

Energetics (answers)

IB CHEMISTRY HL

Michael Sugiyama Jones
WWW.MSJCHEM.COM

15.1 Energy cycles

Understandings:

- Representative equations (eg $M^+_{(g)} \rightarrow M^+_{(aq)}$) can be used for enthalpy/energy of hydration, ionization, atomization, electron affinity, lattice, covalent bond and solution.
- Enthalpy of solution, hydration enthalpy and lattice enthalpy are related in an energy cycle.

Applications and skills:

- Construction of Born-Haber cycles for group 1 and 2 oxides and chlorides.
- Construction of energy cycles from hydration, lattice and solution enthalpy. For example dissolution of solid NaOH or NH_4Cl in water.
- Calculation of enthalpy changes from Born-Haber or dissolution energy cycles.
- Relate size and charge of ions to lattice and hydration enthalpies.
- Perform lab experiments which could include single replacement reactions in aqueous solutions.

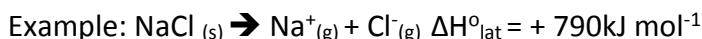
Guidance:

- Polarizing effect of some ions producing covalent character in some largely ionic substances will not be assessed.
- The following enthalpy/energy terms should be covered: ionization, atomization, electron affinity, lattice, covalent bond, hydration and solution.
- Value for lattice enthalpies (section 18), enthalpies of aqueous solutions (section 19) and enthalpies of hydration (section 20).

Lattice enthalpy ($\Delta H^\circ_{\text{lat}}$)

Lattice enthalpy can be defined in two ways:

- 1) The enthalpy change when one mole of solid ionic compound is separated into its gaseous ions under standard conditions (endothermic).



- 2) The enthalpy change when one mole of solid ionic compound is formed from its gaseous ions under standard conditions (exothermic).

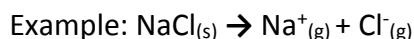


- The IB uses the **first** definition.

$\Delta H^\circ_{\text{lat}}$ is endothermic and has a positive ΔH .

Some important definitions:

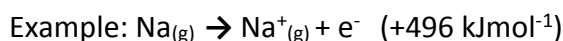
- **Lattice enthalpy ($\Delta H^\circ_{\text{lat}}$)** - the enthalpy change when one mole of a solid ionic compound is separated into gaseous ions under standard conditions (endothermic).



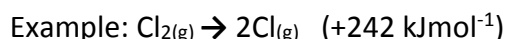
- **Enthalpy of atomisation ($\Delta H^\circ_{\text{atom}}$)** – the enthalpy change when one mole of gaseous atoms is formed from an element in its standard state (endothermic).



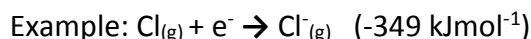
- **First ionisation energy ($\Delta H^\circ_{\text{i}}$)** – the minimum energy required to remove one mole of electrons from one mole of gaseous atoms (endothermic)



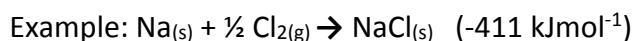
- **Bond dissociation enthalpy (E)** – the enthalpy change when one mole of bonds is broken in the gaseous state (endothermic).



- **First electron affinity ($\Delta H^\circ_{\text{e}}$)** – the enthalpy change when one mole of electrons are added to one mole of gaseous atoms (exothermic).

