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Introduction

Overview

- Thermodynamics studies energy forms and transformations, mainly heat to work and vice versa.
- Laws apply only when system is in equilibrium or moves between equilibrium states.
- Chemical thermodynamics focuses on chemical changes and substances.

Solved Examples

- **Ex 1:** Identify the branch of thermodynamics studying chemical reactions only.
  Answer: Chemical thermodynamics.

Practice Set

- Q1: Thermodynamics mainly studies:
  A. Motion of particles
  B. Energy transformations
  C. Chemical bonding
  D. Atomic structure
- Q2: Laws of thermodynamics apply when:
  A. System is in equilibrium
  B. System is non-equilibrium
  C. System is isolated only
  D. System is open only
- Q3: Chemical thermodynamics deals with:
  A. Physical changes
  B. Chemical changes
  C. Nuclear changes
  D. Mechanical changes

Answer Key

- Q1: B - Thermodynamics studies energy transformations.
- Q2: A - Laws apply only at equilibrium states.
- Q3: B - Chemical thermodynamics studies chemical changes.

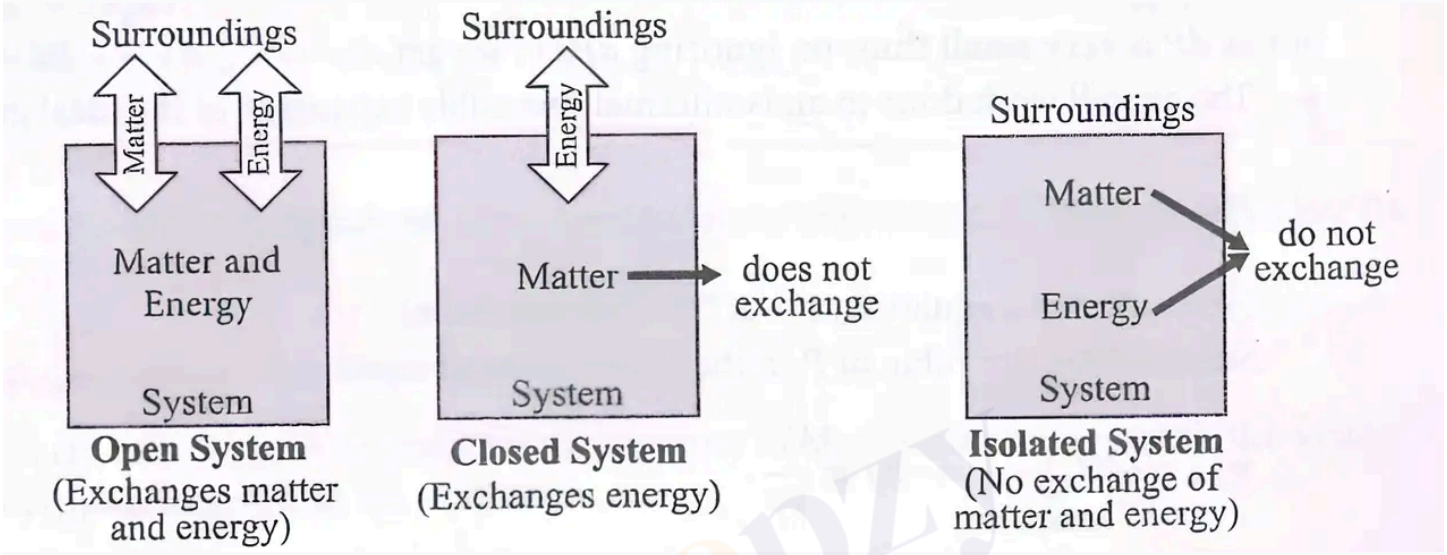
Memory Aids

- Thermodynamics = Energy + Heat + Work
- Equilibrium needed for laws to apply
- Chemical thermodynamics = chemical reactions focus

Thermodynamic Terms

Key Definitions

- System: Part of universe under study; rest is surroundings.
- Boundary: Real or imaginary surface separating system and surroundings.
- Types of systems:
  - Open: Exchange of matter and energy.
  - Closed: Exchange of energy only.
  - Isolated: No exchange of matter or energy.



Solved Examples

- **Ex 1:** Classify a system that exchanges only energy but not matter.
  Answer: Closed system.

Practice Set

- Q1: Which system exchanges both matter and energy?
  A. Open
  B. Closed
  C. Isolated
  D. None
- Q2: A system that exchanges no matter or energy is:
  A. Open
  B. Closed
  C. Isolated
  D. None
- Q3: Boundary is:
  A. Part of surroundings
  B. Surface separating system and surroundings
  C. System itself
  D. Energy flow path

Answer Key

- Q1: A - Open system exchanges both matter and energy.
- Q2: C - Isolated system exchanges neither matter nor energy.
- Q3: B - Boundary separates system and surroundings.

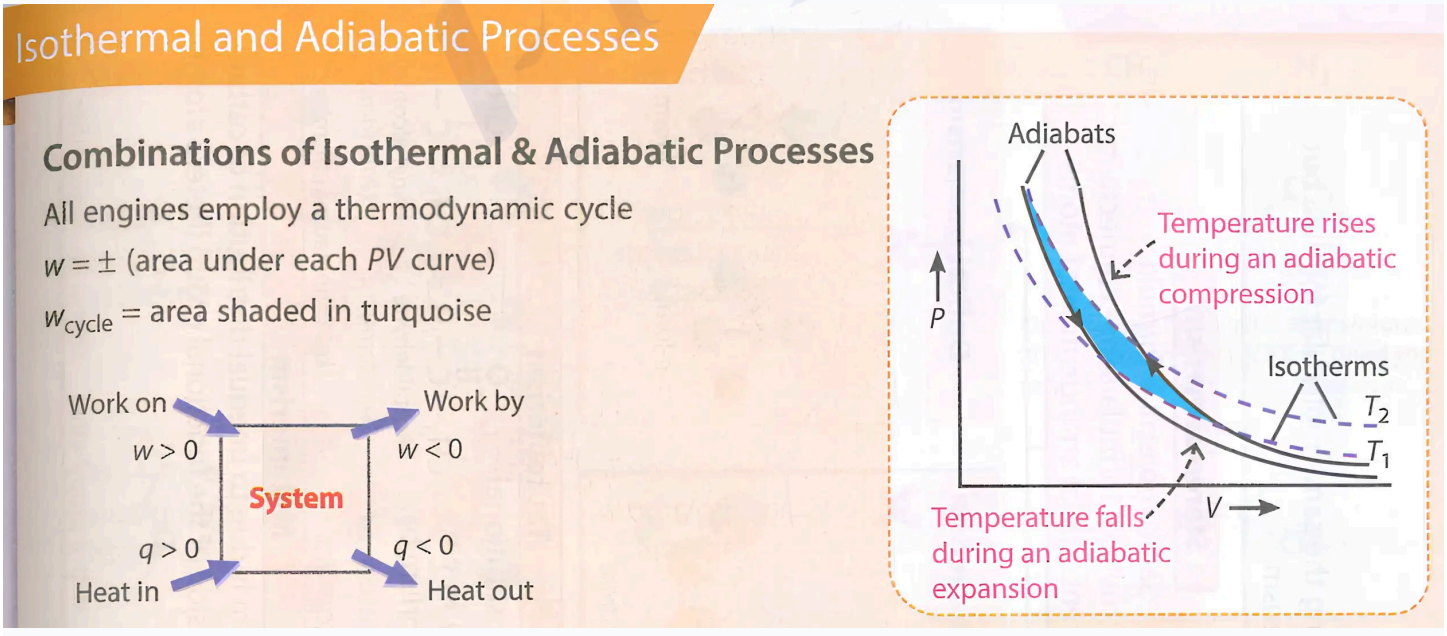
Memory Aids

- Open = Both matter & energy exchange
- Closed = Only energy exchange
- Isolated = No exchange

The State of the System

State and Processes

- Thermodynamic state: measurable properties needed for calculations.
- State variables/functions: pressure (P), temperature (T), volume (V), amount (n); depend only on state, not path.
- Thermodynamic processes:
  - Isothermal: constant temperature (ΔT=0)
  - Adiabatic: no heat exchange (q=0)
  - Isobaric: constant pressure (ΔP=0)
  - Isochoric: constant volume (ΔV=0)
  - Processes can be cyclic, reversible, or irreversible.



