TURBO

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Thermodynamics & Chemical Energetics

Class 11 Chemistry ● Complete Formula Sheet

Sr.	Concept	Formulas	Other Information
**Basic Terminology & Systems**			
1	System Types	**Open:** Exchange matter + energy **Closed:** Exchange energy only **Isolated:** No exchange	**Examples:** Open: Beaker Closed: Sealed flask Isolated: Thermos flask
2	State Functions	Properties: U, H, S, G, P, T, V Independent of path	**Key point:** Value depends only on initial and final state, not on the process
3	Extensive vs Intensive Properties	**Extensive:** Mass, Volume, Energy, Heat capacity (depend on amount) **Intensive:** Density, T, P, Concentration (independent of amount)	**Remember:** Extensive $\div$ Extensive = Intensive Example: Mass/Volume = Density
**Types of Processes**			
4	Isothermal Process	$\Delta T = 0$ Temperature constant	**Condition:** Process carried out at constant temperature Example: Melting ice at 0°C
5	Adiabatic Process	$q = 0$ No heat exchange	**Condition:** System thermally insulated Fast processes are usually adiabatic
6	Isobaric Process	$\Delta P = 0$ Pressure constant	**Condition:** Open system or constant external pressure
7	Isochoric Process	$\Delta V = 0$ Volume constant	**Condition:** Rigid container, no expansion/compression
8	Cyclic Process	Initial state = Final state $\Delta U = 0$ for cycle	**Key point:** All state functions return to original values
**First Law of Thermodynamics**			
9	First Law	$\Delta U = q + w$ Energy is conserved	**Sign conventions:** Heat absorbed: $q > 0$ Heat released: $q < 0$ Work on system: $w > 0$ Work by system: $w < 0$
10	Work Done (General)	$w = -\int_{V_1}^{V_2} P_{\text{ext}} dV$	**Negative sign:** Work done by system is taken as negative
11	Work: Isochoric	$\Delta V = 0$ $w = 0$ $Q_V = \Delta U$	**How to use:** At constant volume, all heat goes to change internal energy

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12	Work: Isobaric	$w = -P_{\text{ext}}\Delta V$ $Q_P = \Delta U + P_{\text{ext}}\Delta V$	**For reversible isobaric:** $w = -P_{\text{gas}}\Delta V$
13	Work: Isothermal (Irreversible)	$w = -P_{\text{ext}}\Delta V$ $\Delta U = 0$ (for ideal gas)	**How to use:** Use when external pressure is constant
14	Work: Isothermal (Reversible)	$w = -2.303\,nRT \log \frac{V_2}{V_1}$ $w = -2.303\,nRT \log \frac{P_1}{P_2}$ $\Delta U = 0$ (ideal gas)	**How to use:** For ideal gas isothermal expansion $V_2 > V_1 \Rightarrow w < 0$ (work by system) $P_1 > P_2$ for expansion
15	Work: Adiabatic Reversible	$w = \frac{nR(T_2 - T_1)}{\gamma - 1} = \frac{P_2 V_2 - P_1 V_1}{\gamma - 1}$	**For adiabatic:** $q = 0$, so $\Delta U = w$ $\gamma = \frac{C_P}{C_V}$

Enthalpy

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16	Enthalpy Definition	$H = U + PV$ $\Delta H = \Delta U + \Delta(PV)$	**Properties:** State function Extensive property
17	Enthalpy: Constant P	$\Delta H = \Delta U + P\Delta V$ $Q_P = \Delta H$	**How to use:** At constant P, heat absorbed = enthalpy change
18	Enthalpy: Constant V	$\Delta H = \Delta U + V\Delta P$	**How to use:** When volume is fixed
19	Enthalpy: For Reactions	$\Delta H = \Delta U + \Delta n_g RT$ where $\Delta n_g =$ moles of gaseous products $-$ moles of gaseous reactants	**Example:** $N_2 + 3H_2 \rightarrow 2NH_3$ $\Delta n_g = 2 - 4 = -2$ $\Delta H = \Delta U - 2RT$

Heat Capacity

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20	Heat Capacity	$Q = C \cdot \Delta T$ $C =$ Heat capacity	**Unit:** J/K or cal/K Extensive property
21	Specific Heat	$Q = m \cdot S \cdot \Delta T$ $S =$ Specific heat capacity	**Unit:** J/(g·K) Intensive property
22	Molar Heat Capacity	$C_m = S \times M$ $M =$ Molar mass	**Unit:** J/(mol·K) Intensive property