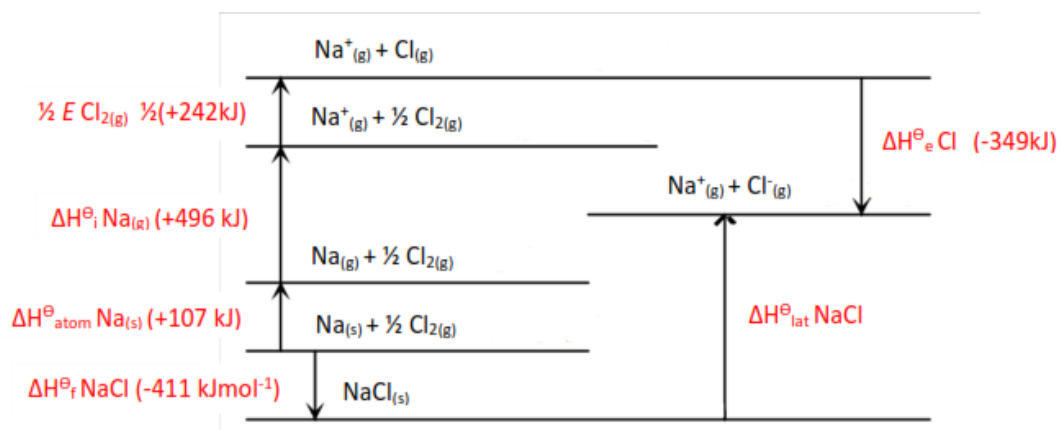


- **Standard enthalpy change of formation ($\Delta H^\circ_{\text{f}}$)** - The enthalpy change when one mole of a compound is formed from its elements in their standard states under standard conditions.



Exercises:

- Using the values given above, construct a Born-Haber cycle using the template below, and calculate the $\Delta H^\ominus_{\text{lat}}$ for NaCl.

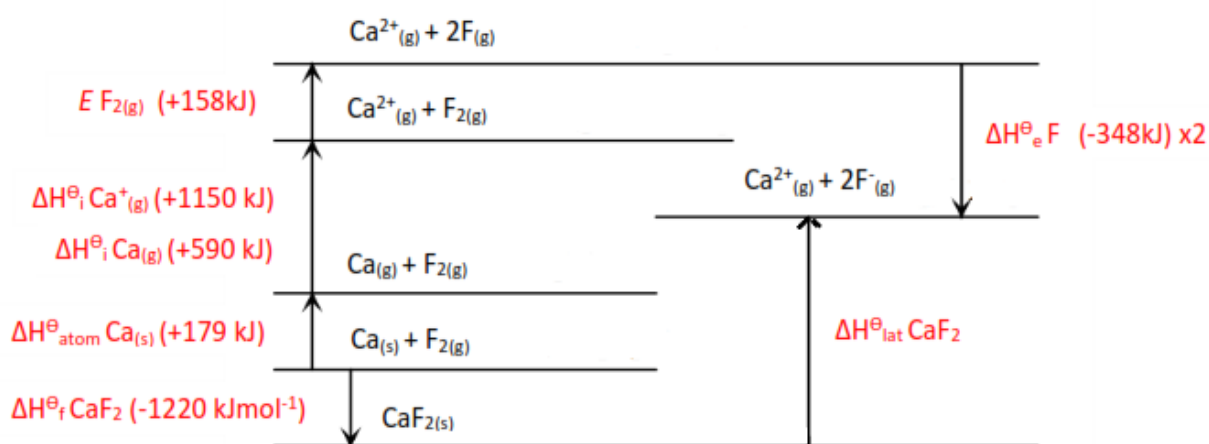


$$\Delta H^\ominus_{\text{lat}} = (\Delta H^\ominus_{\text{atom Na(s)}}) + (\Delta H^\ominus_{\text{i Na(g)}}) + (\frac{1}{2} E \text{Cl}_{2(g)}) + (\Delta H^\ominus_{\text{e Cl}}) - (\Delta H^\ominus_{\text{f NaCl}})$$

$$\Delta H^\ominus_{\text{lat}} = (+107) + (+496) + (+121) + (-349) - (-411)$$

$$\Delta H^\ominus_{\text{lat}} = +786 \text{ kJmol}^{-1}$$

- Using the values given in the table below, construct a Born-Haber cycle and calculate the $\Delta H^\ominus_{\text{lat}}$ for CaF₂



