

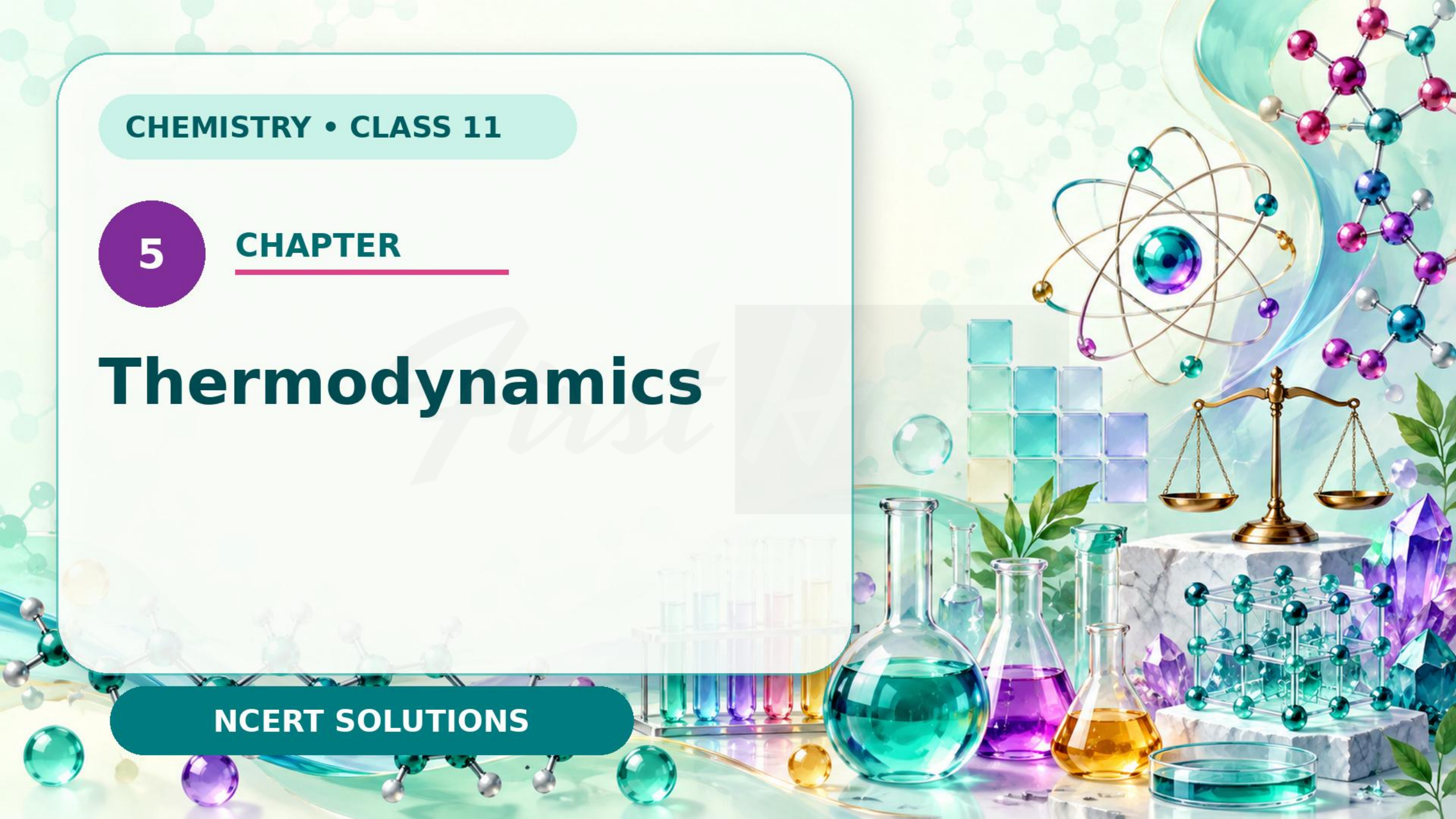
CHEMISTRY • CLASS 11

5

CHAPTER

Thermodynamics

NCERT SOLUTIONS



EXERCISE SOLUTIONS: 5.1-5.2

5.1 Choose the correct answer. A thermodynamic state function is a quantity:

- (i) used to determine heat changes
- (ii) whose value is independent of path
- (iii) used to determine pressure-volume work
- (iv) whose value depends on temperature only.

Answer: (ii) Whose value is independent of path.