Standard Enthalpies

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23	Enthalpy of Formation (ΔH_f°)	Formation of 1 mole from elements in standard state	**Key point:** ΔH_f° of elements in standard state $= 0$ Example: ΔH_f° (C graphite) $= 0$
24	Enthalpy of Combustion (ΔH_c°)	Complete combustion of 1 mole Example: $C_6H_6(l) + \frac{15}{2}O_2(g) \rightarrow 6CO_2(g) + 3H_2O(l)$ $\Delta H_c^\circ = -3268$ kJ/mol	**How to use:** Always exothermic $\Delta H_c < 0$
25	Enthalpy of Neutralization ($\Delta H_{\text{neu}}^\circ$)	$H^+ + OH^- \rightarrow H_2O$ $\Delta H_{\text{neu}}^\circ = -57.1$ kJ/mol	**For weak acids/bases:** $\Delta H_{\text{neu}} = -57.1 + \Delta H_{\text{ion(acid)}} + \Delta H_{\text{ion(base)}}$ kJ
26	Enthalpy of Fusion (ΔH_{fus})	Melting 1 mole solid at m.p. $H_2O(s, 273K) \rightarrow H_2O(l, 273K)$ $\Delta H_{\text{fus}}^\circ = 6.0$ kJ/mol	**Remember:** Endothermic process Ice to water

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27	Enthalpy of Vaporization (ΔH_{vap})	Boiling 1 mole liquid at b.p. $H_2O(l, 373K) \rightarrow H_2O(g, 373K)$ $\Delta H^\circ_{vap} = 40.7$ kJ/mol	**Remember:** Endothermic process Water to steam
28	Enthalpy of Sublimation (ΔH_{sub})	Solid to gas directly $I_2(s) \rightarrow I_2(g)$ $\Delta H^\circ_{sub} = 62.4$ kJ/mol	**Relation:** $\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$
29	Enthalpy of Solution (ΔH_{sol})	Dissolving 1 mole solute $KCl(s) + aq \rightarrow KCl(aq)$ $\Delta H = -4.4$ kcal	**Can be:** Exothermic or endothermic

Hess's Law & Calculations

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30	Hess's Law	ΔH_r is independent of path $\Delta H_r = \Delta H_1 + \Delta H_2 + \ldots$	**How to use:** Add/subtract equations and their ΔH values Enthalpy is state function
31	Calculating ΔH_r from ΔH_f	$\Delta H^\circ_r = \sum \Delta H^\circ_f(\text{products}) - \sum \Delta H^\circ_f(\text{reactants})$	**Example:** $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ $\Delta H_r = [\Delta H_f(CO_2) + 2\Delta H_f(H_2O)] - [\Delta H_f(CH_4)]$
32	Calculating ΔH_r from ΔH_c	$\Delta H_r = \sum \Delta H_c(\text{reactants}) - \sum \Delta H_c(\text{products})$	**Note:** Opposite to formation enthalpies
33	Kirchoff's Equation	$\Delta H_{T_2} = \Delta H_{T_1} + \Delta C_P(T_2 - T_1)$ $\frac{\Delta H_{T_2} - \Delta H_{T_1}}{T_2 - T_1} = \Delta C_P$	**How to use:** To find ΔH at different temperatures $\Delta C_P = \sum C_P(\text{products}) - \sum C_P(\text{reactants})$

Bond Energy

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34	Bond Dissociation Energy	Energy to break 1 mole of bonds in gaseous state $H\text{-}H(g) \rightarrow 2H(g)$, $\Delta H = +433$ kJ/mol	**Always positive:** Bond breaking requires energy (endothermic)
35	Average Bond Energy	For polyatomic molecules: Bond energy = Average of all similar bond dissociation energies Example: O-H in H_2O $\frac{497.8+428.5}{2} = 463.15$ kJ/mol	**Two steps:** $H_2O \rightarrow H + OH$ (+497.8) $OH \rightarrow H + O$ (+428.5)
36	ΔH_r from Bond Energies	$\Delta H_r = \sum \text{B.E.(reactants)} - \sum \text{B.E.(products)}$ = Energy to break bonds − Energy released in forming bonds	**How to use:** Count all bonds broken and formed Breaking bonds: positive Forming bonds: negative
37	Resonance Energy	Resonance Energy = Experimental ΔH_f − Calculated ΔH_f (from bond energies)	**Key point:** Resonance stabilizes molecules Experimental value is more negative