$$\Delta H^\ominus_{\text{lat}} = (\Delta H^\ominus_{\text{atom Ca(s)}}) + (\Delta H^\ominus_{\text{i Ca(g)}}) + (\Delta H^\ominus_{\text{i Ca}^+(g)}) + (E \text{F}_{2(g)}) + 2(\Delta H^\ominus_{\text{e F}_2}) - (\Delta H^\ominus_{\text{f CaF}_2})$$

$$\Delta H^\ominus_{\text{lat}} = (+179) + (+590) + (+1150) + (+158) + 2(-348) - (-1220)$$

$$\Delta H^\ominus_{\text{lat}} = +2601 \text{ kJmol}^{-1}$$

Factors that affect lattice enthalpy

- The charge on the ion – an ion with a higher charge will have a stronger force of attraction, therefore a higher lattice enthalpy.

Example – MgO has a higher lattice enthalpy than NaCl.

- The ionic radius of the ions – an increase in ionic radius decreases the attraction between the ions.

Example - NaCl has a higher lattice enthalpy than KBr.

Exercises:

1. Which combination of ionic radius and ionic charge would result in the highest lattice enthalpy for an ionic compound? **A**

	Ionic radius	Ionic charge
A.	small	high
B.	large	high
C.	small	low
D.	large	low

2. Which is a correct definition of lattice enthalpy? **D**

- A. It is the enthalpy change that occurs when an electron is removed from 1 mol of gaseous atoms.
- B. It is the enthalpy change that occurs when 1 mol of a compound is formed from its elements.
- C. It is the enthalpy change that occurs when 1 mol of solid crystal changes into a liquid.
- D. It is the enthalpy change that occurs when 1 mol of solid crystal is broken into its gaseous ions.

3. Which ionic compound has the greatest lattice enthalpy? **A**

- A. MgO
- B. CaO
- C. NaF
- D. KF

4. Which is the correct order of increasing magnitude of lattice enthalpy (lowest first)? **C**

A. $\text{NaCl} < \text{KCl} < \text{MgS} < \text{MgO}$

B. $\text{MgO} < \text{MgS} < \text{KCl} < \text{NaCl}$

C. $\text{KCl} < \text{NaCl} < \text{MgS} < \text{MgO}$

D. $\text{MgO} < \text{NaCl} < \text{KCl} < \text{MgS}$

5. Which ionic compound has the most endothermic lattice enthalpy? **C**

A. NaCl

B. KCl

C. NaF

D. KF

Enthalpies of solution and hydration

Enthalpy of solution $\Delta H^\ominus_{\text{sol}}$

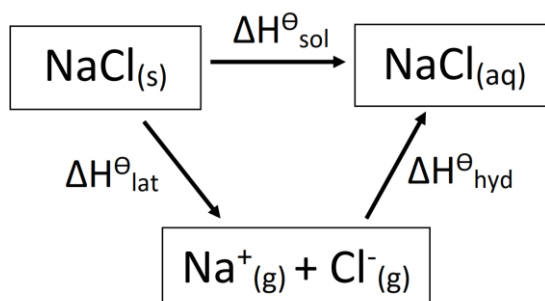
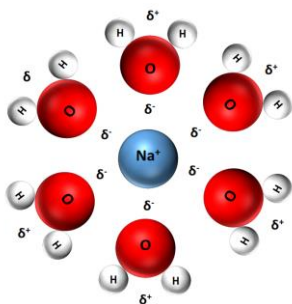
- The enthalpy change of solution ($\Delta H^\ominus_{\text{sol}}$) is the enthalpy change when one mole of an ionic substance dissolves in water to give a solution of infinite dilution.
- A solution of infinite dilution is one where there is a sufficiently large excess of water that adding any more does not cause any further heat to be absorbed or evolved.

Enthalpy of hydration $\Delta H^\ominus_{\text{hyd}}$

- The enthalpy of hydration (ΔH_{hyd}) is the enthalpy change when one mole of gaseous ions dissolve in water to give a solution of infinite dilution.

