CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 5 : Thermodynamics (NCERT)

5.1 Basic Terms & System Types

Term	Definition / Detail
System	Part of the universe under study
Surroundings	Everything outside the system
Open system	Exchanges both matter & energy (e.g. open beaker)
Closed system	Exchanges only energy, not matter (e.g. sealed flask)
Isolated system	Exchanges neither matter nor energy (e.g. thermos flask)
State function	Depends only on initial & final state: U, H, S, G, T, P, V
Path function	Depends on path taken: heat (q) and work (w)
Extensive property	Depends on amount: mass, volume, U, H, S, G, heat capacity
Intensive property	Independent of amount: T, P, density, molar volume, refractive index

5.2 First Law of Thermodynamics

Concept	Formula / Expression
First law (energy conservation)	$\Delta U = q + w$
Sign convention	Heat absorbed BY system: $q = +$; Work done ON system: $w = +$
Pressure–volume work	$w = -P_{ext}\,\Delta V$ (irreversible) ; $\Delta V = V_{final} - V_{initial}$
Reversible isothermal (ideal gas)	$w = -2.303\ nRT \log(V_2/V_1) = -2.303\ nRT \log(P_1/P_2)$
Free expansion (into vacuum)	$P_{ext} = 0 \rightarrow w = 0$
Isothermal ideal gas	$\Delta U = 0$ and $\Delta T = 0 \rightarrow q = -w$
Adiabatic process	$q = 0 \rightarrow \Delta U = w$ (temperature changes)
Isochoric (constant V)	$\Delta V = 0 \rightarrow w = 0$, so $q_v = \Delta U$
Isobaric (constant P)	$q_p = \Delta H$

5.3 Enthalpy (H) & Heat Capacity

Concept	Formula / Expression
Enthalpy definition	$H = U + PV$; $\Delta H = \Delta U + P\Delta V$
For gaseous reactions	$\Delta H = \Delta U + \Delta n_g RT$; Δn_g = moles of gaseous products − gaseous reactants
Heat at constant V	$q_v = \Delta U$; Heat at constant P: $q_p = \Delta H$
Specific heat	$q = m \times c \times \Delta T$ (c = specific heat capacity, J g^{-1} K^{-1})
Molar heat capacity	$q = n \times C \times \Delta T$ (C = molar heat capacity, J mol^{-1} K^{-1})
Relation (ideal gas)	$C_p - C_v = R$ (per mole)
Heat capacity ratio	$\gamma = C_p/C_v$ (monatomic 1.67, diatomic 1.40)

Concept	Formula / Expression
C_v values	Monatomic: (3/2)R ; Diatomic: (5/2)R

5.4 Enthalpies of Reactions (Thermochemistry)

Type of Enthalpy	Definition
Enthalpy of reaction ($\Delta_r H$)	Heat change when reaction occurs in molar quantities
Standard conditions	Standard state: 1 bar pressure, specified T (usually 298 K)
Enthalpy of formation ($\Delta_f H°$)	Heat change forming 1 mol compound from elements in standard states (elements: $\Delta_f H° = 0$)
Enthalpy of combustion ($\Delta_c H°$)	Heat released on complete combustion of 1 mol substance (always negative)
Enthalpy of fusion ($\Delta_{fus} H°$)	Heat to melt 1 mol solid → liquid at m.p.
Enthalpy of vaporisation ($\Delta_{vap} H°$)	Heat to vaporise 1 mol liquid → gas at b.p.
Enthalpy of sublimation ($\Delta_{sub} H°$)	$\Delta_{sub}H = \Delta_{fus}H + \Delta_{vap}H$
Enthalpy of atomisation ($\Delta_a H°$)	Heat to break 1 mol substance into gaseous atoms
Bond enthalpy	Energy to break 1 mol of a particular bond in gaseous state
Lattice / Solution / Dilution	$\Delta_{lattice}H$, $\Delta_{sol}H$, $\Delta_{dil}H$ for ionic solids & solutions

5.5 Hess's Law & Enthalpy Calculations

Concept	Formula
Hess's law of constant heat summation	Total ΔH is the same whether reaction occurs in one step or several steps (since H is a state function)
From formation enthalpies	$\Delta_r H° = \Sigma\, \Delta_f H°(\text{products}) - \Sigma\, \Delta_f H°(\text{reactants})$
From bond enthalpies	$\Delta_r H = \Sigma$ B.E.(bonds broken, reactants) − Σ B.E.(bonds formed, products)
From combustion enthalpies	$\Delta_r H° = \Sigma\, \Delta_c H°(\text{reactants}) - \Sigma\, \Delta_c H°(\text{products})$

5.6 Spontaneity, Entropy & Gibbs Energy

Concept	Formula / Expression
Entropy (S)	Measure of disorder/randomness; a state function. $\Delta S = q_{rev}/T$
Entropy of surroundings	$\Delta S_{surr} = -\Delta H_{sys}/T$
Second law (total entropy)	For a spontaneous process: $\Delta S_{total} = \Delta S_{sys} + \Delta S_{surr} > 0$
At phase equilibrium	$\Delta S = \Delta H/T$ (e.g. melting: $\Delta S_{fus} = \Delta_{fus}H / T_m$)
Gibbs energy (definition)	$G = H - TS$; $\Delta G = \Delta H - T\Delta S$ (at constant T, P)
Gibbs–Helmholtz criterion	$\Delta G < 0$ → spontaneous ; $\Delta G = 0$ → equilibrium ; $\Delta G > 0$ → non-spontaneous
ΔG and equilibrium constant	$\Delta G° = -RT \ln K = -2.303\, RT \log K$
ΔG and useful work	$-\Delta G$ = maximum non-expansion (useful) work obtainable
Third law of thermodynamics	Entropy of a perfectly crystalline substance is zero at 0 K

5.7 Spontaneity from Signs of ΔH and ΔS

ΔH	ΔS	ΔG = ΔH − TΔS	Spontaneity
− (exo)	+	Always −	Spontaneous at ALL temperatures
+ (endo)	−	Always +	Non-spontaneous at ALL temperatures
− (exo)	−	− at low T	Spontaneous only at LOW temperature
+ (endo)	+	− at high T	Spontaneous only at HIGH temperature

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First law signs	'Heat IN is +, Work ON is +' → ΔU = q + w (chemistry convention).
ΔH vs ΔU	$\Delta H = \Delta U + \Delta n_g RT$ — only GASEOUS moles count in Δn_g. If $\Delta n_g = 0$, ΔH = ΔU.
State vs path	State functions = 'destination only' (U,H,S,G,T,P,V). Path functions = q and w.
Cp − Cv = R	For ideal gas only. Cp is always greater (extra energy for expansion work).
Spontaneity grid	'−ΔH and +ΔS = always spontaneous; +ΔH and −ΔS = never.'
ΔG = ΔH − TΔS	Negative ΔG drives reactions. 'Low T favours enthalpy, high T favours entropy.'
ΔG° = −2.303 RT log K	K > 1 → ΔG° negative (products favoured) ; K < 1 → ΔG° positive.
Δsub = Δfus + Δvap	Solid → gas in two imaginary steps (Hess's law shortcut).