Second Law & Entropy

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38	Entropy (S)	Measure of randomness/disorder State function, Extensive property Unit: J/(mol·K)	**Trends:** $S_{gas} > S_{liquid} > S_{solid}$ S increases with temperature
39	Entropy Change (Constant T)	For isothermal reversible: $\Delta S = \frac{Q_{rev}}{T}$	**How to use:** For reversible process only Irreversible: $\Delta S > \frac{Q}{T}$
40	Entropy: Isobaric Process	$\Delta S = 2.303\, nC_P \log \frac{T_2}{T_1}$ or $\Delta S = 2.303\, nRT \log \frac{P_1}{P_2}$	**How to use:** When pressure is constant Choose formula based on given data

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41	Entropy: Isochoric Process	$\Delta S = 2.303\,nC_V \log \frac{T_2}{T_1}$ or $\Delta S = 2.303\,nRT \log \frac{V_2}{V_1}$	**How to use:** When volume is constant
42	Second Law	For spontaneous process: $\Delta S_{\text{total}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}} > 0$ At equilibrium: $\Delta S_{\text{total}} = 0$	**Key point:** Entropy of universe always increases for spontaneous processes

Gibbs Free Energy

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43	Gibbs Free Energy	$G = H - TS$ $\Delta G = \Delta H - T\Delta S$	**Properties:** State function Extensive property
44	Spontaneity Criteria	**At constant T, P:** $\Delta G < 0$: Spontaneous $\Delta G = 0$: Equilibrium $\Delta G > 0$: Non-spontaneous	**Four cases:** (1) $\Delta H < 0$, $\Delta S > 0$: Always spontaneous (2) $\Delta H > 0$, $\Delta S < 0$: Never spontaneous (3) $\Delta H < 0$, $\Delta S < 0$: Spontaneous at low T (4) $\Delta H > 0$, $\Delta S > 0$: Spontaneous at high T
45	Standard Free Energy	$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ $\Delta G^\circ = \sum \Delta G_f^\circ(\text{products}) - \sum \Delta G_f^\circ(\text{reactants})$	**Note:** ΔG_f° of elements = 0
46	Free Energy and Equilibrium	$\Delta G = \Delta G^\circ + RT \ln Q$ At equilibrium: $\Delta G = 0$ $\Delta G^\circ = -RT \ln K = -2.303RT \log K$	**How to use:** If $K > 1$: $\Delta G^\circ < 0$ (spontaneous) If $K < 1$: $\Delta G^\circ > 0$ (non-spontaneous) If $K = 1$: $\Delta G^\circ = 0$ (equilibrium)
47	Free Energy and Cell Potential	$\Delta G = -nFE_{\text{cell}}$ $\Delta G^\circ = -nFE_{\text{cell}}^\circ$	**Where:** n = number of electrons $F = 96500$ C/mol E_{cell} = cell potential
48	Temperature and Equilibrium	At equilibrium: $\Delta G = 0$ $\Delta H = T\Delta S$ $T = \frac{\Delta H}{\Delta S}$	**How to use:** Find temperature at which process is at equilibrium

Third Law & Miscellaneous

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49	Third Law	At absolute zero (0 K), entropy of perfect crystalline solid = 0 $S = S_T - S_0$ where $S_0 = 0$	**Application:** Calculate absolute entropy at any temperature
50	Coupled Reactions	Non-spontaneous reaction ($\Delta G > 0$) can be made spontaneous by coupling with highly spontaneous reaction ($\Delta G << 0$) $\Delta G_{\text{total}} = \Delta G_1 + \Delta G_2 < 0$	**Example:** Glucose + Fructose + ATP $\rightarrow$ Sucrose + ADP $\Delta G = -8$ kJ/mol
51	Clausius-Clapeyron Equation	$\log \frac{P_2}{P_1} = \frac{\Delta H_{\text{vap}}}{2.303R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$	**How to use:** Relates vapor pressure to temperature Used for phase transitions

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