Dissolving

- The solid ionic compound is broken down into its gaseous ions ($\Delta H^\ominus_{\text{lat}}$).
- The gaseous ions are hydrated by water molecules ($\Delta H^\ominus_{\text{hyd}}$).



$$\Delta H^\ominus_{\text{sol}} = \Delta H^\ominus_{\text{lat}} + \Delta H^\ominus_{\text{hyd}}$$

Example: Calculate the enthalpy change of solution for sodium chloride.

$$\Delta H_{\text{sol}}^{\ominus} = \Delta H_{\text{lat}}^{\ominus} + \Delta H_{\text{hyd}}^{\ominus}$$

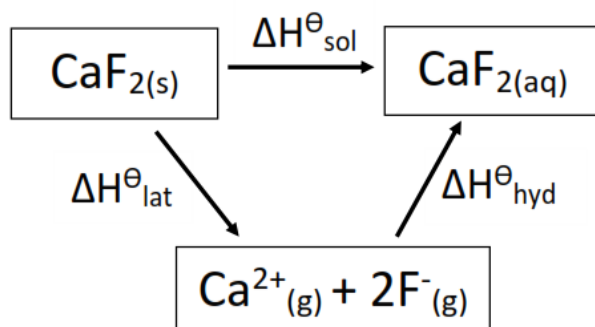
$$\Delta H_{\text{sol}}^{\ominus} (\text{NaCl}) = (+790) + (-424 + -359)$$

$$\Delta H_{\text{sol}}^{\ominus} (\text{NaCl}) = +7 \text{ kJmol}^{-1}$$

Exercises:

1. Construct an enthalpy cycle and calculate the enthalpy change of solution ($\Delta H_{\text{sol}}^{\ominus}$) for calcium fluoride (CaF_2).

1) Enthalpy cycle for CaF_2



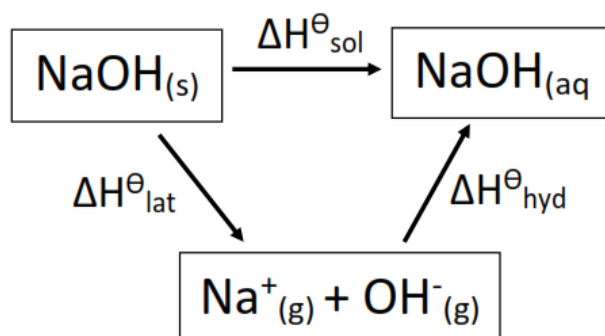
$$\Delta H_{\text{sol}}^{\ominus} = \Delta H_{\text{lat}}^{\ominus} + \Delta H_{\text{hyd}}^{\ominus}$$

$$\Delta H_{\text{sol}}^{\ominus} = +2651 + (-1616 + 2 \times -504)$$

$$\Delta H_{\text{sol}}^{\ominus} = +27 \text{ kJmol}^{-1}$$

2. Construct an enthalpy cycle and calculate the lattice enthalpy ($\Delta H_{\text{lat}}^{\ominus}$) for NaOH.

2) Enthalpy cycle for NaOH



$$\Delta H_{\text{lat}}^{\ominus} = \Delta H_{\text{sol}}^{\ominus} - \Delta H_{\text{hyd}}^{\ominus}$$

$$\Delta H_{\text{lat}}^{\ominus} (\text{NaOH}) = -44.5 - (-424 + -519)$$

$$\Delta H_{\text{lat}}^{\ominus} (\text{NaOH}) = +898 \text{ kJmol}^{-1}$$

Factors that affect $\Delta H^\ominus_{\text{hyd}}$

- The size of the hydration enthalpy is determined by the attraction between the ions and the water molecules.
- The size of the ion – the smaller the ion, the stronger the attraction and the higher the $\Delta H^\ominus_{\text{hyd}}$
- The charge on the ion – the greater the charge on the ion, the stronger the attraction and the higher the $\Delta H^\ominus_{\text{hyd}}$

Exercises:

1) Explain why the enthalpy change of hydration ($\Delta H^\ominus_{\text{hyd}}$) of Na^+ is greater than K^+

Na^+ has a smaller ionic radius than K^+ , therefore a greater $\Delta H^\ominus_{\text{hyd}}$ (stronger electrostatic attraction between the Na^+ ion and water molecule).