Solved Examples

- **Ex 1:** Identify process with no heat exchange.
  Answer: Adiabatic process.

Practice Set

- Q1: Which process occurs at constant temperature?
  A. Isobaric
  B. Isothermal
  C. Adiabatic
  D. Isochoric
- Q2: In which process is volume constant?
  A. Isobaric
  B. Isochoric
  C. Isothermal
  D. Adiabatic
- Q3: A cyclic process:
  A. Cannot be reversed
  B. Returns system to original state
  C. Has $\Delta U \neq 0$
  D. Is irreversible

Answer Key

- Q1: B - Isothermal means constant temperature.
- Q2: B - Isochoric means constant volume.
- Q3: B - Cyclic process returns system to original state.

Memory Aids

- Isothermal = T constant
- Adiabatic = No heat exchange
- Isobaric = P constant
- Isochoric = V constant

Internal Energy

Concept and First Law

- Internal energy (U): total energy inside system (chemical, electrical, mechanical, etc.), a state function.
- U changes when:
  - Heat enters or leaves system
  - Work is done on or by system
  - Matter enters or leaves system
- First law: Energy of isolated system is constant (conservation of energy).
- Equation: $\Delta U = q + w$ or $q = \Delta U - w$
- q and w are path functions; ΔU is state function.
- Sign convention:
  - Heat absorbed by system = +ve
  - Heat evolved by system = −ve
  - Work done on system = +ve
  - Work done by system = −ve

- For adiabatic change: $\Delta U = U_2 - U_1 = w_{ad}$ (q=0)
- At constant volume (no work): $\Delta U = q_v$

Solved Examples

- **Ex 1:** Calculate ΔU if 100 J heat is absorbed and 50 J work is done by system.
  Solution: $\Delta U = q + w = +100 + (-50) = +50$ J

Practice Set

- Q1: Internal energy is:
  A. Path function
  B. State function
  C. Neither
  D. Both
- Q2: If system absorbs 200 J heat and does 80 J work, ΔU is:
  A. 280 J
  B. 120 J
  C. −120 J
  D. −280 J
- Q3: For adiabatic process:
  A. q ≠ 0
  B. w = 0
  C. ΔU = w
  D. ΔU = q

Answer Key

- Q1: B - Internal energy is a state function.
- Q2: B - $\Delta U = q + w = 200 - 80 = 120$ J.
- Q3: C - For adiabatic, ΔU = work done (q=0).

Memory Aids

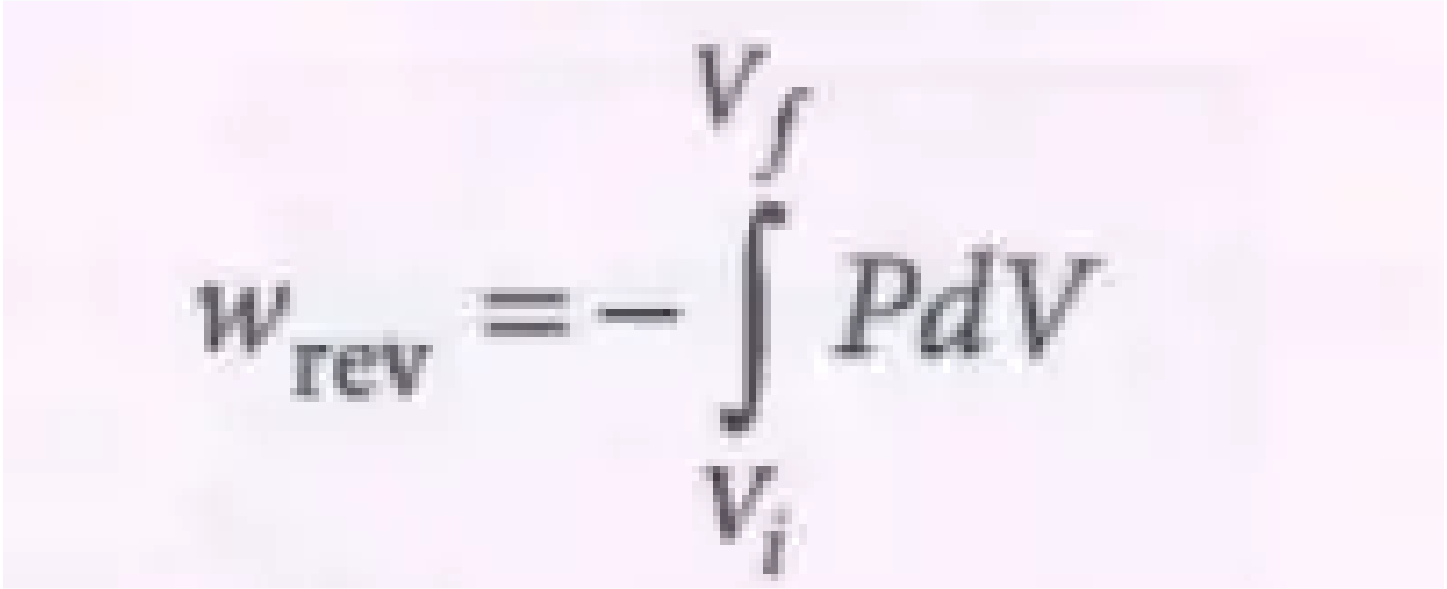
- $\Delta U = q + w$
- Heat in (+), work done by system (−)
- Adiabatic: no heat exchange

Work

Isothermal reversible expansion of an ideal gas

- Gas expands in a cylinder with a frictionless piston; volume increases, pressure decreases.
- External pressure initially equals internal pressure; piston moves when external pressure decreases by dP.
- Work done in infinitesimal expansion: $dw = -P_{ex}dV = -PdV$.
- Total work done from volume V_i to V_f: $w_{rev} = -\int_{V_i}^{V_f} PdV$.

- Using ideal gas law $P = \frac{nRT}{V}$, work done: $w_{rev} = -nRT \ln \frac{V_f}{V_i} = nRT \ln \frac{P_i}{P_f}$.
- Work done during isothermal reversible compression: $w_{rev} = 2.303nRT \log \frac{P_i}{P_f}$.



$$w_{\text{rev}} = -RT \cdot \left(\int_{V_i}^{V_f} \frac{dV}{V} = -RT \ln \frac{V_f}{V_i} \right.$$

$$\text{For } n \text{ moles, } w_{\text{rev}} = --nRT \ln \frac{V_f}{V_i}$$

$$\therefore \; P_i V_i = P_f V_f$$

$$\therefore \; w_{\text{rev}} = -nRT \ln \frac{P_i}{P_f}$$

$$\therefore \; w_{\text{rev}} = -nRT \ln \frac{P_i}{P_f} = -2.303nRT \log \frac{P_i}{P_f}$$

Free expansion

- Expansion of gas into vacuum ($P_{ex} = 0$) is free expansion.
- No work done during free expansion, reversible or irreversible.
- At constant volume ($\Delta V = 0$), internal energy change $\Delta U = q_v$.

Isothermal and free expansion of an ideal gas

- Isothermal expansion into vacuum: work $w = 0$ since $P_{ex} = 0$.
- Isothermal irreversible change: $q = -P_{ex}(V_f - V_i)$.
- Isothermal reversible change: $q = w = 2.303nRT \log \frac{V_f}{V_i}$.

Adiabatic change

- Heat exchange $q = 0$.
- Change in internal energy equals work done: $\Delta U = w_{ad}$.

Solved Examples

- **Ex 1:** Calculate work done in isothermal expansion of 1 mole ideal gas from 10 L to 20 L at 300 K.
  - Use $w = -nRT \ln \frac{V_f}{V_i}$.
  - $w = -1 \times 8.314 \times 300 \times \ln(2) = -1729\ \text{J}$.
- **Ex 2:** Work done in isothermal compression from 20 atm to 10 atm at 300 K for 2 moles.
  - $w = 2.303nRT \log \frac{P_i}{P_f} = 2.303 \times 2 \times 8.314 \times 300 \times \log(2) \approx 7910\ \text{J}$.

