Reactivity 1: What Drives Chemical Reactions?

IB SL Study Guide

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IB TIP

Data booklet: You can use the IB Chemistry Data Booklet in the exam — all constants, the periodic table, and key equations are provided.

IB Chemistry SL — Thermodynamics & Gibbs Free Energy

Complete Study Guide

Topics Covered

Videos on this page: Enthalpy Changes & Calorimetry

1. Energy Changes in Chemical Reactions (R1.1)

Every chemical reaction involves energy. When bonds break, energy is absorbed; when bonds form, energy is released. The net balance of these two processes determines whether a reaction releases energy to its surroundings or absorbs energy from them. This is one of the most fundamental ideas in chemistry, and it underpins everything from why fuels burn to why ice melts.

Exothermic and Endothermic Reactions

	Exothermic	Endothermic
Energy change	Energy released to surroundings	Energy absorbed from surroundings
Temperature of surroundings	Increases	Decreases
Sign of ΔH	Negative ($\Delta H < 0$)	Positive ($\Delta H > 0$)
Bond energy balance	Energy released forming bonds > Energy absorbed breaking bonds	Energy absorbed breaking bonds > Energy released forming bonds
Examples	Combustion, neutralisation, most oxidation reactions	Photosynthesis, thermal decomposition, dissolving ammonium nitrate

MEMORISE THIS

Exothermic = Exit — energy exits the system into the surroundings ($\Delta H < 0$, temperature rises).

Endothermic = Enter — energy enters the system from the surroundings ($\Delta H > 0$, temperature falls).

Enthalpy Diagrams

Enthalpy diagrams (also called energy profile diagrams) show the enthalpy of the reactants and products on a vertical energy axis, with the reaction pathway on the horizontal axis.

Exothermic reaction:

- Reactants are at a **higher** enthalpy level than products.
- The arrow for ΔH points **downwards**.
- There is an activation energy (E_a) hump that must be overcome for the reaction to start.

Endothermic reaction:

- Reactants are at a **lower** enthalpy level than products.
- The arrow for ΔH points **upwards**.
- The activation energy hump is measured from the reactant level to the top of the curve.

EXAM ALERT

When drawing enthalpy diagrams, always label: (1) the y-axis as "Enthalpy / kJ", (2) the x-axis as "Reaction pathway" or "Progress of reaction", (3) reactants and products, (4) ΔH with its sign, and (5) E_a (activation energy). Missing labels lose marks.

Activation Energy

Activation energy (E_a) is the **minimum energy** that colliding particles must possess for a reaction to occur. Even exothermic reactions need activation energy to get started — think of a match: once lit (activation energy supplied), the combustion reaction releases far more energy than was initially needed.

A catalyst lowers the activation energy by providing an alternative reaction pathway. It does **not** change ΔH — the enthalpy difference between reactants and products stays the same.

2. Calorimetry — Measuring Enthalpy Changes

Calorimetry is the experimental technique used to measure enthalpy changes. The basic idea is simple: carry out a reaction in or near water, measure the temperature change of the water, and use this to calculate the energy transferred.

The Calorimetry Equation

$$q = mc\Delta T$$

where:

- q = energy transferred (in joules, J)
- m = mass of the solution being heated (in grams, g) — usually assumed to be the mass of water
- c = specific heat capacity (for water, $c = 4.18$ J g^{-1}K^{-1})
- ΔT = temperature change (in K or °C — the magnitude is the same)

WORKED EXAMPLE

Example: 50.0 cm^3 of 1.00 mol dm^{-3} HCl is mixed with 50.0 cm^3 of 1.00 mol dm^{-3} NaOH. The temperature rises from 22.0°C to 28.8°C. Calculate the enthalpy of neutralisation per mole.

Step 1: Total volume = 100.0 cm^3. Assume density = 1.00 g cm^{-3}, so $m = 100.0$ g.

