MA99 EDUCATION

Drivers of Reactions: Topic Notes 04

NSW CHEMISTRY STAGE 6 SYLLABUS 2017 FOR HSC 2028

1. Exothermic and endothermic reactions

Every reaction involves an energy change. Enthalpy (H) is the heat content of a substance, and the enthalpy change (ΔH) is the heat released or absorbed at constant pressure.

- **Exothermic:** releases heat to the surroundings, so ΔH is negative. Combustion and neutralisation are exothermic.
- **Endothermic:** absorbs heat from the surroundings, so ΔH is positive. Photosynthesis and the decomposition of limestone are endothermic.

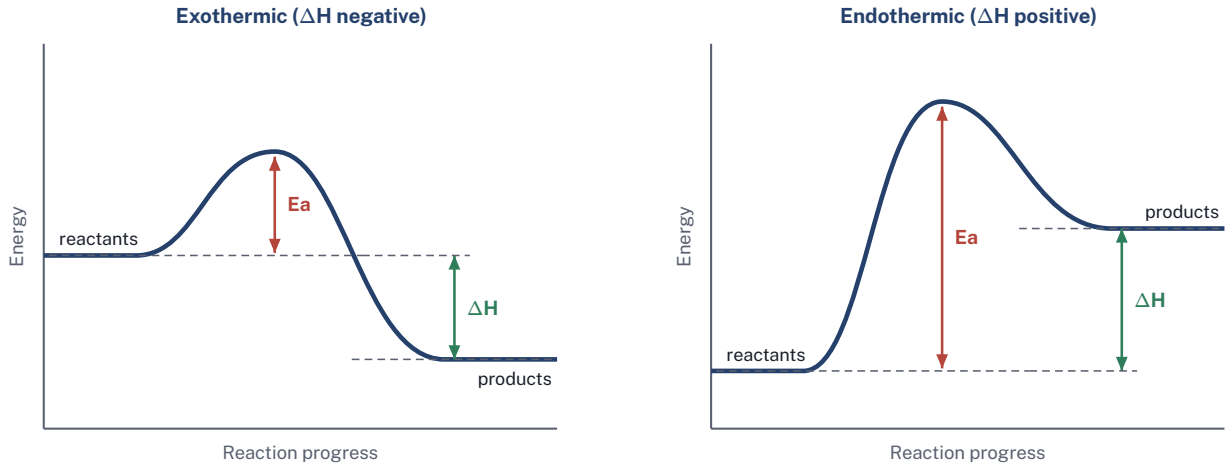


Figure 1 · In an exothermic reaction the products lie below the reactants; in an endothermic reaction they lie above. The peak is the activation energy.

2. Measuring enthalpy: calorimetry

Calorimetry measures the heat released by a reaction from the temperature change it causes in a known mass of water, using:

The calorimetry formula

$$q = m\,c\,\Delta T$$

q = heat energy in joules (J); m = mass of water in grams (g); c = specif ic heat capacity of water = 4.18 J/g/K; ΔT = change in temperature (degrees C or K)

Worked example: heat of combustion

A spirit burner burning 0.500 g of ethanol raises the temperature of 200 g of water by 15.0 degrees C. Find the heat released, then the molar heat of combustion.

$$q = m\,c\,\Delta T = 200 \times 4.18 \times 15.0 = 12540\ \text{J} = 12.54\ \text{kJ}$$

$$n(\text{ethanol}) = m / M = 0.500 / 46.07 = 0.01085\ \text{mol}$$

$$\Delta H = -q / n = -12.54 / 0.01085 = -1160\ \text{kJ/mol}\ (3\ \text{significant figures})$$

The value is negative because combustion is exothermic. It is less than the accepted value because heat is lost to the surroundings.

Worked example: heat of neutralisation (hypothetical (MA99) data)

50.0 mL of 1.00 mol/L hydrochloric acid is mixed with 50.0 mL of 1.00 mol/L sodium hydroxide in an insulated cup. The temperature rises by 6.7 degrees C. Take the solution as 100 g with the specific heat capacity of water.

$$q = m\,c\,\Delta T = 100 \times 4.18 \times 6.7 = 2.8 \times 10^3\ \text{J} = 2.8\ \text{kJ}$$

$$n(\text{HCl}) = 1.00 \times 0.0500 = 0.0500\ \text{mol, which reacts with } 0.0500\ \text{mol NaOH}$$

$$\Delta H = -q / n = -2.8 / 0.0500 = -56\ \text{kJ/mol}$$

The heat is divided by the moles of **water formed** (0.0500 mol here), not by the total volume. The result is close to the standard value for strong acid and strong base, about −56 kJ/mol. Heat losses and the assumptions about density and heat capacity explain small differences.