Practice Set

- Q1: Work done in free expansion of an ideal gas is:
  - A) Positive
  - B) Negative
  - C) Zero
  - D) Depends on volume
- Q2: For isothermal reversible expansion, work done is:
  - A) $-nRT \ln \frac{V_f}{V_i}$
  - B) $nRT \ln \frac{V_f}{V_i}$
  - C) Zero
  - D) $nRT \ln \frac{P_f}{P_i}$
- Q3: Work done in isothermal compression of 1 mole gas from 5 atm to 10 atm at 300 K is:
  - A) Negative
  - B) Zero
  - C) Positive
  - D) Cannot be determined

Answer Key

- Q1: C) Zero - No work is done in free expansion as external pressure is zero.
- Q2: A) $-nRT \ln \frac{V_f}{V_i}$ - Work done is negative for expansion.
- Q3: C) Positive - Work done on the gas during compression is positive.

Memory Aids

- Work done in isothermal expansion: Negative sign, gas does work on surroundings.
- Free expansion: No work done, $w = 0$.
- Compression work is positive, expansion work is negative.

Enthalpy

Basic concepts

- Total heat content at constant pressure.
- Extensive and state function.
- Mathematically: $H = U + PV$.
- Change in enthalpy: $\Delta H = \Delta U + P\Delta V$.
- Exothermic: $\Delta H < 0$, Endothermic: $\Delta H > 0$.

Relation with first law

- First law: $q = \Delta U - w$, with $w = -P\Delta V$.
- At constant volume: $w = 0$, so $q_v = \Delta U$.
- At constant pressure: $q_p = \Delta U + P\Delta V = \Delta H$.
- $\Delta H = \Delta U + \Delta n_g RT$, where Δn_g is change in moles of gas.
- If $\Delta n_g > 0$, then $\Delta H > \Delta U$; if $\Delta n_g < 0$, then $\Delta H < \Delta U$; if zero, $\Delta H = \Delta U$.
- For solids/liquids or closed vessel reactions, $\Delta H = \Delta U$.

Solved Examples

- **Ex 1:** Calculate ΔH for reaction with $\Delta U = -100\ \text{kJ}$, $\Delta n_g = 2$, at 298 K.
  - $\Delta H = \Delta U + \Delta n_g RT = -100 + 2 \times 8.314 \times 298/1000 = -95.05\ \text{kJ}$.
- **Ex 2:** For reaction with no gas change, $\Delta H = \Delta U = -50\ \text{kJ}$.

Practice Set

- Q1: Enthalpy is a:
  - A) Intensive property
  - B) Extensive property
  - C) Neither
  - D) Depends on temperature
- Q2: At constant volume, heat exchanged equals:
  - A) ΔH
  - B) ΔU
  - C) Zero
  - D) Work done
- Q3: For reaction with $\Delta n_g = 1$, ΔH is:
  - A) Less than ΔU
  - B) Equal to ΔU
  - C) Greater than ΔU

- D) Cannot say

Answer Key

- Q1: B) Extensive property - Enthalpy depends on system size.
- Q2: B) ΔU - At constant volume, heat equals internal energy change.
- Q3: C) Greater than ΔU - Positive gas moles increase enthalpy.

Memory Aids

- $\Delta H = \Delta U + P\Delta V$.
- Exothermic: Negative ΔH, Endothermic: Positive ΔH.
- At constant pressure, heat equals enthalpy change.

Extensive and Intensive Properties

Definitions and examples

- **Extensive:** Depends on amount of matter (mass, volume, internal energy, enthalpy, heat capacity).
- **Intensive:** Independent of amount (temperature, density, pressure, molar volume, molar heat capacity).

Solved Examples

- **Ex 1:** Identify property: mass (Extensive), temperature (Intensive).
- **Ex 2:** Heat capacity of 2 moles vs 1 mole: doubles (Extensive).

Practice Set

- Q1: Which is an intensive property?
  - A) Volume
  - B) Mass
  - C) Temperature
  - D) Enthalpy
- Q2: Heat capacity depends on:
  - A) Amount of substance
  - B) Temperature only
  - C) Pressure only
  - D) None
- Q3: Molar heat capacity is:
  - A) Extensive
  - B) Intensive
  - C) Neither
  - D) Depends on mass

Answer Key

- Q1: C) Temperature - Intensive property does not depend on amount.
- Q2: A) Amount of substance - Heat capacity increases with quantity.
- Q3: B) Intensive - Molar heat capacity is per mole, independent of amount.

Memory Aids

- Extensive: Depends on size (mass, volume).
- Intensive: Independent of size (temperature, pressure).

Heat Capacity

Types and formulas

- **Heat capacity (C):** Heat required to raise temperature by 1°C, $C = \frac{Q}{\Delta T}$.
- **Molar heat capacity (C_m):** Heat to raise 1 mole by 1°C, $C_m = \frac{C}{n}$.
- **Specific heat capacity (c):** Heat to raise 1 g by 1°C, $q = mc\Delta T$.
- Water's specific heat: 1 cal g^{-1} K^{-1} = 4.18 J g^{-1} K^{-1}.
- Relation for ideal gas: $C_p = \frac{\Delta H}{\Delta T}, C_v = \frac{\Delta U}{\Delta T}, C_p - C_v = R$.
- Ratio $\gamma = \frac{C_p}{C_v}$: monatomic = 1.66, diatomic = 1.40, triatomic = 1.30.

Solved Examples

- **Ex 1:** Calculate C_m if $C = 100\ \text{J/K}$ and $n = 2$ moles.
  - $C_m = \frac{100}{2} = 50\ \text{J/mol K}$.
- **Ex 2:** Find heat required to raise 10 g water by 5°C.
  - $q = mc\Delta T = 10 \times 4.18 \times 5 = 209\ \text{J}$.

Practice Set

- Q1: Heat capacity is heat required to raise temperature by:
  - A) 1 K
  - B) 1°C
  - C) 10°C
  - D) 100°C
- Q2: Specific heat capacity is heat required to raise temperature of:
  - A) 1 mole
  - B) 1 gram
  - C) 1 kg
  - D) 1 liter
- Q3: For ideal gas, $C_p - C_v =$?
  - A) 0
  - B) R
  - C) C_p
  - D) C_v

Answer Key

- Q1: B) 1°C - Definition of heat capacity.
- Q2: B) 1 gram - Specific heat is per gram basis.
- Q3: B) R - Difference of heat capacities equals gas constant.

Memory Aids

- Heat capacity: Heat per 1°C rise.
- Specific heat: Heat per gram per °C.
- $C_p - C_v = R$ for ideal gases.

Efficiency of Heat Engine

Basic formula and terms

- Efficiency $\eta = \frac{w}{q_2} = \frac{T_2 - T_1}{T_2} = \frac{q_2 - q_1}{q_2}$.
- q_2: Heat absorbed at high temperature T_2.
- q_1: Heat rejected at low temperature T_1.
- w: Net work done by the engine.
- Unit conversions: 1 J = 0.2390 cal, 1 cal = 4.184 J, 1 litre-atm = 101.3 J.

Solved Examples

- **Ex 1:** Calculate efficiency if $T_2 = 500K$, $T_1 = 300K$.
  - $\eta = \frac{500 - 300}{500} = 0.4 = 40\%$.

Practice Set

- Q1: Efficiency of heat engine depends on:
  - A) T_1 only
  - B) T_2 only
  - C) Both T_1 and T_2
  - D) Neither
- Q2: If $q_1 = q_2$, efficiency is:
  - A) 1
  - B) 0
  - C) $\frac{T_1}{T_2}$
  - D) $\frac{T_2}{T_1}$
- Q3: Net work done w is:
  - A) $q_1 + q_2$
  - B) $q_2 - q_1$
  - C) $q_1 - q_2$
  - D) Zero

Answer Key

- Q1: C) Both T_1 and T_2 - Efficiency depends on temperature difference.
- Q2: B) 0 - No heat difference means no work done.
- Q3: B) $q_2 - q_1$ - Work done equals heat absorbed minus heat rejected.

Memory Aids

- Efficiency $\eta = \frac{T_2 - T_1}{T_2}$.
- Heat engine converts heat difference into work.

Joule–Thomson Effect

Definition and coefficient

- Temperature change during adiabatic expansion from high to low pressure through a small opening.
- Joule-Thomson coefficient: $\mu_{JT} = \left(\frac{\partial T}{\partial P}\right)_H$.
- Zero for ideal gases; no heat absorbed or evolved during expansion in vacuum.

Adiabatic reversible expansion

- Relations: $PV^\gamma = \text{constant}$, $TV^{\gamma-1} = \text{constant}$, $T^r P^{1-\gamma} = \text{constant}$.
- Work done: $w = \frac{nR(T_1 - T_2)}{\gamma - 1}$.