Step 2: $\Delta T = 28.8 - 22.0 = 6.8$ K

Step 3: $q = mc\Delta T = 100.0 \times 4.18 \times 6.8 = 2842$ J $= 2.84$ kJ

Step 4: Moles of HCl = $0.0500 \times 1.00 = 0.0500$ mol

Step 5: $\Delta H = -\dfrac{q}{n} = -\dfrac{2.84}{0.0500} = -56.8$ kJ mol^{-1}

The negative sign indicates the reaction is exothermic (temperature rose).

EXAM ALERT

Common calorimetry mistakes: (1) Forgetting to make ΔH negative for exothermic reactions — if the temperature goes UP, ΔH is NEGATIVE. (2) Using the mass of the solute instead of the mass of the solution/water. (3) Forgetting to convert J to kJ (divide by 1000). (4) Not dividing by moles to get kJ mol^{-1}.

Sources of Error in Calorimetry

In practice, experimental ΔH values are always less accurate than literature values because:

- **Heat loss to surroundings** — the calorimeter is not perfectly insulated.

- **Incomplete combustion** (for combustion experiments) — soot formation means not all fuel reacted completely.
- **Assumption that the solution has the density and specific heat capacity of water** — dilute solutions are close, but not exact.
- **Evaporation of volatile substances** — some fuel or solvent may evaporate.

▸**Watch: Enthalpy Changes & Calorimetry** VIDEO

3. Enthalpy of Combustion, Formation, and Neutralisation

The IB syllabus requires you to know several standard enthalpy definitions. All are measured under **standard conditions**: 298 K (25°C), 100 kPa, solutions at 1 mol dm^{-3}. The symbol is ΔH° (the superscript $^\circ$ means standard).

MEMORISE THIS

Standard enthalpy of combustion (ΔH_c°): The enthalpy change when **one mole** of a substance is completely burned in excess oxygen under standard conditions. Always **negative** (exothermic).

Standard enthalpy of formation (ΔH_f°): The enthalpy change when **one mole** of a compound is formed from its elements in their standard states under standard conditions. Can be positive or negative.

Standard enthalpy of neutralisation (ΔH_n°): The enthalpy change when an acid and base react to form **one mole of water** under standard conditions. Approximately -57.1 kJ mol^{-1} for strong acid + strong base.

EXAM ALERT

The definitions always specify "one mole" — this is critical. If you burn 2 moles of methane, the energy released is $2 \times \Delta H_c^\circ$, but the standard enthalpy of combustion is defined per mole. When writing thermochemical equations, make sure the coefficient of the substance being defined matches "1 mole." For example:

$$\mathrm{CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)} \quad \Delta H_c^\circ = -890 \text{ kJ mol}^{-1}$$

Key fact: The standard enthalpy of formation of any element in its standard state is **zero** by definition. For example, $\Delta H_f^\circ[\mathrm{O_2(g)}] = 0$, $\Delta H_f^\circ[\mathrm{C(graphite)}] = 0$.

4. Hess's Law & Energy Cycles (R1.2)

What is Hess's Law?

Hess's Law states that **the total enthalpy change for a reaction is independent of the route taken**, provided the initial and final conditions are the same. In plain language: it

does not matter whether a reaction happens in one step or in five steps — the total energy change is the same.

This is a consequence of the law of conservation of energy and is enormously useful because it lets us calculate enthalpy changes for reactions that are difficult or impossible to measure directly.

Using Enthalpy of Formation to Calculate Reaction Enthalpy

$$\Delta H^\circ_{\text{rxn}} = \sum \Delta H^\circ_f(\text{products}) - \sum \Delta H^\circ_f(\text{reactants})$$

In words: add up the enthalpies of formation of all the products, subtract the sum of the enthalpies of formation of all the reactants.