3. Energy profiles and catalysts

An energy profile diagram plots energy against reaction progress. The difference between reactants and the peak is the activation energy, and the difference between reactants and products is ΔH.

A catalyst provides an alternative pathway with a lower activation energy, so the reaction is faster, but it does not change ΔH, because the energy of the reactants and products is unchanged.

4. Bond enthalpy

Breaking bonds absorbs energy and forming bonds releases energy. The overall enthalpy change is the energy to break the reactant bonds minus the energy released when the product bonds form:

Key idea: bond enthalpy

$$\Delta H = (\text{bonds broken}) \text{ minus } (\text{bonds formed})$$

Use bond enthalpies (kJ/mol) from a data table. A negative answer means the reaction is exothermic.

Worked example: hydrogen and chlorine

For $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$, using H-H = 436, Cl-Cl = 242 and H-Cl = 431 kJ/mol:

$$\text{bonds broken} = 436 + 242 = 678\ \text{kJ/mol}$$

$$\text{bonds formed} = 2 \times 431 = 862\ \text{kJ/mol}$$

$$\Delta H = 678 - 862 = -184\ \text{kJ/mol (exothermic)}$$

5. Hess's law

Hess's law states that the enthalpy change of a reaction is the same whether it happens in one step or several, because enthalpy is a state function. This lets us add known reactions to find an unknown ΔH.

When an equation is reversed, the sign of ΔH is reversed. When an equation is multiplied by a number, ΔH is multiplied by the same number. Adding the steps of an energy cycle gives the overall enthalpy change.

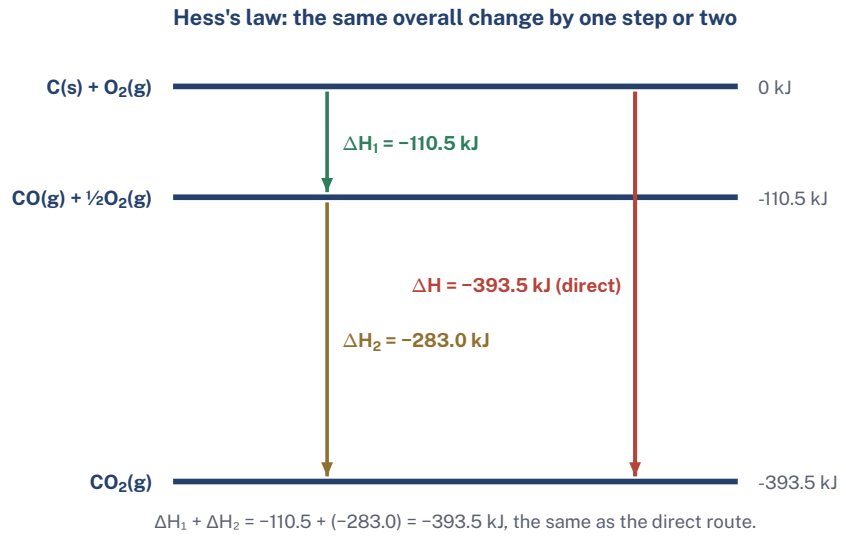


Figure 2 · The direct combustion of carbon to carbon dioxide has the same ΔH as burning it first to carbon monoxide and then burning the carbon monoxide.

Method: how to apply Hess's law

1. Write the target equation.
2. For each given equation, ask: does the substance I need appear on the correct side? If not, reverse the equation and change the sign of ΔH.
3. Multiply an equation if the target needs more moles of that substance, and multiply its ΔH by the same number.
4. Add the equations. Substances that appear on both sides cancel. Check that the sum is the target.
5. Add the ΔH values.

Worked example 1: reversing an equation

Carbon burns to carbon monoxide only partly, so ΔH cannot be measured directly. Find ΔH for $C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$ from

(a) $C(s) + O_2(g) \rightarrow CO_2(g)$, $\Delta H = -393.5$ kJ

(b) $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$, $\Delta H = -283.0$ kJ

Keep (a). CO is a reactant in (b) but a product in the target, so reverse (b): $CO_2(g) \rightarrow CO(g) + \frac{1}{2}O_2(g)$, $\Delta H = +283.0$ kJ. Add:

$$C + O_2 + CO_2 \rightarrow CO_2 + CO + \tfrac{1}{2}O_2, \text{ which cancels to } C + \tfrac{1}{2}O_2 \rightarrow CO$$

$$\Delta H = -393.5 + 283.0 = -110.5 \text{ kJ}$$

The accepted formation enthalpy of carbon monoxide is −110.53 kJ/mol (NIST Chemistry WebBook).

