat per degree rise; for ideal gas $C_p - C_v = R$
Hess's law	Net ΔH is the same whether a reaction occurs in one or several steps
Bond enthalpy	Energy to break 1 mole of a bond in the gaseous state
Entropy (S)	Measure of randomness/disorder: $\Delta S = q_{rev}/T$
Gibbs free energy (G)	Energy free to do work: $\Delta G = \Delta H - T\Delta S$; $\Delta G < 0$ means spontaneous

Solved Examples

Example 1

A system absorbs 200 J of heat and does 50 J of work on the surroundings. Find ΔU .

Answer: $q = +200 \text{ J}$, $w = -50 \text{ J}$ (work done by system). $\Delta U = q + w = 200 + (-50) = \mathbf{+150 \text{ J}}$.

Example 2

For the reaction $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$ at 298 K, $\Delta U = -92.0 \text{ kJ}$. Calculate ΔH . ($R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$)

Answer: $\Delta n_g = 2 - (1 + 3) = -2$. $\Delta H = \Delta U + \Delta n_g RT = -92.0 + (-2)(8.314 \times 10^{-3})(298) = -92.0 - 4.95 = \mathbf{-96.95 \text{ kJ}}$.

Example 3

Calculate $\Delta_r H^\circ$ for $\text{C}(\text{graphite}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$ given $\Delta_f H^\circ(\text{CO}_2) = -393.5 \text{ kJ/mol}$.

Answer: $\Delta_r H^\circ = \Delta_f H^\circ(\text{CO}_2) - [\Delta_f H^\circ(\text{C}) + \Delta_f H^\circ(\text{O}_2)] = -393.5 - (0 + 0) = \mathbf{-393.5 \text{ kJ/mol}}$ (elements have $\Delta_f H^\circ = 0$).

Example 4

Two moles of an ideal gas expand isothermally and reversibly from 1 L to 10 L at 300 K. Find the work done. ($R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$)

Answer: $w = -2.303 nRT \log(V_2/V_1) = -2.303 \times 2 \times 8.314 \times 300 \times \log(10/1) = -2.303 \times 2 \times 8.314 \times 300 \times 1 = \mathbf{-11488 \text{ J} \approx -11.49 \text{ kJ}}$.

Example 5

For a reaction $\Delta H = +30 \text{ kJ/mol}$ and $\Delta S = +100 \text{ J K}^{-1} \text{ mol}^{-1}$. Find the temperature above which the reaction becomes spontaneous.

Answer: At the changeover $\Delta G = 0$, so $T = \Delta H/\Delta S = 30000/100 = \mathbf{300 \text{ K}}$. Above 300 K, $T\Delta S > \Delta H$, so $\Delta G < 0$ and the reaction is spontaneous.

Example 6

Calculate $\Delta_r H^\circ$ for $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g})$ using bond enthalpies: $\text{H-H} = 436$, $\text{Cl-Cl} = 242$, $\text{H-Cl} = 431 \text{ kJ/mol}$.

Answer: $\Delta_r H^\circ = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed}) = (436 + 242) - (2 \times 431) = 678 - 862 = \mathbf{-184 \text{ kJ/mol}}$.

Important Questions for Board Exams

1-Mark Questions (VSA)

1. Define a state function. Give two examples.
2. What is the value of ΔU for one complete cycle of a cyclic process?
3. Why is the standard enthalpy of formation of $O_2(g)$ taken as zero?
4. State the sign of ΔH for an exothermic reaction.
5. Predict the sign of ΔS when a gas condenses to a liquid.

2–3-Mark Questions (SA)

1. State the first law of thermodynamics and express it mathematically. Explain the sign conventions of q and w .
2. Derive the relation $\Delta H = \Delta U + \Delta n_g RT$ for a reaction involving gases.
3. State Hess's law and explain how it is used to calculate the enthalpy of formation of CO .
4. Show that for an ideal gas $C_p - C_v = R$, and explain why $C_p > C_v$.

5-Mark Questions (LA)

1. Explain the terms entropy and Gibbs free energy. Derive $\Delta G = \Delta H - T\Delta S$ and discuss how the signs of ΔH and ΔS decide spontaneity at low and high temperatures.
 2. Define enthalpy of formation, combustion and neutralization with one example each, and state the sign of ΔH in each case.
 3. Derive the relation $\Delta G^\circ = -2.303 RT \log K$ and explain its significance for chemical equilibrium.
-

Quick Revision Points

- System types: open (matter + energy), closed (energy only), isolated (neither)
- State functions (U, H, S, G, P, V, T) depend on state; heat and work are path functions
- First law: $\Delta U = q + w$; for a cycle $\Delta U = 0$, $q = -w$
- Enthalpy $H = U + PV$; $q_p = \Delta H$; $\Delta H = \Delta U + \Delta n_g RT$
- Exothermic: $\Delta H < 0$; endothermic: $\Delta H > 0$
- $C_p - C_v = R$ (ideal gas); $C_p > C_v$
- $\Delta_f H^\circ$ of an element in its stable form = 0
- Bond enthalpy: $\Delta_r H^\circ = \Sigma(\text{broken}) - \Sigma(\text{formed})$
- Hess's law: net ΔH is path-independent; $\Delta_r H^\circ = \Delta_f H^\circ(\text{products}) - \Delta_f H^\circ(\text{reactants})$
- Entropy $\Delta S = q_{\text{rev}}/T$; $\Delta S_{\text{universe}} > 0$ for spontaneous change (second law)

- Gibbs energy $\Delta G = \Delta H - T\Delta S$; $\Delta G < 0$ spontaneous, $= 0$ equilibrium, > 0 non-spontaneous
 - $\Delta G^\circ = -2.303 RT \log K$
-

Next Chapter: Chapter 6 — Equilibrium



Got this PDF from a friend?

Everything here is free at **chapternotes.in** — 134 chapters of notes, a gamified practice app, and full NEET mock & prediction papers.



Scan to practice this chapter free in the app