IAL Chem Unit 4 Entropy and Energetics Cheat Sheet

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Key Terms and Definitions

Term	Definition
Entropy (ΔS)	A measure of the disorder in a system. Units: $J\ K^{-1}\ mol^{-1}$
Standard Entropy Change (ΔS^o)	$\Delta S^o = \Sigma S^o(\text{products}) - \Sigma S^o(\text{reactants})$
Spontaneous Process	A process where $\Delta G < 0$; can occur without external input.
Gibbs Free Energy (ΔG)	The energy available to do work: $\Delta G = \Delta H - T\Delta S$
Lattice Enthalpy of Formation (ΔH^o_{LE})	Enthalpy change when 1 mole of an ionic solid forms from its gaseous ions (exothermic).
Lattice Enthalpy of Dissociation	Enthalpy change when 1 mole of ionic solid dissociates into gaseous ions (endothermic).
Atomisation Enthalpy (ΔH^o_{at})	Enthalpy change to form 1 mole of gaseous atoms from the element.
Electron Affinity (ΔH^o_{ea})	Enthalpy change when 1 mol of electrons is added to 1 mole of gaseous atoms to form 1^- ions.
Hydration Enthalpy (ΔH^o_{hyd})	Enthalpy change when gaseous ions dissolve in water to form aqueous ions.

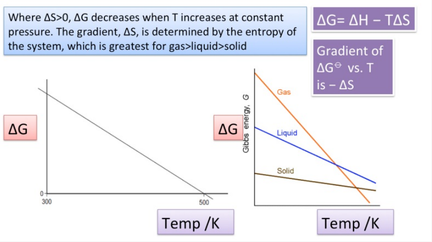
Key Terms and Definitions (cont)

Solution Enthalpy (ΔH^o_{sol})	Enthalpy change when 1 mol of solute dissolves to infinite dilution.
Perfect Ionic Model	Assumes ions are spherical and charge is evenly distributed with no covalent character.

Entropy Effect

Change	Entropy Effect
Melting / Boiling	Entropy increases significantly
Dissolving ionic lattice	Entropy increases
Increase in gas moles in products	Entropy increases
Cooling a gas	Entropy decreases
Formation of solid from gas	Entropy decreases significantly

Gibbs Free Energy vs. Temperature



Where $\Delta S>0$, ΔG decreases when T increases at constant pressure. The gradient, ΔS, is determined by the entropy of the system, which is greatest for gas>liquid>solid

Equation: $\Delta G = -\Delta S \cdot T + \Delta H$ ($y = mx + c$)

Slope = $-\Delta S$

Y-intercept = ΔH

Key Formulas

✅ Overall Entropy of a system $\Delta S_{total} = \Delta S_{system} + \Delta S_{surroundings}$

✅ Entropy of surroundings $\Delta S_{surroundings} = -\Delta H/T$

✅ Entropy Change of a reaction within a system

$\Delta S^\circ_{system} = \sum S^\circ_{products} - \sum S^\circ_{reactants}$

If entropy change is +ve, products more disordered than reactants (natural direction of change)

& if entropy change is -ve, reactants more disordered than products.

✅ Gibbs Free Energy $\Delta G = \Delta H - T\Delta S_{system}$

- ΔG: Gibbs free energy ($kJ \cdot mol^{-1}$)
- ΔH: Enthalpy change ($kJ \cdot mol^{-1}$)
- T: Temperature (K)
- ΔS: Entropy change ($J \cdot K^{-1} \cdot mol^{-1}$; convert to kJ by dividing by 1000)

✅ Feasibility Conditions

A reaction is feasible when: $\Delta G < 0$

To find minimum temperature for spontaneity: $\Delta G = 0 \Rightarrow T = \Delta H/\Delta S_{system}$

✅ Born-Haber Cycle Equation

For lattice formation enthalpy:

$\Delta H^\circ_f = \sum$enthalpies (atomisation, ionisation, electron affinity, etc.) + ΔH_{LE}

✅ Hydration & Solution Enthalpy (Hess's Cycle) $\Delta H^\circ_{sol} = \sum \Delta H^\circ - \Delta H^\circ_{LE}$

Factors Affecting Lattice & Hydration Enthalpies

Factor	Effect on ΔH (more negative)
Smaller ionic radius	Increases attraction → More negative ΔH
Higher ionic charge	Increases attraction → More negative ΔH
Higher ionic charge Increases attraction → More negative ΔH Greater hydration energy	Stronger interactions with water → More negative ΔH

Kinetics vs. Thermodynamics

A reaction may be thermodynamically feasible ($\Delta G < 0$) but may not occur due to:

- High activation energy
- Slow reaction rate due to kinetic barriers