WORKED EXAMPLE

Example: Calculate ΔH° for the reaction:

$$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$$

Given: $\Delta H^\circ_f[\text{CH}_4(\text{g})] = -74.8$, $\Delta H^\circ_f[\text{CO}_2(\text{g})] = -393.5$, $\Delta H^\circ_f[\text{H}_2\text{O}(\text{l})] = -285.8$ (all in kJ mol^{-1}).

Solution:

$$\Delta H^\circ_{\text{rxn}} = \left[\Delta H^\circ_f(\text{CO}_2) + 2 \times \Delta H^\circ_f(\text{H}_2\text{O})\right] - \left[\Delta H^\circ_f(\text{CH}_4) + 2 \times \Delta H^\circ_f(\text{O}_2)\right]$$

$$= \left[-393.5 + 2(-285.8)\right] - \left[-74.8 + 0\right]$$

$$= (-393.5 - 571.6) - (-74.8) = -965.1 + 74.8 = -890.3 \text{ kJ mol}^{-1}$$

The large negative value confirms this is a highly exothermic combustion reaction.

Using Enthalpy of Combustion to Calculate Reaction Enthalpy

$$\Delta H^\circ_{\text{rxn}} = \sum \Delta H^\circ_c(\text{reactants}) - \sum \Delta H^\circ_c(\text{products})$$

EXAM ALERT

Notice the sign is **reversed** compared to the formation route: for combustion data, it is **reactants minus products**. A common mistake is applying the same "products minus reactants" rule to both — this only works for formation data. For combustion data, think of it as: the enthalpy cycle goes *down* from reactants via combustion, and *down* from products via combustion, to the same set of combustion products.

WORKED EXAMPLE

Example: Calculate ΔH° for the reaction:

$$\mathrm{C(s) + 2H_2(g) \rightarrow CH_4(g)}$$

Given: $\Delta H_c^\circ[\mathrm{C(s)}] = -393.5$, $\Delta H_c^\circ[\mathrm{H_2(g)}] = -285.8$, $\Delta H_c^\circ[\mathrm{CH_4(g)}] = -890.3$ (all in kJ $\mathrm{mol^{-1}}$).

Solution:

$$\Delta H^\circ_{\text{rxn}} = \left[\Delta H_c^\circ(\mathrm{C}) + 2 \times \Delta H_c^\circ(\mathrm{H_2})\right] - \left[\Delta H_c^\circ(\mathrm{CH_4})\right]$$

$$= \left[-393.5 + 2(-285.8)\right] - \left[-890.3\right]$$

$$= -965.1 + 890.3 = -74.8 \text{ kJ mol}^{-1}$$

This equals the standard enthalpy of formation of methane, as expected from the equation.

Practice: Hess's Law — Fading Sequence

These problems build your confidence step by step. The first shows the complete solution, the second hides the final calculation, and the third gives only the setup.

WORKED EXAMPLE *Full worked example — all steps shown*

Calculate ΔH° for: $\mathrm{2C(s) + 3H_2(g) + \frac{1}{2}O_2(g) \rightarrow C_2H_5OH(l)}$ using combustion data.

Step 1

Write the formula: $\Delta H^\circ_{\text{rxn}} = \sum \Delta H_c^\circ(\text{reactants}) - \sum \Delta H_c^\circ(\text{products})$

Step 2

List the combustion enthalpies: $\Delta H_c^\circ[\mathrm{C(s)}] = -393.5$, $\Delta H_c^\circ[\mathrm{H_2(g)}] = -285.8$, $\Delta H_c^\circ[\mathrm{C_2H_5OH(l)}] = -1367.3$ kJ $\mathrm{mol^{-1}}$. Note: $\mathrm{O_2}$ does not combust, so it has no ΔH_c° term.

Step 3

Substitute: $\Delta H^\circ_{\text{rxn}} = \left[2(-393.5) + 3(-285.8)\right] - \left[-1367.3\right]$

Step 4

Calculate: $= (-787.0 - 857.4) - (-1367.3) = -1644.4 + 1367.3 = -277.1 \text{ kJ mol}^{-1}$

This is the enthalpy of formation of ethanol.