2) Explain why the enthalpy change of hydration ($\Delta H^\ominus_{\text{hyd}}$) of Mg^{2+} is greater than Li^+

Mg^{2+} has a higher charge on the ion, therefore a greater $\Delta H^\ominus_{\text{hyd}}$ (same reason as above).

15.2 Entropy and spontaneity

Understandings:

- Entropy (S) refers to the distribution of available energy among the particles. The more ways the energy can be distributed the higher the entropy.
- Gibbs free energy (G) relates the energy that can be obtained from a chemical reaction to the change in enthalpy (ΔH), change in entropy (ΔS), and absolute temperature (T).
- Entropy of gas > liquid > solid under same conditions.

Applications and skills:

- Prediction of whether a change will result in an increase or decrease in entropy by considering the states of the reactants and products.
- Calculation of entropy changes (ΔS) from given standard entropy values (S^\ominus).
- Application of $\Delta G = \Delta H - T\Delta S$
- Relation of ΔG to position of equilibrium.

Guidance:

- Examine various reaction conditions that affect ΔG
- ΔG is a convenient way to take into account both the direct entropy change resulting from the transformation of the chemicals, and the indirect entropy change of the surroundings as a result of the gain/loss of heat energy.

Spontaneity

- A spontaneous process is a process that occurs without adding energy (other than the energy required to overcome the energy barrier).

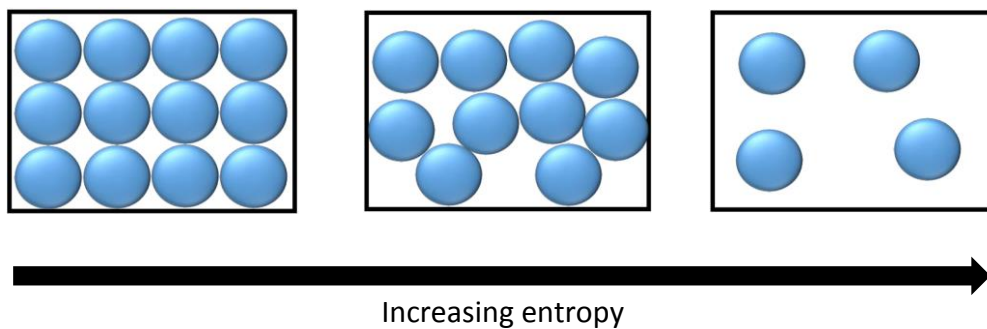


- Spontaneous processes can be either fast or slow.



Entropy

- Entropy (S) refers to the distribution of available energy among the particles in a system.
- The more ways the energy can be distributed, the greater the entropy.
- Gases have a higher entropy than liquids which have a higher entropy than solids.



Exercises:

1. What is the correct order of decreasing entropy for a pure substance?

- A. gas > liquid > solid
- B. solid > liquid > gas
- C. solid > gas > liquid
- D. liquid > solid > gas

A – gases have the highest entropy, followed by liquids and then solids which have the lowest entropy.

2. Which change will **not** increase the entropy of a system?

- A. Increasing the temperature
- B. Changing the state from liquid to gas
- C. Mixing different types of particles
- D. A reaction where four moles of gaseous reactants changes to two moles of gaseous products

D – there is a decrease in the number of moles of gas in the products, therefore a decrease in entropy.

3. Which reaction has the greatest increase in entropy?

- A. $\text{SO}_2(\text{g}) + 2\text{H}_2\text{S}(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + 3\text{S}(\text{s})$
- B. $\text{CaO}(\text{s}) + \text{CO}_2(\text{g}) \rightarrow \text{CaCO}_3(\text{s})$
- C. $\text{CaC}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{OH})_2(\text{s}) + \text{C}_2\text{H}_2(\text{g})$
- D. $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$

C – in this reaction, one mole of solid and one mole of liquid react to form one mole of solid and one mole of gas. Gases have a higher entropy than solids therefore entropy increases.