Worked example 2: reversing and multiplying

Find ΔH for $2C(s) + O_2(g) \rightarrow 2CO(g)$ from

(a) $C(s) + O_2(g) \rightarrow CO_2(g)$, $\Delta H = -393.5$ kJ

(c) $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$, $\Delta H = -566.0$ kJ

The target has 2 mol of C, so multiply (a) by 2: $2C + 2O_2 \rightarrow 2CO_2$, $\Delta H = -787.0$ kJ. The target has 2CO as a product, but (c) has it as a reactant, so reverse (c): $2CO_2 \rightarrow 2CO + O_2$, $\Delta H = +566.0$ kJ. Add the two:

$$2C + 2O_2 + 2CO_2 \rightarrow 2CO_2 + 2CO + O_2, \text{ which cancels to } 2C + O_2 \rightarrow 2CO$$

$$\Delta H = -787.0 + 566.0 = -221.0 \text{ kJ}$$

This is twice the value in example 1 ($2 \times -110.5 = -221.0$), as it must be for twice the amount.

Worked example 3: from enthalpies of formation

The enthalpy change of a reaction is the sum of the formation enthalpies of the products minus the sum for the reactants (an element in its standard state is 0): $\Delta H = \Sigma \Delta_f H(\text{products}) - \Sigma \Delta_f H(\text{reactants})$. For the combustion of methane at the Kerrindal Works boiler (invented), $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$, use $\Delta_f H$ of $CH_4(g) = -74.6$, $CO_2(g) = -393.51$ and $H_2O(l) = -285.83$ kJ/mol (NIST Chemistry WebBook).

$$\text{products: } -393.51 + 2(-285.83) = -965.17 \text{ kJ}$$

$$\text{reactants: } -74.6 + 2(0) = -74.6 \text{ kJ}$$

$$\Delta H = -965.17 - (-74.6) = -890.6 \text{ kJ}$$

The coefficient 2 multiplies the formation enthalpy of water. The accepted value for the enthalpy of combustion of methane is about −890 kJ/mol.

Check yourself: the water-gas shift

Use $\Delta_f H$ values in kJ/mol of $CO(g) = -110.53$, $H_2O(g) = -241.83$, $CO_2(g) = -393.51$ and $H_2(g) = 0$ to find ΔH for $CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2(g)$. Is the reaction exothermic or endothermic?

Answer. Products: $-393.51 + 0 = -393.51$. Reactants: $-110.53 + (-241.83) = -352.36$. $\Delta H = -393.51 - (-352.36) = -41.15$ kJ, so exothermic.

6. Entropy

Entropy (S) is a measure of the disorder or randomness of a system. Entropy increases when a solid melts, a liquid boils, a solid dissolves, or when a reaction produces more gas molecules than it consumes.

The entropy change of a reaction is ΔS = (entropy of products) minus (entropy of reactants). A positive ΔS means disorder increases.

SUBSTANCE	S° AT 298 K (J MOL^{-1} K^{-1})
C (graphite)	5.6
$H_2O(l)$	69.95
$H_2O(g)$	188.84
CO(g)	197.66
$CO_2(g)$	213.79

Solids have the smallest entropies and gases the largest. Standard molar entropies are from the NIST Chemistry WebBook (CODATA and Chase 1998 values), accessed 7 October 2026.

Worked example: calculating ΔS

Evaporation. $H_2O(l) \rightarrow H_2O(g)$: $\Delta S = 188.84 - 69.95 = +118.9$ J mol^{-1} K^{-1}. A gas is far more disordered than a liquid, so ΔS is large and positive.

The Kerrindal furnace (invented). $C(s) + CO_2(g) \rightarrow 2CO(g)$: $\Delta S = 2(197.66) - (5.6 + 213.79) = 395.32 - 219.39 = +175.9$ J mol^{-1} K^{-1}. One mole of gas becomes two, so disorder rises sharply.

7. Gibbs free energy and spontaneity

Two drivers decide whether a reaction is spontaneous: a tendency toward lower energy (negative ΔH) and a tendency toward greater disorder (positive ΔS). Gibbs free energy combines them:

Key idea: Gibbs free energy

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

ΔG and ΔH are in kJ mol^{-1}, ΔS is in J mol^{-1} K^{-1} (divide by 1000 for kJ), T is in kelvin (K). Negative ΔG: the reaction is spontaneous. Positive ΔG: it is non-spontaneous.