Solved Examples

- **Ex 1:** Calculate work done in adiabatic expansion of 1 mole gas from 400 K to 300 K, $\gamma = 1.4$.
  - $w = \frac{1 \times 8.314 \times (400 - 300)}{1.4 - 1} = 2078.5\ \text{J}$.

Practice Set

- Q1: Joule-Thomson effect is zero for:
  - A) Real gases
  - B) Ideal gases
  - C) Liquids
  - D) Solids
- Q2: Work done in adiabatic expansion is proportional to:
  - A) Temperature difference
  - B) Pressure difference
  - C) Volume difference
  - D) Number of moles only
- Q3: For adiabatic reversible expansion, $PV^\gamma =$?
  - A) Variable
  - B) Constant
  - C) Zero

- D) Depends on temperature

Answer Key

- Q1: B) Ideal gases - No temperature change during expansion.
- Q2: A) Temperature difference - Work depends on $T_1 - T_2$.
- Q3: B) Constant - Adiabatic condition relation.

Memory Aids

- Joule-Thomson effect zero for ideal gases.
- Adiabatic expansion: $PV^\gamma = \text{constant}$.
- Work $w = \frac{nR(T_1 - T_2)}{\gamma - 1}$.

Measurement of ΔU and ΔH: Calorimetry

Measurement of ΔU

- Performed using a bomb calorimeter.
- Formula: ΔU = C × Δt × (M / m)
- Where,
  - m = mass of substance taken
  - M = molar mass of substance
  - C = heat capacity of the calorimeter
  - Δt = rise in temperature

Measurement of ΔH

- Reaction carried out in an insulated vessel (calorimeter).
- Temperature change (rise or fall) is recorded.
- Temperature change indicates whether reaction is exothermic or endothermic.

Solved Examples

- **Ex 1:** Calculate ΔU if C = 500 J/K, Δt = 2 K, M = 100 g/mol, m = 50 g.
  - ΔU = 500 × 2 × (100 / 50) = 2000 J
- **Ex 2:** In a calorimeter, temperature rises by 3 K for a reaction. If C = 400 J/K, M = 80 g/mol, m = 40 g, find ΔU.
  - ΔU = 400 × 3 × (80 / 40) = 2400 J

Practice Set

- Q1: Which instrument is used to measure ΔU?
  - A. Bomb calorimeter
  - B. Thermometer
  - C. Barometer

    - D. Voltmeter
- Q2: In the formula ΔU = C × Δt × (M / m), what does Δt represent?
    - A. Mass of substance
    - B. Molar mass
    - C. Heat capacity
    - D. Temperature rise
- Q3: A reaction in a calorimeter causes temperature to fall. What type of reaction is it?
    - A. Exothermic
    - B. Endothermic
    - C. Neutral
    - D. None of these

Answer Key

- Q1: A - Bomb calorimeter is specifically designed to measure internal energy change (ΔU).
- Q2: D - Δt represents the rise in temperature during the reaction.
- Q3: B - Temperature fall indicates an endothermic reaction absorbing heat.

Memory Aids

- Bomb calorimeter → Measures ΔU (internal energy change).
- Insulated vessel calorimeter → Measures ΔH (enthalpy change).
- Temperature rise → Exothermic reaction.
- Temperature fall → Endothermic reaction.

Enthalpy Change Reaction Enthalpy

Definition and Formula

- **Reaction enthalpy ($\Delta_r H$)**: Enthalpy change during a chemical reaction.
- Calculated as: $\Delta_r H = \sum a_i H_{products} - \sum b_i H_{reactants}$
- Here, a_i and b_i are stoichiometric coefficients of products and reactants.
- Example: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
- Reaction enthalpy: $\Delta_r H = [H_m(CO_2) + 2H_m(H_2O)] - [H_m(CH_4) + 2H_m(O_2)]$
- H_m is molar enthalpy.

Standard Enthalpy of Reaction

- Enthalpy change when all substances are in their *standard states* (pure form at 1 bar pressure).
- Example: Liquid ethanol at 298 K and 1 bar; solid iron at 500 K and 1 bar.

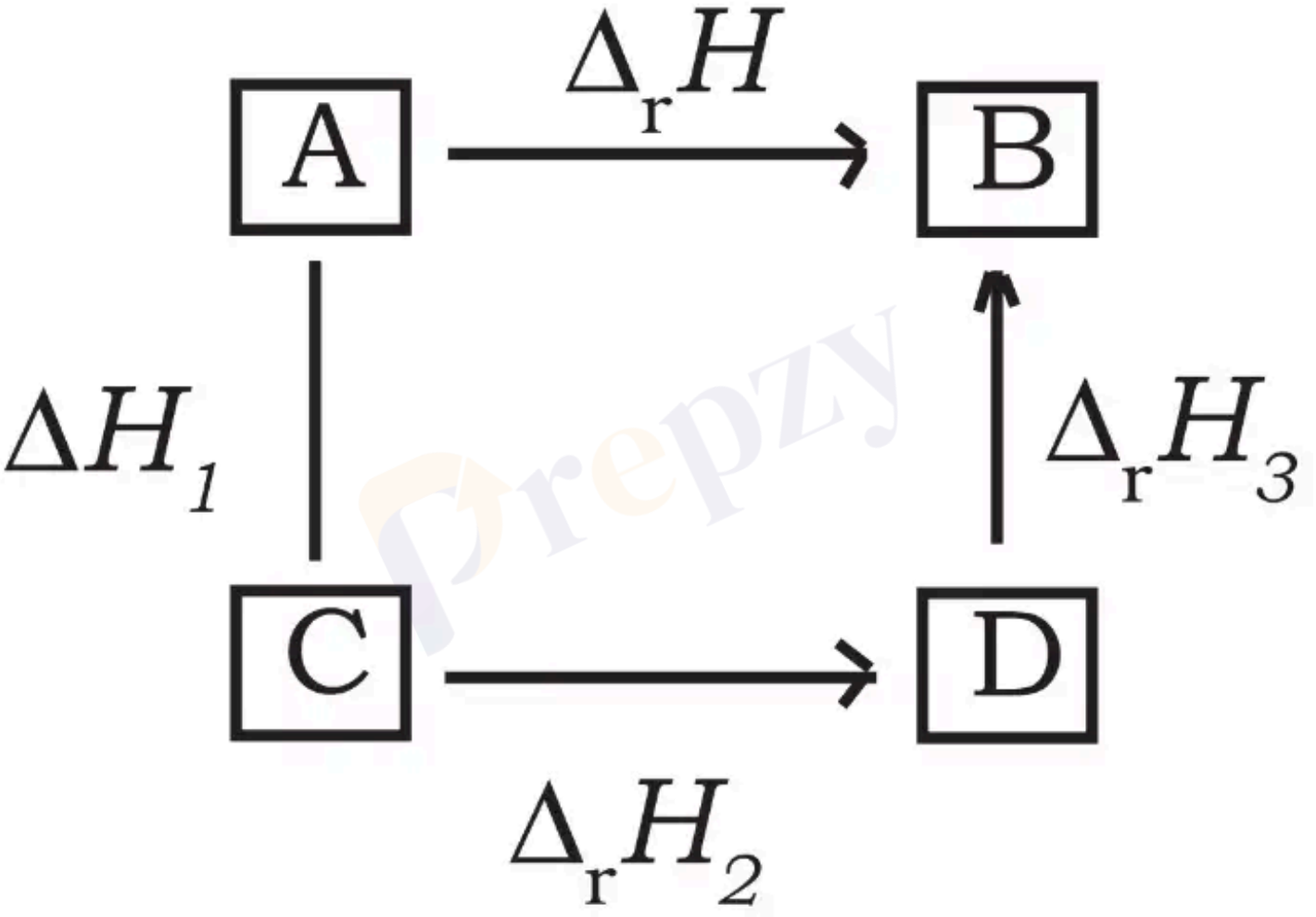
Enthalpy Changes During Phase Transformations

- **Enthalpy of fusion:** Energy to convert 1 mole solid to liquid at melting point.
    - Example: $H_2O_{(s)} \rightarrow H_2O_{(l)}$ at 273 K, $\Delta_{fus}H^\circ = +6.0$ kJ/mol
- **Enthalpy of vaporisation:** Energy to convert 1 mole liquid to vapor at boiling point.

