CHEMICAL THERMODYNAMICS
FORMULA SHEET (NEET)

Dev Chemistry ♡

1. BASIC CONCEPTS

- System : Part of universe under study
- Surroundings : rest of the universe
- Types of system :
  - Open – exchange of matter & energy
  - Closed – exchange only energy
  - Isolated – no exchange of matter/energy
- State functions : U, H, S, G
  (depend only on initial & final state)
- Path functions : q, w
  (depend on path)

2. FIRST LAW OF THERMODYNAMICS

$$\Delta U = q + w$$

- q : heat absorbed by system (+)
- w : work done on system (+)
  (work done by system is −ve)

For only P–V work : $w = -P_{ext}\,\Delta V$

For reversible process : $w = -\int_{V_1}^{V_2} P\,dV$

3. WORK DONE (IDEAL GAS)

(i) Isobaric : $w = -P(V_2 - V_1)$

(ii) Isochoric : $w = 0$

(iii) Isothermal (reversible) :

$$w = -nRT\ln\left(\frac{V_2}{V_1}\right) = -nRT\ln\left(\frac{P_1}{P_2}\right)$$

(iv) Adiabatic (reversible) :

$$w = \frac{P_2V_2 - P_1V_1}{\gamma - 1} \qquad \gamma = \frac{C_p}{C_v}$$

4. INTERNAL ENERGY (U)

For ideal gas : $\Delta U = nC_v\,\Delta T$

Molar heat capacity (C_v) :

- Monatomic : $C_v = \frac{3}{2}R$
- Diatomic : $C_v = \frac{5}{2}R$
- Polyatomic : $C_v = 3R$

$$C_p = C_v + R$$

In cyclic process : $\Delta U = 0$

In free expansion (ideal gas) :

$$\Delta U = 0,\quad w = 0,\quad q = 0$$

5. ENTHALPY (H)

$$H = U + PV$$

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

At constant pressure : $q_p = \Delta H$

At constant volume : $q_v = \Delta U$

Relation with heat capacities :

$$C_p - C_v = R$$

6. STANDARD ENTHALPY CHANGES

(i) Formation : ΔH_f° (from elements)

(ii) Combustion : ΔH_c° (complete combustion)

(iii) Neutralisation : ΔH°_{neut}
(strong acid + strong base)
$= -57.1$ kJ mol^{-1}

(iv) Lattice enthalpy : ΔH_{latt} (formation from gaseous ions)

(v) Hydration enthalpy : ΔH_{hyd} (gaseous ion → aqueous)

(vi) Hess's Law : Enthalpy change is path independent.

7. GIBBS ENERGY & SPONTANEITY

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

- $\Delta G < 0$: spontaneous
- $\Delta G = 0$: at equilibrium
- $\Delta G > 0$: non-spontaneous
- At equilibriom : $\Delta G = 0,\ K = 1$

$$\Delta G^\circ = -RT\ln K$$

$R = 8.314$ J mol^{-1} K^{-1}
(or 0.008314 kJ mol^{-1} K^{-1})

8. ENTROPY (S)

$$\Delta S = \frac{q_{rev}}{T}$$ Units : JK^{-1}mol^{-1}

(i) Phase changes :

$$\Delta S_{fus} = \frac{\Delta H_{fus}}{T} \qquad \Delta S_{vap} = \frac{\Delta H_{vap}}{T}$$

(ii) Ideal gas (isothermal) :

$$\Delta S = nR\ln\left(\frac{V_2}{V_1}\right) = nR\ln\left(\frac{P_1}{P_2}\right)$$

(iii) At constant T : $\Delta S = nC\ln\left(\frac{T_2}{T_1}\right)$

(iv) Third law : S = 0 for perfectly crystaline substance at 0 K

9. DIFFERENT PROCESSES (IDEAL GAS)

Process	Condition	q	w	ΔU	ΔH
Isothermal	T = const	$q = -w$	$w = -nRT\ln\left(\frac{V_2}{V_1}\right)$	0	0
Isochoric	V = const	$q = nC_v\Delta T$	0	$nC_v\Delta T$	$nC_p\Delta T$
Isobaric	P = const	$q = nC_p\Delta T$	$w = -P\Delta V$	$nC_v\Delta T$	$nC_p\Delta T$
Adiabatic (reversible)		$q = 0$	$w = \frac{P_2V_2 - P_1V_1}{\gamma - 1}$	$nC_v\Delta T$	$nC_p\Delta T$
Free expansion (ideal gas)		$q = 0$	$w = 0$	$\Delta U = 0$	$\Delta H = 0$