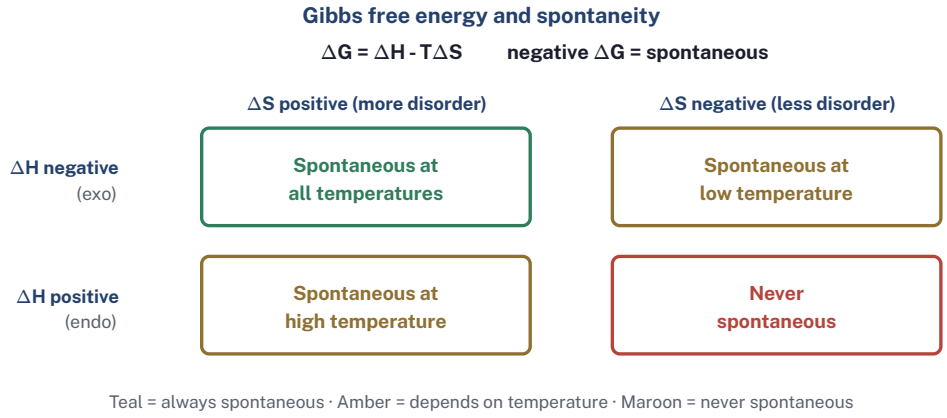


Figure 3 · The signs of ΔH and ΔS, and the temperature, together decide whether a reaction is spontaneous.

Worked example: is the reaction spontaneous?

For $N_2(g) + O_2(g) \rightarrow 2NO(g)$, $\Delta H = +180.5$ kJ mol^{-1} and $\Delta S = +24.77$ J mol^{-1} K^{-1}. Find ΔG at 298 K.

Convert ΔS to kJ: $24.77 \text{ J mol}^{-1} \text{ K}^{-1} \div 1000 = 0.02477 \text{ kJ mol}^{-1} \text{ K}^{-1}$

$$\Delta G = \Delta H - T\Delta S = 180.5 - (298 \times 0.02477)$$

$$\Delta G = 180.5 - 7.38 = +173.1 \text{ kJ mol}^{-1}$$

ΔG is positive, so the reaction is non-spontaneous at 298 K.

Worked example: the temperature at which a reaction becomes spontaneous

For C(s) + $CO_2(g)$ → 2CO(g) in the furnace: ΔH = 2(-110.53) - (-393.51) = +172.45 kJ mol^{-1} and ΔS = +175.9 J mol^{-1} K^{-1} = +0.1759 kJ mol^{-1} K^{-1} (175.9 ÷ 1000).

At 298 K: ΔG = 172.45 - 298 × 0.1759 = 172.45 - 52.4 = +120 kJ mol^{-1}, non-spontaneous.

Changeover. ΔG = 0 when T = ΔH / ΔS = 172.45 / 0.1759 = 980 K (ΔS in kJ mol^{-1} K^{-1}) (about 707 °C). Above 980 K the TΔS term is larger than ΔH, so ΔG is negative and the reaction is spontaneous. Here ΔH is positive and ΔS is positive, the "high temperature" cell of the figure.

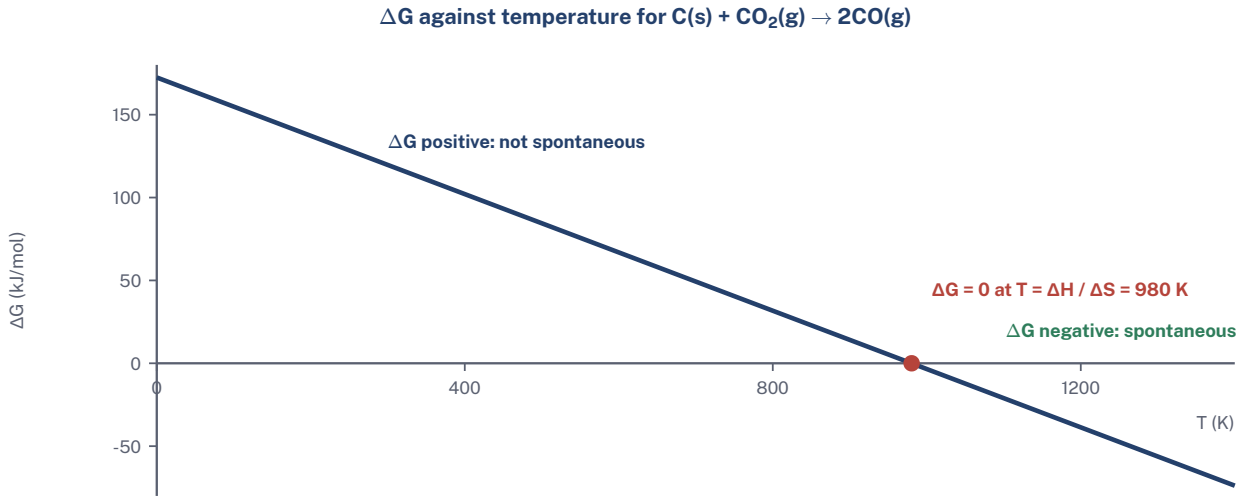


Figure 4 · ΔG for the furnace reaction falls as the temperature rises and crosses zero near 980 K.