YOUR TURN (PARTIAL) *Partial example — try the last step yourself*

Calculate ΔH° for: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ using formation data.

Steps 1–2 are shown. Try the calculation yourself, then reveal.

Step 1

Write the formula: $\Delta H^\circ_{\text{rxn}} = \sum \Delta H^\circ_f(\text{products}) - \sum \Delta H^\circ_f(\text{reactants})$

Step 2

Identify the values: $\Delta H^\circ_f[NH_3(g)] = -46.1$ kJ mol^{-1}. Both N_2 and H_2 are elements in their standard states, so $\Delta H^\circ_f = 0$.

Try it yourself, then click to reveal — Step 3 — Substitute into the formula and calculate

YOUR TURN (SCAFFOLDED) *Scaffolded — only the setup is given*

Use bond enthalpies to estimate ΔH for: $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$.

Bond enthalpies: H-H = 436, Cl-Cl = 242, H-Cl = 431 kJ mol^{-1}. Work through the full calculation, then reveal each step to check.

Try it yourself, then click to reveal — Step 1 — List bonds broken and bonds formed
Try it yourself, then click to reveal — Step 2 — Apply the formula and calculate

5. Bond Enthalpies

What are Bond Enthalpies?

A bond enthalpy (also called bond energy) is the energy required to break **one mole** of a specific covalent bond in the gaseous state, averaged over many different molecules. Because they are averages, calculations using bond enthalpies give **approximate** values for ΔH.

Breaking bonds is **endothermic** (requires energy input). Forming bonds is **exothermic** (releases energy).

Calculating ΔH from Bond Enthalpies

$\Delta H \approx \sum(\text{bond enthalpies of bonds broken}) - \sum(\text{bond enthalpies of bonds formed})$

In words: add up the energy needed to break all bonds in the reactants, then subtract the energy released when all bonds in the products form.

MEMORISE THIS

"Break minus Make" — break is positive (costs energy), make is negative (releases energy). If breaking costs more than making releases, the reaction is endothermic. If making releases more than breaking costs, the reaction is exothermic.

WORKED EXAMPLE

Example: Use bond enthalpies to estimate ΔH for the combustion of methane:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

Bond enthalpies: C-H = 414, O=O = 498, C=O (in CO_2) = 804, O-H = 463 (all in kJ mol^{-1}).

Bonds broken (reactants):

- $4 \times$ C-H $= 4 \times 414 = 1656$
- $2 \times$ O=O $= 2 \times 498 = 996$
- **Total broken = 2652 kJ**

Bonds formed (products):

- $2 \times$ C=O $= 2 \times 804 = 1608$
- $4 \times$ O-H $= 4 \times 463 = 1852$
- **Total formed = 3460 kJ**

$$\Delta H \approx 2652 - 3460 = -808 \text{ kJ mol}^{-1}$$

The value is approximate because bond enthalpies are averages. The data booklet value (-890 kJ mol^{-1}) is more accurate.

EXAM ALERT

Bond enthalpy calculations give **approximate** answers because the values are averaged over many molecules. The IB expects you to explain this when asked why calculated values differ from experimental values. Also: bond enthalpy data applies only to **gaseous** substances. If water is produced as a liquid, the bond enthalpy method does not account for the energy of condensation.

Why Bond Enthalpy Calculations are Only Estimates

1. Bond enthalpies are **average values** — the actual energy of a C-H bond in methane is slightly different from a C-H bond in ethane.
2. They apply to the **gaseous state** only — if reactants or products are liquids or solids, additional energy changes (vaporisation, condensation) are not included.
3. In molecules with resonance or delocalisation (like benzene), the actual bond strengths differ from the tabulated average.

6. Entropy & the Second Law (R1.3)

What is Entropy?