10. USEFUL RELATIONS & CONSTANTS

- $\Delta H = \Delta U + \Delta n_g RT$
- $C_p = C_v + R$
- $\gamma = \frac{C_p}{C_v}$
- $PV = nRT$ (ideal gas)
- $nR = 2$ cal K^{-1} mol^{-1} $= 8.314$ J K^{-1} mol^{-1}
- 1 cal $= 4.184$ J
- $\Delta G^\circ = -RT\ln K$
- $\ln K = -\frac{\Delta G^\circ}{RT}$
- For isothermal (rev.) : $w = -\int P\,dV$

11. CLAUSIUS–CLAPEYRON EQUATION

$$\ln\left(\frac{P_2}{P_1}\right) = -\frac{\Delta H_{vap}}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

For small temperature range :

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{vap}}{R} \times \frac{T_2 - T_1}{T_1^2}$$

12. GIBBS–HELMHOLTZ EQUATION

$$\left(\frac{\partial(\Delta G)}{\partial T}\right)_P = -\frac{\Delta H}{T^2}$$

If ΔH is constant :

$$\frac{d(\Delta G)}{dT} = -\frac{\Delta H}{T^2} \qquad \Delta G = \Delta H - T\left(\frac{d(\Delta G)}{dT}\right)_P$$

13. ENTROPY & SPONTANEITY

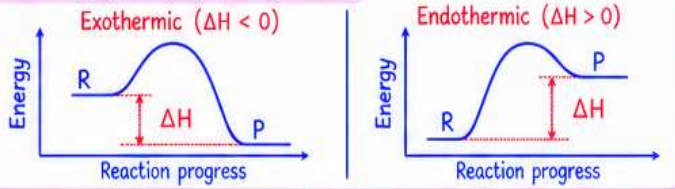
- $\Delta S_{total} = \Delta S_{system} + \Delta S_{surroundings}$
- Spontaneous process : $\Delta S_{total} > 0$
- At equilibrium : $\Delta S_{total} = 0$
- Non-spontaneous : $\Delta S_{total} < 0$
- For surroundings (at constant T) :

$$\Delta S_{surr} = -\frac{q_{system}}{T}$$

14. PHASE CHANGES (1 mol)

- Fusion : $\Delta S = \frac{\Delta H_{fus}}{T_m}$
- Vapourisation : $\Delta S = \frac{\Delta H_{vap}}{T_b}$
- Sublimation : $\Delta S = \frac{\Delta H_{sub}}{T}$
- $\Delta S_{vap} > \Delta S_{fus}$
- $\Delta S_{sub} = \Delta S_{fus} + \Delta S_{vap}$

15. ENERGY PROFILE DIAGRAMS



16. SIGNS (QUICK CHECK)

Quantity	Endothermic	Exothermic
q (system)	+	−
ΔH	+	−
w (on system)	+	+
ΔS (system)	+	−
ΔG	−	+

17. IMPORTANT POINTS FOR PYQS

1. Use correct sign convention for q and w.
2. For ideal gas, ΔU and ΔH depend only on temperature.
3. For cyclic process, $\Delta U = 0$ and $w_{total} = -q_{total}$.
4. Free expansion (ideal gas) : $\Delta U = 0,\ w = 0,\ q = 0,\ \Delta S = nR\ln\left(\frac{V_2}{V_1}\right)$.

18. COMMON VALUES

- $R = 8.314$ J mol^{-1} K^{-1} $= 0.008314$ kJ mol^{-1}K^{-1}
  $= 2$ cal K^{-1} mol^{-1}
  $= $ cal $= 4.184$ J

19. THERMODYNAMIC IDENTITIES

- $dU = TdS - PdV$
- $dH = TdS + VdP$