4. Which reaction has the most negative change in entropy?

- A. $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g})$
- B. $\text{NH}_4\text{Cl}(\text{s}) \rightarrow \text{NH}_3(\text{g}) + \text{HCl}(\text{g})$
- C. $\text{PbCl}_2(\text{s}) \rightarrow \text{Pb}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq})$
- D. $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$

A – there is a decrease in the number of moles of in the products, therefore the entropy change will be the most negative.

5. Which reaction has the largest increase in entropy?

- A. $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g})$
- B. $\text{Al}(\text{OH})_3(\text{s}) + \text{NaOH}(\text{aq}) \rightarrow \text{Al}(\text{OH})_4^{-}(\text{aq}) + \text{Na}^{+}(\text{aq})$
- C. $\text{Na}_2\text{CO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
- D. $\text{BaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{NaCl}(\text{aq})$

C – solid and aqueous reactants become gases and liquids. Gases and liquids have higher entropy than solids, therefore entropy increases.

6. Which reaction has the greatest positive entropy change?

- A. $\text{CH}_4(\text{g}) + 1\frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CO}(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
- B. $\text{CH}_4(\text{g}) + 1\frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CO}(\text{g}) + 2\text{H}_2\text{O}(\text{l})$
- C. $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
- D. $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

A – 2 ½ moles of gas in the reactants form 3 moles of gas in the products which is an increase in entropy.

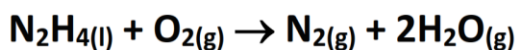
Standard entropy change (ΔS^\ominus)

- The standard entropy change of reaction can be calculated using this equation:

$$\Delta S^\ominus_{\text{reaction}} = \sum S^\ominus_{\text{products}} - \sum S^\ominus_{\text{reactants}}$$

- A perfectly ordered crystal at zero kelvin has an entropy of zero.
- Absolute entropy values are positive.

Example: Calculate the entropy change for this reaction:



	$S^\ominus (\text{J K}^{-1} \text{mol}^{-1})$
$\text{N}_2\text{H}_4(\text{l})$	121
$\text{O}_2(\text{g})$	205
$\text{N}_2(\text{g})$	192
$\text{H}_2\text{O}(\text{g})$	188

$$\Delta S^\ominus = (192 + 2 \times 188) - (205 + 121) = 242 \text{ J K}^{-1} \text{mol}^{-1}$$

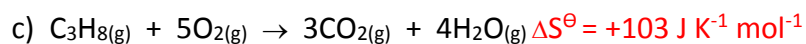
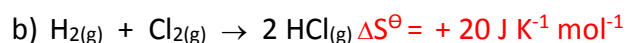
Exercise: Use the data in the table to answer the questions.

Substance	$S^\ominus (\text{J K}^{-1}\text{mol}^{-1})$	Substance	$S^\ominus (\text{J K}^{-1}\text{mol}^{-1})$
$\text{H}_2\text{O}(\text{g})$	189	$\text{CO}_2(\text{g})$	214
$\text{H}_2\text{O}(\text{l})$	70	$\text{C}_3\text{H}_8(\text{g})$	270
$\text{H}_2(\text{g})$	131	$\text{O}_2(\text{g})$	205
$\text{Cl}_2(\text{g})$	223		
$\text{HCl}(\text{g})$	187		

1) Predict the sign of ΔS^\ominus for the following reactions and explain your reasoning.

- a) $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$ **negative ΔS^\ominus - gases have higher entropy than liquids.**
- b) $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g})$ **small change in entropy – same number of moles of gas in reactants and products.**
- c) $\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$ **positive ΔS^\ominus - six moles of gas in reactants, seven moles of gas in products.**

2) Calculate ΔS^\ominus for the above reactions.