Entropy (S) is a measure of the **disorder** or **dispersal of energy** in a system. A system with high entropy has many possible arrangements of its particles and energy — it is

more “spread out” or disordered. A system with low entropy is highly ordered with fewer possible arrangements.

Think of it this way: a tidy room has low entropy (everything in its place, only one arrangement). A messy room has high entropy (items scattered in many possible positions). Nature tends towards the messy room — systems naturally move towards greater disorder.

Predicting Entropy Changes

You can predict whether entropy increases or decreases by looking at:

Change	Effect on entropy
Solid → Liquid → Gas	Entropy **increases** (particles become more disordered)
Fewer moles of gas → More moles of gas	Entropy **increases**
More moles of gas → Fewer moles of gas	Entropy **decreases**
Dissolving a solid in water	Entropy usually **increases** (particles become more dispersed)
Temperature increases	Entropy **increases** (particles have more energy, more possible arrangements)

MEMORISE THIS

Quick rules for entropy change (ΔS):

- More gas molecules in products than reactants $\Rightarrow \Delta S > 0$ (positive)
- Fewer gas molecules in products than reactants $\Rightarrow \Delta S < 0$ (negative)
- Change of state from solid/liquid to gas $\Rightarrow$ large positive ΔS
- Dissolving $\Rightarrow$ usually positive ΔS

Calculating Entropy Changes

$$\Delta S^\circ_{\text{rxn}} = \sum S^\circ(\text{products}) - \sum S^\circ(\text{reactants})$$

Note: unlike enthalpy of formation, the standard entropy of elements is **not** zero. Every substance has a positive standard entropy value (the third law of thermodynamics states that entropy is zero only at absolute zero, 0 K).

WORKED EXAMPLE

Example: Calculate ΔS° for the decomposition of calcium carbonate:

$$\mathrm{CaCO_3(s) \rightarrow CaO(s) + CO_2(g)}$$

Given: $S^\circ[\mathrm{CaCO_3(s)}] = 92.9$, $S^\circ[\mathrm{CaO(s)}] = 39.7$, $S^\circ[\mathrm{CO_2(g)}] = 213.7$ (all in J K^{-1} mol^{-1}).

Solution:

$$\Delta S^\circ = [39.7 + 213.7] - [92.9] = 253.4 - 92.9 = +160.5 \text{ J K}^{-1}\text{mol}^{-1}$$

The large positive value makes sense: a gas ($\mathrm{CO_2}$) is produced from a solid, greatly increasing disorder.

EXAM ALERT

Entropy is measured in **J K^{-1} mol^{-1}** (joules, not kilojoules!). When using entropy in Gibbs free energy calculations, you must convert to kJ by dividing by 1000, or convert ΔH to J by multiplying by 1000. Mixing up units is one of the most common errors in thermodynamics calculations.

7. Gibbs Free Energy & Spontaneity (R1.3)

What is Gibbs Free Energy?

Gibbs free energy (G) combines enthalpy and entropy into a single quantity that tells us whether a reaction is **spontaneous** (thermodynamically favourable) at a given temperature. A spontaneous reaction is one that can occur without continuous external input of energy — but "spontaneous" does **not** mean "fast." Diamond converting to graphite is spontaneous but takes millions of years.

The Gibbs Equation

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

where:

- ΔG = Gibbs free energy change (kJ mol^{-1})
- ΔH = enthalpy change (kJ mol^{-1})
- T = temperature in **kelvin** (K = °C + 273)
- ΔS = entropy change (must be in **kJ** K^{-1} mol^{-1} — divide J values by 1000)

MEMORISE THIS

Spontaneity rules:

- $\Delta G < 0$: reaction is **spontaneous** (thermodynamically favourable) in the forward direction

- $\Delta G > 0$: reaction is **non-spontaneous** — the reverse reaction is spontaneous
- $\Delta G = 0$: system is at **equilibrium**