A state function depends only on the state of the system. Its change depends on the initial and final states, not on the route followed.

$$\Delta U = U(\text{final}) - U(\text{initial})$$

5.2 For a process under adiabatic conditions, which condition is correct?

- (i) $\Delta T = 0$
- (ii) $\Delta p = 0$
- (iii) $q = 0$
- (iv) $w = 0$

Answer: (iii) $q = 0$

An adiabatic process has no heat exchange with the surroundings. Work can still be done, so the internal energy and temperature may change.

$$\Delta U = q + w = w$$

EXERCISE SOLUTIONS: 5.3–5.4

5.3 The enthalpies of all elements in their standard states are:

- (i) unity
- (ii) zero
- (iii) less than zero
- (iv) different for each element.

Answer: Intended answer: (ii) zero.

More precisely, the standard enthalpy of formation of an element in its reference standard state is zero by convention. This does not mean its absolute enthalpy is zero.

$$\Delta_f H^\circ (\text{element in reference state}) = 0$$

5.4 The standard internal-energy change of combustion of methane is $-X \text{ kJ mol}^{-1}$. How does ΔH° compare with ΔU° ?

- (i) $\Delta H^\circ = \Delta U^\circ$
- (ii) $\Delta H^\circ > \Delta U^\circ$
- (iii) $\Delta H^\circ < \Delta U^\circ$
- (iv) $\Delta H^\circ = 0$

Answer: (iii) $\Delta H^\circ < \Delta U^\circ$.

For complete combustion with liquid water as product:



$$\Delta n_g = 1 - (1 + 2) = -2$$

$$\begin{aligned}\Delta H^\circ &= \Delta U^\circ + \Delta n_g RT \\ &= \Delta U^\circ - 2RT\end{aligned}$$

The enthalpy change is therefore more negative.

EXERCISE SOLUTIONS: 5.5

5.5 At 298 K, the combustion enthalpies of methane, graphite and hydrogen are -890.3 , -393.5 and -285.8 kJ mol^{-1} . Find the formation enthalpy of $\text{CH}_4(\text{g})$.

(i) -74.8

(ii) -52.27

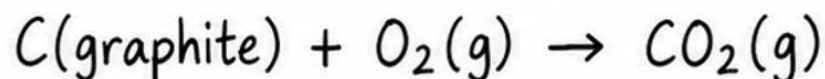
(iii) $+74.8$

(iv) $+52.26$ kJ mol^{-1}

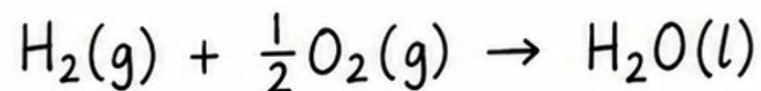
Given combustion reactions:



$$\Delta_c H^\circ = -890.3 \text{ kJ mol}^{-1}$$

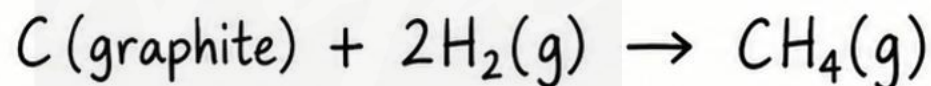


$$\Delta_c H^\circ = -393.5 \text{ kJ mol}^{-1}$$



$$\Delta_c H^\circ = -285.8 \text{ kJ mol}^{-1}$$

Answer: Use Hess's law. Add the graphite reaction and twice the hydrogen reaction, then subtract methane combustion.



$$\begin{aligned}\Delta_f H^\circ &= -393.5 + 2(-285.8) - (-890.3) \\ &= -74.8 \text{ kJ mol}^{-1}\end{aligned}$$

Correct option: (i)

Formation refers to one mole of methane made from elements in their reference states.

EXERCISE SOLUTIONS: 5.6–5.7

5.6 $A + B \rightarrow C + D + q$ has a positive entropy change. The reaction is:

- (i) possible at high temperature
- (ii) possible only at low temperature
- (iii) not possible at any temperature
- (iv) possible at any temperature.

Answer: (iv) Possible at any temperature.

Heat appears on the product side, so the reaction is exothermic. A positive entropy change also favours spontaneity.

$$\Delta H < 0; \quad \Delta S > 0$$

$$\Delta G = \Delta H - T\Delta S < 0$$

Thus it is thermodynamically spontaneous at all positive temperatures, assuming these signs remain unchanged.

5.7 A system absorbs 701 J of heat and does 394 J of work. Find its internal-energy change.

Answer: Heat absorbed is positive. Work done by the system is negative in the chemistry sign convention.

$$q = +701 \text{ J}$$

$$w = -394 \text{ J}$$

$$\Delta U = q + w$$

$$= 701 - 394$$

$$= +307 \text{ J}$$

The internal energy increases.



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