- Example: $H_2O_{(l)} \rightarrow H_2O_{(g)}$ at 373 K, $\Delta_{vap}H^\circ = +40.79$ kJ/mol
- **Enthalpy of sublimation:** Energy to convert 1 mole solid directly to vapor.
  - Example: Iodine, $\Delta_{sub}H^\circ = +62.4$ kJ/mol
  - Relation: $\Delta_{sub}H^\circ = \Delta_{fus}H^\circ + \Delta_{vap}H^\circ$
- **Enthalpy of formation:** Energy change when 1 mole compound forms from elements in stable states.
  - Example: $C_{graphite,s} + 2H_2(g) \rightarrow CH_4(g)$, $\Delta_f H^\circ = -74.8$ kJ/mol

Hess's Law of Constant Heat Summation

- If a reaction occurs in steps, total enthalpy change is sum of enthalpy changes of steps at same temperature.
- Formula: $\Delta_r H = \Delta_r H_1 + \Delta_r H_2 + \Delta_r H_3$



Solved Examples

- **Ex 1:** Calculate $\Delta_r H$ for $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ given molar enthalpies: $H_m(CH_4) = -75$, $H_m(CO_2) = -393$, $H_m(H_2O) = -286$, $H_m(O_2) = 0$ kJ/mol.
  - $\Delta_r H = [-393 + 2(-286)] - [-75 + 2(0)] = (-393 - 572) - (-75) = -965 + 75 = -890$ kJ/mol
- **Ex 2:** Use Hess's law to find $\Delta_r H$ if $\Delta_r H_1 = -100$, $\Delta_r H_2 = +50$, $\Delta_r H_3 = -30$ kJ/mol.
  - $\Delta_r H = -100 + 50 - 30 = -80$ kJ/mol

Practice Set

- Q1: What is the reaction enthalpy for $N_2 + 3H_2 \rightarrow 2NH_3$ if $\sum H_{products} = -92$ kJ and $\sum H_{reactants} = 0$ kJ?
  - A. +92 kJ
  - B. -92 kJ
  - C. 0 kJ
  - D. -46 kJ
- Q2: Enthalpy of vaporisation of water is:
  - A. +6.0 kJ/mol
  - B. +40.79 kJ/mol
  - C. +62.4 kJ/mol
  - D. -74.8 kJ/mol
- Q3: Hess's law states that total enthalpy change is:
  - A. Product of enthalpy changes of steps
  - B. Sum of enthalpy changes of steps
  - C. Difference of enthalpy changes of steps
  - D. None of the above

Answer Key

- Q1: B. -92 kJ; Reaction enthalpy = products - reactants = -92 - 0 = -92 kJ (exothermic)
- Q2: B. +40.79 kJ/mol; Energy to convert liquid water to vapor at boiling point.
- Q3: B. Sum of enthalpy changes of steps; Hess's law defines total enthalpy as sum of steps.

Memory Aids

- Enthalpy of sublimation = fusion + vaporisation
- Standard state = pure substance at 1 bar
- Hess's law: Add enthalpy changes of intermediate steps

Enthalpy of Combustion

Concept

- Change in enthalpy when 1 mole of substance burns completely in oxygen.
- Example: $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$; $\Delta_c H° = -393.5$ kJ mol^{-1}
- $\Delta_c H°$ is always negative (exothermic reaction).

Solved Examples

- **Ex 1:** Calculate heat released when 2 moles of carbon burn.
  - $\Delta_c H°$ for 1 mole C = −393.5 kJ
  - For 2 moles: Heat = 2 × (−393.5) = −787 kJ

Practice Set

- Q1: What is the sign of enthalpy of combustion?
  - A. Positive
  - B. Negative
  - C. Zero
  - D. Depends on substance
- Q2: Which of the following represents enthalpy of combustion?
  - A. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
  - B. $C + O_2 \rightarrow CO_2$
  - C. $H_2 + 1/2\ O_2 \rightarrow H_2O$
  - D. $N_2 + O_2 \rightarrow 2NO$
- Q3: Heat released on burning 3 moles of carbon is?
  - A. −1180.5 kJ
  - B. −393.5 kJ
  - C. +1180.5 kJ
  - D. +393.5 kJ

Answer Key

- Q1: B. Negative – Combustion releases heat, so enthalpy change is negative.
- Q2: B. $C + O_2 \rightarrow CO_2$ – Represents combustion of 1 mole carbon.
- Q3: A. −1180.5 kJ – 3 × (−393.5) = −1180.5 kJ heat released.

Memory Aids

- Combustion enthalpy always negative (exothermic).
- Complete burning means full oxidation to CO_2 and H_2O.

Enthalpy of Atomization

Concept

- Enthalpy change when 1 mole of bonds breaks to form gaseous atoms.
- Example: $CH_{4(g)} \rightarrow C_{(g)} + 4H_{(g)}$; $\Delta_a H = 1665\ kJ\ mol^{-1}$
- Bond dissociation enthalpy is enthalpy to break 1 mole of covalent bonds in gas phase.

Solved Examples

- **Ex 1:** Calculate bond dissociation enthalpy if breaking 1 mole C–H bonds requires 1665 kJ.
  - Given: $CH_4 \rightarrow C + 4H$, $\Delta_a H = 1665$ kJ
  - Bond dissociation enthalpy per C–H bond = 1665/4 = 416.25 kJ/mol

Practice Set

- Q1: Enthalpy of atomization involves breaking bonds to form?
  - A. Molecules
  - B. Ions
  - C. Gaseous atoms
  - D. Liquids
- Q2: If $\Delta_a H$ for CH_4 is 1665 kJ, bond dissociation enthalpy for one C–H bond is?
  - A. 416.25 kJ/mol
  - B. 1665 kJ/mol
  - C. 332.5 kJ/mol
  - D. 665 kJ/mol
- Q3: Bond dissociation enthalpy is enthalpy change when?
  - A. One mole of ionic bonds break
  - B. One mole of covalent bonds break in gas phase
  - C. One mole of atoms form molecules
  - D. One mole of molecules ionize

Answer Key

- Q1: C. Gaseous atoms – Atomization forms gaseous atoms by bond breaking.
- Q2: A. 416.25 kJ/mol – Total enthalpy divided by number of bonds.
- Q3: B. One mole of covalent bonds break in gas phase – Definition of bond dissociation enthalpy.

Memory Aids

- Atomization = bond breaking to atoms in gas phase.
- Bond dissociation enthalpy = energy to break covalent bonds.

Bond Enthalpy

Concept

- Standard enthalpy of reaction related to bond enthalpies of reactants and products.
- Formula: $\Delta_r H^\circ = \Sigma \text{ bond enthalpies}_{\text{reactants}} - \Sigma \text{ bond enthalpies}_{\text{products}}$
- Used for gas phase reactions involving covalent bonds.

Solved Examples

- **Ex 1:** Calculate $\Delta_r H^\circ$ for $H_2 + Cl_2 \rightarrow 2HCl$
  - Bond enthalpies: H–H = 436, Cl–Cl = 243, H–Cl = 431 kJ/mol
  - $\Delta_r H^\circ = (436 + 243) - 2 \times 431 = 679 - 862 = -183$ kJ/mol

Practice Set

- Q1: Bond enthalpy is energy required to break bonds in?
  - A. Liquids
  - B. Solids
  - C. Gaseous molecules
  - D. Ionic compounds
- Q2: $\Delta_r H^\circ$ = Σ bond enthalpies reactants – products means reaction is?
  - A. Exothermic if negative
  - B. Endothermic if negative
  - C. Always positive
  - D. Zero
- Q3: Calculate $\Delta_r H^\circ$ for $CH_4 + 2Cl_2 \rightarrow CH_2Cl_2 + 2HCl$ given bond enthalpies (C–H=412, Cl–Cl=243, C–Cl=338, H–Cl=431 kJ/mol).
  - A. –123 kJ/mol
  - B. +123 kJ/mol
  - C. –200 kJ/mol
  - D. +200 kJ/mol

Answer Key

- Q1: C. Gaseous molecules – Bond enthalpy applies to gaseous covalent bonds.
- Q2: A. Exothermic if negative – Negative ΔH means heat release.
- Q3: A. –123 kJ/mol – Calculate reactants and products bond enthalpy difference.