The Four Combinations of ΔH and ΔS

ΔH	ΔS	$\Delta G = \Delta H - T\Delta S$	Spontaneous?
Negative (exothermic)	Positive (more disorder)	Always negative	**Always spontaneous** at all temperatures
Negative (exothermic)	Negative (less disorder)	Depends on T	**Spontaneous at low** T (enthalpy wins)
Positive (endothermic)	Positive (more disorder)	Depends on T	**Spontaneous at high** T (entropy wins)
Positive (endothermic)	Negative (less disorder)	Always positive	**Never spontaneous** at any temperature

⚠ EXAM ALERT

The IB frequently asks: "At what temperature does this reaction become spontaneous?" Set $\Delta G = 0$ and solve for T:

$$0 = \Delta H - T\Delta S \quad \Rightarrow \quad T = \frac{\Delta H}{\Delta S}$$

This gives the temperature at which the reaction switches between spontaneous and non-spontaneous. Make sure ΔH and ΔS use the **same energy units** (both kJ or both J).

WORKED EXAMPLE

Example: For the decomposition of calcium carbonate:

$$\mathrm{CaCO_3(s) \rightarrow CaO(s) + CO_2(g)}$$

$\Delta H^\circ = +178 \text{ kJ mol}^{-1}$, $\Delta S^\circ = +160.5 \text{ J K}^{-1}\text{mol}^{-1}$

(a) Is this reaction spontaneous at 298 K?

$$\Delta G = 178 - (298 \times 0.1605) = 178 - 47.8 = +130.2 \text{ kJ mol}^{-1}$$

$\Delta G > 0$, so the reaction is **not spontaneous** at 298 K.

(b) At what temperature does it become spontaneous?

$$T = \frac{\Delta H}{\Delta S} = \frac{178}{0.1605} = 1109 \text{ K} \approx 836^\circ\text{C}$$

Above 1109 K (836°C), the entropy term ($T\Delta S$) becomes large enough to overcome the positive ΔH, making ΔG negative.

WORKED EXAMPLE

Example: The Haber process:

$\mathrm{N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)}$

$\Delta H^\circ = -92.2\ \mathrm{kJ\ mol^{-1}}$, $\Delta S^\circ = -198.7\ \mathrm{J\ K^{-1} mol^{-1}}$

This is an exothermic reaction ($\Delta H < 0$) with a decrease in entropy ($\Delta S < 0$, because 4 moles of gas become 2 moles). At low temperatures, $\Delta G < 0$ (spontaneous). At high temperatures, the $-T\Delta S$ term becomes a large positive number that overcomes the negative ΔH.

$T = \frac{-92.2}{-0.1987} = 464\ \mathrm{K} \approx 191^\circ\mathrm{C}$

Above 464 K, $\Delta G > 0$ and the forward reaction is non-spontaneous. This is why the Haber process uses a compromise temperature (around 450°C) — high enough for a reasonable rate but near the thermodynamic limit.

Spontaneity vs. Rate

IB TIP

A spontaneous reaction ($\Delta G < 0$) is NOT necessarily a fast reaction. Spontaneity tells you about the **thermodynamic tendency** — whether the reaction is energetically favourable. The **rate** depends on the activation energy and kinetics (a separate topic). For example:

- Diamond → graphite: $\Delta G < 0$ (spontaneous) but extremely slow at room temperature.
- Combustion of petrol: $\Delta G \ll 0$ (highly spontaneous) but does not occur without a spark (activation energy).