Worked example: enthalpy favourable, entropy unfavourable, and a trap

Combustion of methane, $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$: ΔH = -890.6 kJ mol^{1}, and ΔS = 213.79 + 2(69.95) - 186.25 - 2(205.15) = -242.9 J mol^{-1} K^{-1} (NIST values). ΔS is negative, because three gas molecules become one and liquid water forms.

Converting to kJ, ΔS = -0.2429 kJ mol^{1} K^{-1}. ΔG at 298 K = -890.6 - 298 × (-0.2429) = -890.6 + 72.4 = -818 kJ mol^{1}, strongly spontaneous. ΔH is negative and ΔS is negative, so the reaction is spontaneous at low temperature, and only a very high temperature, ΔH / ΔS = 890.6 / 0.2429 = 3667 K, would reverse it.

Trap. Spontaneous does not mean fast: methane and oxygen can sit mixed for years until a spark supplies the activation energy. ΔG says whether a reaction can go; the rate is a separate question.

Exam focus: Module 4

State the sign of ΔH: negative for exothermic, positive for endothermic, and say what it means.

In calorimetry, use q = mcΔT for the water, then divide by moles for the molar enthalpy, and explain heat loss.

For Hess's law, reverse and multiply the given equations until they add to the target, and change ΔH to match each step.

For spontaneity, use ΔG = ΔH - TΔS, convert ΔS from J mol^{-1} K^{-1} to kJ mol^{-1} K^{-1} by dividing by 1000, and judge the sign of ΔG.

Check your understanding

1. Is a reaction with ΔH =−286 kJ/mol exothermic or endothermic?
2. Calculate the heat gained by 150 g of water that warms by 12.0 degrees C.
3. Using H-H = 436 and Cl-Cl = 242 and H-Cl = 431, find ΔH for $H_2 + Cl_2$ gives 2HCl.
4. Predict the sign of ΔS for a reaction that produces a gas from solids.
5. A reaction has ΔH negative and ΔS positive. Is it spontaneous? Explain.
6. Given $C(s) + O_2(g) \rightarrow CO_2(g)$, ΔH = -393.5 kJ, and $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$, ΔH = -285.8 kJ, and $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$, ΔH = -1411 kJ, find ΔH for $2C(s) + 2H_2(g) \rightarrow C_2H_4(g)$.
7. A reaction has ΔH = +60.0 kJ mol^{-1} and ΔS = +200 J mol^{-1} K^{-1}. Find ΔG at 298 K and the temperature above which it is spontaneous.

Answers to checks 6 and 7

6. 2 × (-393.5) + 2 × (-285.8) = -787.0 - 571.6 = -1358.6 kJ; reverse the ethene equation (+1411 kJ): ΔH = -1358.6 + 1411 = +52.4 kJ, endothermic. **7.** ΔS = 200 J mol^{-1} K^{-1} = 0.200 kJ mol^{-1} K^{-1}. ΔG = 60.0 - 298 × 0.200 = 60.0 - 59.6 = +0.4 kJ mol^{-1} at 298 K, which is just non-spontaneous. T = ΔH / ΔS = 60.0 / 0.200 = 300 K, so it is spontaneous above 300 K. Answers 1 to 5 follow the worked examples above: exothermic; 7524 J (7.52 kJ); −184 kJ/mol; positive; yes, negative ΔG at every temperature.

About this resource

MA99 EDUCATION

Scope

NSW Chemistry, Year 11. Written to the NSW Chemistry Stage 6 Syllabus (2017), Module 4, for students sitting the HSC in 2028. A new Chemistry syllabus is first examined in HSC 2029. Every number in it was recomputed from its inputs before printing. Thermodynamic data are from the NIST Chemistry WebBook (CODATA and Chase 1998 values), accessed 7 October 2026; contexts at the Kerrindal Works are invented.

Permitted use

Print it, annotate it and keep it on any device you own, for your own study.

Not permitted

Sharing, posting, selling or redistributing it, in whole or in part. Each copy is licensed to one student.

Copyright MA99 Education.
ma99education.com.au