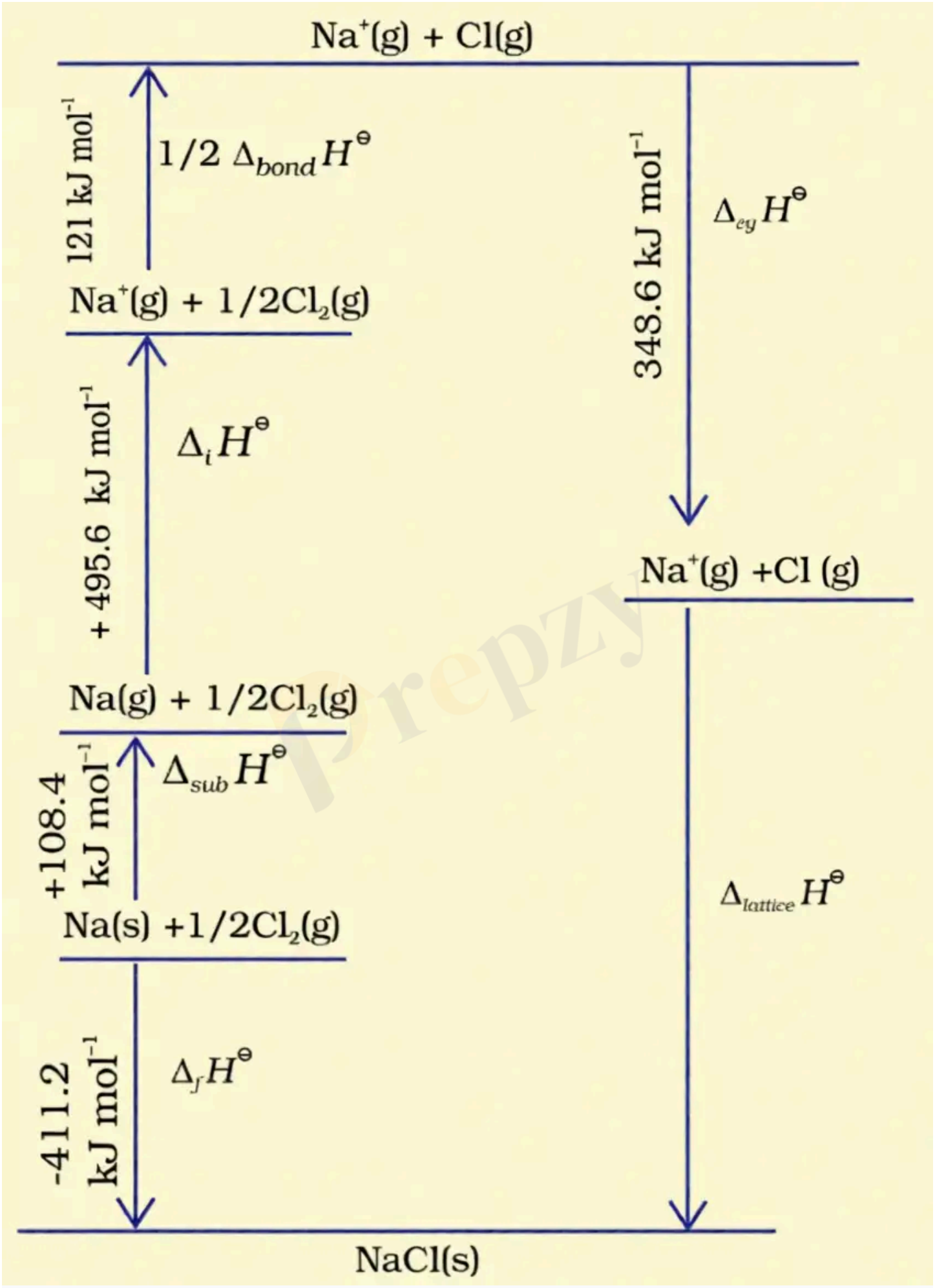
Memory Aids

- $\Delta_r H^\circ$ = Σ bonds broken – Σ bonds formed.
- Negative ΔH means exothermic reaction.

Lattice Enthalpy

Concept

- Enthalpy change when 1 mole ionic compound dissociates into gaseous ions.
- Calculated using Born–Haber cycle.
- Example: NaCl formation enthalpy equation:
  - $\Delta_f H^\circ(NaCl) = \Delta_{sub} H^\circ(Na) + 1/2\ \Delta_{diss} H^\circ(Cl_2) + \Delta_{ion} H^\circ(Na) + \Delta_{eg} H^\circ(Cl) + \Delta_{lattice} H^\circ(NaCl)$


Na⁺(g) + Cl(g)
121 kJ mol⁻¹
1/2 Δbond H⊖
348.6 kJ mol⁻¹
Δeg H⊖
Na⁺(g) + 1/2Cl2(g)
+ 495.6 kJ mol⁻¹
Δi H⊖
Na⁺(g) +Cl (g)
Na(g) + 1/2Cl2(g)
+108.4 kJ mol⁻¹
Δsub H⊖
Δlattice H⊖
Na(s) +1/2Cl2(g)
-411.2 kJ mol⁻¹
Δf H⊖
NaCl(s)

Solved Examples

- **Ex 1:** Use Born-Haber cycle to find lattice enthalpy if other values are known.

Practice Set

- Q1: Lattice enthalpy is enthalpy change when?
  - A. Ionic compound forms solid
  - B. Ionic compound dissociates into gaseous ions
  - C. Covalent bonds break
  - D. Molecules ionize
- Q2: Born-Haber cycle helps calculate?
  - A. Bond enthalpy
  - B. Lattice enthalpy
  - C. Enthalpy of combustion
  - D. Enthalpy of atomization
- Q3: In NaCl Born-Haber cycle, which step involves ionization energy?
  - A. $Na(s) \rightarrow Na(g)$
  - B. $Cl_2(g) \rightarrow 2Cl(g)$
  - C. $Na(g) \rightarrow Na^+(g) + e^-$
  - D. $Cl(g) + e^- \rightarrow Cl^-(g)$

Answer Key

- Q1: B. Ionic compound dissociates into gaseous ions – Definition of lattice enthalpy.
- Q2: B. Lattice enthalpy – Born-Haber cycle calculates lattice enthalpy.
- Q3: C. $Na(g) \rightarrow Na^+(g) + e^-$ – Ionization energy step.

Memory Aids

- Born-Haber cycle = sum of steps to form ionic solid.
- Lattice enthalpy is always exothermic (negative) when ions form solid.

Enthalpy of Solution

Concept

- Enthalpy change when 1 mole of substance dissolves in specified solvent amount.
- Formula: $\Delta_{sol}H° = \Delta_{lattice}H° + \Delta_{hyd}H°$
- Most ionic compounds have positive $\Delta_{sol}H°$ (endothermic dissociation).
- Solubility increases with temperature if $\Delta_{sol}H°$ is positive.
- High lattice enthalpy may prevent dissociation.

Solved Examples

- **Ex 1:** Calculate $\Delta_{sol}H°$ if lattice enthalpy = 700 kJ/mol and hydration enthalpy = –650 kJ/mol.
  - $\Delta_{sol}H°$ = 700 + (–650) = +50 kJ/mol (endothermic)

Practice Set

- Q1: Enthalpy of solution is heat change when?
  - A. Substance melts
  - B. Substance dissolves in solvent
  - C. Substance reacts with acid
  - D. Substance ionizes
- Q2: If $\Delta_{sol}H°$ is positive, solubility of salt?
  - A. Decreases with temperature
  - B. Increases with temperature
  - C. Remains constant
  - D. Zero
- Q3: What happens if lattice enthalpy is very high?
  - A. Salt dissolves easily
  - B. Salt does not dissociate
  - C. Salt reacts with water
  - D. Salt melts

Answer Key

- Q1: B. Substance dissolves in solvent – Definition of enthalpy of solution.
- Q2: B. Increases with temperature – Endothermic dissolution favors solubility with heat.
- Q3: B. Salt does not dissociate – High lattice enthalpy prevents dissociation.

Memory Aids

- $\Delta_{sol}H°$ = lattice enthalpy + hydration enthalpy.
- Positive $\Delta_{sol}H°$ means endothermic dissolution.

Enthalpy of Dilution

Concept

- Heat change when additional solvent is added to a solution at constant temperature and pressure.
- Heat withdrawn from surroundings on dilution.
- Depends on original concentration of solution.

Solved Examples

- **Ex 1:** Adding water to concentrated acid causes heat change due to dilution.