8. Quick Reference — Key Formulas and Sign Conventions

MEMORISE THIS

Essential Equations:

Equation	Use
$q = mc\Delta T$	Calorimetry — calculating energy transferred
$\Delta H = -\frac{q}{n}$	Converting calorimetry data to molar enthalpy (negative sign for exothermic)
$\Delta H^\circ_{\text{rxn}} = \sum \Delta H^\circ_f(\text{products}) - \sum \Delta H^\circ_f(\text{reactants})$	Hess's law using formation data
$\Delta H^\circ_{\text{rxn}} = \sum \Delta H^\circ_c(\text{reactants}) - \sum \Delta H^\circ_c(\text{products})$	Hess's law using combustion data
$\Delta H \approx \sum(\text{bonds broken}) - \sum(\text{bonds formed})$	Bond enthalpy calculation
$\Delta S^\circ_{\text{rxn}} = \sum S^\circ(\text{products}) - \sum S^\circ(\text{reactants})$	Entropy change
$\Delta G = \Delta H - T\Delta S$	Gibbs free energy
$T = \frac{\Delta H}{\Delta S}$	Temperature at which $\Delta G = 0$ (spontaneity switch)

MEMORISE THIS

Sign Convention Cheat Sheet:

Quantity	Positive means…	Negative means…
ΔH	Endothermic (heat absorbed)	Exothermic (heat released)
ΔS	More disorder (favourable)	Less disorder (unfavourable)
ΔG	Non-spontaneous	Spontaneous (favourable)
ΔT in calorimetry	Temperature rose (exothermic reaction)	Temperature fell (endothermic reaction)

Mixed Practice — Exam Style

IB TIP

How to use this section: Unlike topic-specific practice, these questions are interleaved — they mix all topics from this guide in random order. Before answering, identify *which concept or topic area* the question is testing. This is exactly the skill you need on Paper 1 and Paper 2, where you don't know in advance which topic each question covers.

1. **[Exothermic/Endothermic]** The dissolution of ammonium nitrate (NH_4NO_3) in water causes the temperature of the solution to decrease. Which row correctly describes this process?

	ΔH	Energy transfer
A	Negative	Energy released to surroundings
B	Positive	Energy absorbed from surroundings
C	Negative	Energy absorbed from surroundings
D	Positive	Energy released to surroundings

2. **[Calorimetry]** When 0.500 g of ethanol ($\mathrm{C_2H_5OH}$, $M_r = 46.08$) is burned and the heat is used to warm 200.0 g of water from 20.0°C to 33.2°C, the experimental enthalpy of combustion of ethanol is approximately:

   A. $-1020\ \mathrm{kJ\ mol^{-1}}$

   B. $-509\ \mathrm{kJ\ mol^{-1}}$

   C. $+1020\ \mathrm{kJ\ mol^{-1}}$

   D. $-11.0\ \mathrm{kJ\ mol^{-1}}$

3. **[Hess's Law — Formation Data]** Given the following standard enthalpies of formation:

   $\Delta H_f^\circ[\mathrm{SO_3(g)}] = -395.7\ \mathrm{kJ\ mol^{-1}}$, $\Delta H_f^\circ[\mathrm{SO_2(g)}] = -296.8\ \mathrm{kJ\ mol^{-1}}$

   Calculate ΔH° for: $\mathrm{2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)}$

   A. $-98.9\ \mathrm{kJ}$

   B. $-197.8\ \mathrm{kJ}$

   C. $+197.8\ \mathrm{kJ}$

   D. $-692.5\ \mathrm{kJ}$

4. **[Bond Enthalpies]** The bond enthalpy of H-H is 436 kJ mol^{-1} and Cl-Cl is 242 kJ mol^{-1}. The bond enthalpy of H-Cl is 431 kJ mol^{-1}. What is the approximate ΔH for $\mathrm{H_2(g) + Cl_2(g) \rightarrow 2HCl(g)}$?

   A. $-184\ \mathrm{kJ}$

   B. $+184\ \mathrm{kJ}$

   C. $-247\ \mathrm{kJ}$

   D. $-862\ \mathrm{kJ}$

5. **[Entropy — Predicting Sign]** For which reaction is ΔS most likely to be large and positive?

   A. $\mathrm{H_2O(l) \rightarrow H_2O(s)}$

   B. $\mathrm{N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)}$

   C. $\mathrm{CaCO_3(s) \rightarrow CaO(s) + CO_2(g)}$

   D. $\mathrm{2CO(g) + O_2(g) \rightarrow 2CO_2(g)}$

6. **[Gibbs Free Energy — Calculation]** A reaction has $\Delta H = -120\ \mathrm{kJ\ mol^{-1}}$ and $\Delta S = +45\ \mathrm{J\ K^{-1} mol^{-1}}$ at 298 K. What is ΔG?

A. $-133.4\ \text{kJ mol}^{-1}$

B. $-106.6\ \text{kJ mol}^{-1}$

C. $+106.6\ \text{kJ mol}^{-1}$

D. $-13410\ \text{kJ mol}^{-1}$

7. **[Gibbs — Spontaneity at Temperature]** A reaction has $\Delta H = +50.0\ \text{kJ mol}^{-1}$ and $\Delta S = +120\ \text{J K}^{-1}\text{mol}^{-1}$. Above what temperature is the reaction spontaneous?

   A. 417 K

   B. 240 K

   C. 0.42 K

   D. 6000 K

8. **[Enthalpy Diagrams]** In an enthalpy level diagram for an exothermic reaction with a catalyst, which statement is correct?

   A. The catalyst lowers both the activation energy and ΔH

   B. The catalyst lowers the activation energy but does not change ΔH

   C. The catalyst raises the activation energy but makes the reaction more exothermic

   D. The catalyst has no effect on the activation energy

9. **[Calorimetry — Sources of Error]** A student measures the enthalpy of combustion of propan-1-ol using a spirit burner and gets a value significantly less negative than the literature value. The most likely reason is:

   A. The thermometer was reading too high

   B. Heat was lost to the surroundings, so not all energy was transferred to the water

   C. The student used too much water

   D. The specific heat capacity of water was too low

10. **[Gibbs — Four Combinations]** A reaction has $\Delta H > 0$ and $\Delta S < 0$. Which statement about spontaneity is correct?

    A. The reaction is spontaneous at all temperatures

    B. The reaction is spontaneous at high temperatures

    C. The reaction is spontaneous at low temperatures

D. The reaction is never spontaneous at any temperature

11. **[Hess's Law — Combustion Data]** Given:

    $\Delta H_c^\circ[\mathrm{C(s)}] = -393.5\ \mathrm{kJ\ mol^{-1}}$, $\Delta H_c^\circ[\mathrm{H_2(g)}] = -285.8\ \mathrm{kJ\ mol^{-1}}$, $\Delta H_c^\circ[\mathrm{C_2H_6(g)}] = -1560.7\ \mathrm{kJ\ mol^{-1}}$

    Calculate ΔH_f° for ethane: $2\mathrm{C(s)} + 3\mathrm{H_2(g)} \rightarrow \mathrm{C_2H_6(g)}$

    A. $-84.7\ \mathrm{kJ\ mol^{-1}}$

    B. $+84.7\ \mathrm{kJ\ mol^{-1}}$

    C. $-881.4\ \mathrm{kJ\ mol^{-1}}$

    D. $-1560.7\ \mathrm{kJ\ mol^{-1}}$

12. **[Entropy Calculation]** Given: $S^\circ[\mathrm{N_2(g)}] = 191.6$, $S^\circ[\mathrm{H_2(g)}] = 130.7$, $S^\circ[\mathrm{NH_3(g)}] = 192.5$ (all in J K^{-1} mol^{-1}). What is ΔS° for $\mathrm{N_2(g)} + 3\mathrm{H_2(g)} \rightarrow 2\mathrm{NH_3(g)}$?

    A. $-198.7\ \mathrm{J\ K^{-1}mol^{-1}}$

    B. $+198.7\ \mathrm{J\ K^{-1}mol^{-1}}$

    C. $-129.8\ \mathrm{J\ K^{-1}mol^{-1}}$

    D. $+61.9\ \mathrm{J\ K^{-1}mol^{-1}}$

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