Practice Set

- Q1: Enthalpy of dilution is heat change when?
  - A. Solute dissolves
  - B. Solvent added to solution
  - C. Solute reacts
  - D. Solution evaporates
- Q2: Enthalpy of dilution depends on?
  - A. Temperature only
  - B. Original concentration
  - C. Volume only
  - D. Pressure only
- Q3: Heat withdrawn on dilution means process is?
  - A. Exothermic
  - B. Endothermic
  - C. Neutral
  - D. Variable

Answer Key

- Q1: B. Solvent added to solution – Definition of enthalpy of dilution.
- Q2: B. Original concentration – Dilution enthalpy depends on concentration.
- Q3: A. Exothermic – Heat withdrawn means heat released to surroundings.

Memory Aids

- Dilution enthalpy depends on concentration.
- Heat withdrawn means exothermic process.

Enthalpy of Solution at Infinite Dilution

Concept

- Enthalpy change when substance dissolves in infinite solvent amount.
- Ion interactions negligible at infinite dilution.

Solved Examples

- **Ex 1:** Enthalpy change measured when solute dissolves in large solvent volume.

Practice Set

- Q1: Infinite dilution means solvent amount is?
  - A. Very small
  - B. Infinite

    - C. Equal to solute
    - D. None
- Q2: At infinite dilution, ion interactions are?
    - A. Strong
    - B. Negligible
    - C. Moderate
    - D. Variable
- Q3: Enthalpy of solution at infinite dilution is?
    - A. Dependent on concentration
    - B. Independent of concentration
    - C. Zero
    - D. Negative always

Answer Key

- Q1: B. Infinite – Definition of infinite dilution.
- Q2: B. Negligible – Ion interactions minimal at infinite dilution.
- Q3: B. Independent of concentration – No ion interaction effect.

Memory Aids

- Infinite dilution means very large solvent volume.
- Ion interactions negligible at infinite dilution.

Enthalpy of Neutralisation

Concept

- Heat change when 1 gram equivalent acid neutralised by 1 gram equivalent base in dilute solution.
- Example: $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$; $\Delta_{neu}H = -57.1$ kJ mol^{-1}
- Usually exothermic (heat released).

Solved Examples

- **Ex 1:** Calculate heat released when 1 mole HCl neutralised by NaOH.
    - Heat = –57.1 kJ

Practice Set

- Q1: Enthalpy of neutralisation is heat change when?
    - A. Acid reacts with base
    - B. Acid evaporates
    - C. Base melts
    - D. Acid ionizes
- Q2: Heat change in neutralisation is usually?
    - A. Positive

    - B. Negative
    - C. Zero
    - D. Variable
- Q3: Reaction: HCl + NaOH → ?
    - A. NaCl + H_2O
    - B. NaOH + HCl
    - C. H_2 + Cl_2
    - D. Na + Cl

Answer Key

- Q1: A. Acid reacts with base – Definition of neutralisation.
- Q2: B. Negative – Heat is released in neutralisation.
- Q3: A. NaCl + H_2O – Products of acid-base neutralisation.

Memory Aids

- Neutralisation always releases heat (exothermic).
- One equivalent acid reacts with one equivalent base.

Enthalpy of Transition

Concept

- Enthalpy change when 1 mole substance changes from one allotrope to another.
- Example: C(graphite) → C(diamond); $\Delta_{trans}H$ = 1.9 kJ mol^{-1}

Solved Examples

- **Ex 1:** Calculate heat absorbed when 1 mole graphite converts to diamond.
    - $\Delta_{trans}H$ = +1.9 kJ/mol (endothermic)

Practice Set

- Q1: Enthalpy of transition involves change between?
    - A. Different elements
    - B. Different allotropes
    - C. Different compounds
    - D. Different ions
- Q2: $\Delta_{trans}H$ for graphite to diamond is?
    - A. –1.9 kJ/mol
    - B. +1.9 kJ/mol
    - C. Zero
    - D. Variable
- Q3: Transition enthalpy is usually?
    - A. Exothermic

- B. Endothermic
- C. Zero
- D. None

Answer Key

- Q1: B. Different allotropes – Transition is between allotropes.
- Q2: B. +1.9 kJ/mol – Heat absorbed to convert graphite to diamond.
- Q3: B. Endothermic – Transition requires energy input.

Memory Aids

- Graphite to diamond is endothermic transition.
- Allotropes differ in structure, not composition.

Enthalpy of Hydration

Concept

- Heat change when 1 mole anhydrous substance combines with water to form hydrated compound.
- Example: $CuSO_{4(s)} + 5H_2O \rightarrow CuSO_4 \cdot 5H_2O_{(s)}$; $\Delta_{hyd}H = -78.2$ kJ mol^{-1}

Solved Examples

- **Ex 1:** Calculate heat released when 1 mole $CuSO_4$ forms pentahydrate.
  - $\Delta_{hyd}H = -78.2$ kJ/mol (exothermic)

Practice Set

- Q1: Enthalpy of hydration is heat change when?
  - A. Substance melts
  - B. Anhydrous substance forms hydrate
  - C. Substance ionizes
  - D. Substance evaporates
- Q2: Heat change in hydration is usually?
  - A. Positive
  - B. Negative
  - C. Zero
  - D. Variable
- Q3: Reaction: $CuSO_4 + 5H_2O \rightarrow$?
  - A. $CuSO_4 \cdot 5H_2O$
  - B. $CuSO_4$
  - C. $Cu + H_2O$
  - D. $CuSO_4 + H_2O$

Answer Key

- Q1: B. Anhydrous substance forms hydrate – Definition of hydration enthalpy.
- Q2: B. Negative – Hydration usually releases heat.
- Q3: A. $CuSO_4 \cdot 5H_2O$ – Hydrated compound formed.

Memory Aids

- Hydration releases heat (exothermic).
- Hydrates contain water molecules in crystal lattice.

Spontaneity

Key Points

- A spontaneous process is irreversible without external influence.
- Entropy (S) measures system randomness or disorder.
- For reversible reactions, $\Delta S = \frac{q_{rev}}{T}$.
- Total entropy change for spontaneous process: $\Delta S_{total} = \Delta S_{system} + \Delta S_{surroundings} > 0$.
- At equilibrium, entropy is maximum and $\Delta S = 0$.
- Entropy change for reversible process: $\Delta S_{sys} = \frac{q_{sys,rev}}{T}$.
- Entropy changes during phase changes:
  - Fusion: $\Delta S = \frac{\Delta H_{fus}}{T_m}$
  - Vaporization: $\Delta S = \frac{\Delta H_{vap}}{T_b}$
  - Sublimation: $\Delta S = \frac{\Delta H_{sub}}{T}$
- T_m: melting point, T_b: boiling point, T: sublimation temperature.
- Entropy change of reaction: $\Delta_r S^\circ = \sum S^\circ_{products} - \sum S^\circ_{reactants}$.

$$\Delta_{fus}S = \frac{\Delta_{fus}H}{T_m}, \quad \Delta_{vap}S = \frac{\Delta_{vap}H}{T_b}, \quad \Delta_{sub}S = \frac{\Delta_{sub}H}{T}$$

Solved Examples

- **Ex 1:** Calculate entropy change when 1 mole of ice melts at 273 K. Given $\Delta H_{fus} = 6.01\ \text{kJ/mol}$.
  - $\Delta S = \frac{\Delta H_{fus}}{T_m} = \frac{6010}{273} = 22.02\ \text{J/K}$

Practice Set

- Q1: What is the entropy change for vaporization at boiling point?
  A) $\frac{\Delta H_{vap}}{T_b}$ B) $\frac{\Delta H_{fus}}{T_m}$ C) $\frac{\Delta H_{sub}}{T}$ D) Zero
- Q2: At equilibrium, entropy change is:
  A) Positive B) Negative C) Zero D) Infinite

- Q3: For a spontaneous process, total entropy change is:
  A) < 0 B) $= 0$ C) > 0 D) Undefined

Answer Key

- Q1: A - Entropy change for vaporization is heat absorbed divided by boiling temperature.
- Q2: C - At equilibrium, entropy change is zero.
- Q3: C - Spontaneous processes have positive total entropy change.

Memory Aids

- Entropy increases with phase change: solid < liquid < gas.
- Spontaneous process: total entropy change > 0.
- At equilibrium, entropy change = 0.

Gibbs Free Energy

Key Points

- Gibbs free energy (G) = energy available for work: $G = H - TS$.
- Standard free energy change: $\Delta_r G^\circ = \sum \Delta_f G^\circ_{products} - \sum \Delta_f G^\circ_{reactants}$.
- Gibbs–Helmholtz equation: $\Delta G = \Delta H - T\Delta S$.
- Relation with emf: $\Delta G^\circ = -nFE^\circ$ where E° is standard emf.
- Spontaneity criteria at constant P and T:
  - $\Delta G < 0$: spontaneous
  - $\Delta G > 0$: non-spontaneous
  - $\Delta G = 0$: equilibrium

$\Delta_r H^\ominus$	$\Delta_r S^\ominus$	$\Delta_r G^\ominus$	Description*
–	+	–	Reaction spontaneous at all temperatures
–	–	– (at low T)	Reaction spontaneous at low temperature
–	–	+ (at high T)	Reaction nonspontaneous at high temperature
+	+	+ (at low T)	Reaction nonspontaneous at low temperature
+	+	– (at high T)	Reaction spontaneous at high temperature
+	–	+ (at all T)	Reaction nonspontaneous at all temperatures

Solved Examples

- **Ex 1:** Determine spontaneity if $\Delta H^\circ = -40\ \text{kJ}$, $\Delta S^\circ = -100\ \text{J/K}$ at 300 K.
  - $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = -40000 - 300 \times (-100) = -40000 + 30000 = -10000\ \text{J}$
  - Since $\Delta G^\circ < 0$, reaction is spontaneous.

Practice Set

- Q1: If $\Delta G < 0$, the process is:
  A) Non-spontaneous B) Spontaneous C) At equilibrium D) Irreversible
- Q2: Gibbs free energy equation is:
  A) $G = H + TS$ B) $G = H - TS$ C) $G = TS - H$ D) $G = H/TS$
- Q3: For $\Delta G^\circ = -nFE^\circ$, what does E° represent?
  A) Standard emf B) Entropy C) Enthalpy D) Temperature

Answer Key

- Q1: B - Negative ΔG means spontaneous process.
- Q2: B - Gibbs free energy is $G = H - TS$.
- Q3: A - E° is standard emf of the cell.

Memory Aids

- Spontaneous if $\Delta G < 0$.
- Use $\Delta G = \Delta H - T\Delta S$ to check spontaneity.
- Negative ΔH and positive ΔS always spontaneous.

Thermodynamics Laws

Key Points

- **Second law:** Total entropy change $\Delta S_{total} > 0$ for spontaneous processes.
- **Third law:** Entropy of pure crystalline substance approaches zero at absolute zero temperature.
- More symmetrical molecules have lower entropy than less symmetrical ones with similar mass and heat capacity.
- Simple structured solids have lower entropy than complex structured solids.
- Entropy decreases when rubber band stretches (more order), increases when egg boils (protein denaturation).
- Entropy of liquids/solids changes little with pressure due to volume stability.
- Entropy change of ideal gas (1 mole):
  - $\Delta S = C_v \ln \frac{T_2}{T_1} + R \ln \frac{V_2}{V_1} = C_p \ln \frac{T_2}{T_1} + R \ln \frac{P_1}{P_2}$
  - Isothermal: $\Delta S_T = R \ln \frac{V_2}{V_1} = R \ln \frac{P_1}{P_2}$
  - Isochoric: $\Delta S_V = C_v \ln \frac{T_2}{T_1} = C_v \ln \frac{P_2}{P_1}$
  - Isobaric: $\Delta S_P = C_p \ln \frac{T_2}{T_1} = C_p \ln \frac{V_2}{V_1}$
- Gibbs Helmholtz equations:
  - $\Delta G = \Delta H + T\left[\frac{d(\Delta G)}{dT}\right]_P$
  - $\frac{\Delta G - \Delta H}{T} = \left[\frac{d(\Delta G)}{dT}\right]_P$
  - $-\Delta S = \left[\frac{d(\Delta G)}{dT}\right]_P$

Solved Examples

- **Ex 1:** Calculate entropy change for 1 mole ideal gas expanding isothermally from volume V_1 to V_2.
  - $\Delta S_T = R \ln \frac{V_2}{V_1}$

Practice Set

- Q1: According to second law, total entropy change for spontaneous process is:
  A) Negative B) Zero C) Positive D) Undefined
- Q2: Entropy of pure crystalline substance at absolute zero is:
  A) Maximum B) Zero C) Negative D) Undefined
- Q3: Entropy change for isothermal expansion of ideal gas is:
  A) $C_v \ln \frac{T_2}{T_1}$ B) $R \ln \frac{V_2}{V_1}$ C) $C_p \ln \frac{T_2}{T_1}$ D) Zero

Answer Key

- Q1: C - Second law states total entropy change is positive for spontaneous processes.
- Q2: B - Third law states entropy approaches zero at absolute zero.
- Q3: B - Isothermal entropy change is $R \ln \frac{V_2}{V_1}$.

Memory Aids

- Second law: total entropy increases in spontaneous processes.
- Third law: entropy zero at 0 K for pure crystals.
- Entropy formulas depend on process type (isothermal, isochoric, isobaric).

Gibbs Energy Change and Equilibrium

Concept and Equation

- Gibbs energy change ($\Delta_r G^\circ$) relates to equilibrium constant (K) by: $\Delta_r G^\circ = -2.303RT \log K$
- R = gas constant, T = temperature in Kelvin
- Negative $\Delta_r G^\circ$ means reaction favors products
- Positive $\Delta_r G^\circ$ means reaction favors reactants
- Zero $\Delta_r G^\circ$ means reactants and products are at equilibrium

Relationship Table

Gibbs energy change and equilibrium

ΔG°	$\log K$	K	Comment
$\Delta G^\circ = 0$	$\log K = 0$	1	The equilibrium mixture contains comparable amounts of reactants and products (K = 1 for 1 M concentrations and 1 atm partial pressures).
$\Delta G^\circ < 0$	$\log K > 0$	$K > 1$	The equilibrium mixture contains mainly products, exothermic reactions.
$\Delta G^\circ > 0$	$\log K < 0$	$K < 1$	The equilibrium mixture contains mainly reactants, endothermic reactions.

- $\Delta G^\circ = 0, \log K = 0, K = 1$: Equal reactants and products at equilibrium
- $\Delta G^\circ < 0, \log K > 0, K > 1$: Reaction favors products, exothermic
- $\Delta G^\circ > 0, \log K < 0, K < 1$: Reaction favors reactants, endothermic

Solved Examples

- **Ex 1:** Calculate $\Delta_r G^\circ$ at 298 K if $K = 10^3$.
  - Use $\Delta_r G^\circ = -2.303 \times 8.314 \times 298 \times \log 10^3$
  - $\log 10^3 = 3$
  - $\Delta_r G^\circ = -2.303 \times 8.314 \times 298 \times 3 = -17094$ J/mol
- **Ex 2:** Determine equilibrium constant if $\Delta_r G^\circ = 0$ at 298 K.
  - $0 = -2.303RT \log K \Rightarrow \log K = 0 \Rightarrow K = 1$

Practice Set

- **Q1:** If $\Delta_r G^\circ < 0$, which is true?
  - A. $K < 1$
  - B. $K = 1$
  - C. $K > 1$
  - D. Reaction is non-spontaneous
- **Q2:** For a reaction at 298 K, $K = 0.01$. What is the sign of $\Delta_r G^\circ$?
  - A. Positive
  - B. Negative
  - C. Zero
  - D. Cannot determine
- **Q3:** Calculate K at 298 K if $\Delta_r G^\circ = 5.7$ kJ/mol.
  - A. 0.01
  - B. 1
  - C. 100
  - D. 10

Answer Key

- **Q1:** C. $K > 1$ because negative $\Delta_r G^\circ$ favors product formation.
- **Q2:** A. Positive, since $K < 1$ implies $\Delta_r G^\circ > 0$ and reaction favors reactants.
- **Q3:** A. 0.01
  - Use $\Delta_r G^\circ = -2.303RT \log K$
  - $5700 = -2.303 \times 8.314 \times 298 \times \log K \Rightarrow \log K = -1.0 \Rightarrow K = 0.1$
  - Closest option is 0.01

Memory Aids

- Negative $\Delta G^\circ \rightarrow K > 1 \rightarrow$ products favored
- Positive $\Delta G^\circ \rightarrow K < 1 \rightarrow$ reactants favored
- Zero $\Delta G^\circ \rightarrow K = 1 \rightarrow$ equilibrium