Gibbs free energy

- The energy associated with a chemical reaction that can be used to do work.
- ΔG must be negative for a spontaneous process.

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

ΔG – change in Gibbs free energy

ΔH – change in enthalpy

T – temperature in kelvin (K)

ΔS – change in entropy

Example: Calculate ΔG for this reaction at 500 K, using the data in the table. State and explain if the reaction is spontaneous at that temperature.



Compound	$\Delta H_f^\ominus (\text{kJ mol}^{-1})$	$\Delta S^\ominus (\text{J K}^{-1} \text{ mol}^{-1})$
$\text{CaCO}_{3(\text{s})}$	-1207	92.9
$\text{CaO}_{(\text{s})}$	-635	39.7
$\text{CO}_{2(\text{g})}$	-394	214

See video for solution.

Calculate the temperature at which the reaction will be spontaneous:

See video for solution.

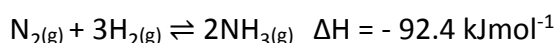
Exercises:

1. A reaction has a standard enthalpy change, ΔH , of +10.00 kJ mol⁻¹ at 298 K. The standard entropy change, ΔS , for the same reaction is +10.00 J K⁻¹ mol⁻¹. What is the value of ΔG for the reaction in kJ mol⁻¹?

$$\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus = 10000 - (298 \times 10.00) = 7020 \text{ kJmol}^{-1}$$

$$\Delta G^\ominus = +7.02 \text{ kJmol}^{-1}$$

2. Consider the following reaction:



- (i) The absolute entropy values, S , at 238 K for $\text{N}_{2(\text{g})}$, $\text{H}_{2(\text{g})}$ and $\text{NH}_{3(\text{g})}$ are 192, 131 and 193 J K⁻¹ mol⁻¹ respectively. Calculate ΔS^\ominus for the reaction and explain the sign of ΔS^\ominus .

$$\Delta S^\ominus = [2 \times 193] - [192 + 3 \times 131] = -199 \text{ J K}^{-1} \text{ mol}^{-1}$$

ΔS^\ominus is negative - four gaseous molecules forming two gaseous molecules/fewer molecules of gas

- (ii) Calculate ΔG^\ominus for the reaction at 298 K. State and explain whether the reaction is spontaneous.

$$\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus = -76.0 - 298(-0.199) = -16.7 \text{ kJmol}^{-1}$$

Reaction is spontaneous (ΔG is negative)

3. Hex-1-ene gas, C_6H_{12} , burns in oxygen to produce carbon dioxide and water vapour.

- (a) Write an equation to represent this reaction.



- (b) Use the data below to calculate the values of ΔH_c^\ominus and ΔS^\ominus for the combustion of hex-1-ene.

Substance	$\text{O}_{2(\text{g})}$	$\text{C}_6\text{H}_{12(\text{g})}$	$\text{CO}_{2(\text{g})}$	$\text{H}_2\text{O}_{(\text{g})}$
Standard enthalpy of formation, ΔH_f^\ominus (kJmol ⁻¹)	0.0	-43	-394	-242
Entropy, S^\ominus (J K ⁻¹ mol ⁻¹)	205	385	214	189

(i) Value of ΔH_c^\ominus

$$\Delta H_c^\ominus = \sum \Delta H_f^\ominus \text{ products} - \sum \Delta H_f^\ominus \text{ reactants}$$

$$\Delta H_c^\ominus = (6 \times -394 + 6 \times -242) - (-43)$$

$$\Delta H_c^\ominus = -3773 \text{ kJ mol}^{-1} / -3.8 \times 10^3 \text{ kJ mol}^{-1} \text{ (2 SF)}$$

(ii) Value of ΔS^\ominus

$$\Delta S^\ominus = \sum \Delta S^\ominus \text{ products} - \sum \Delta S^\ominus \text{ reactants}$$

$$\Delta S^\ominus = (6 \times 189 + 6 \times 214) - (385 + 9 \times 205)$$

$$\Delta S^\ominus = +188 \text{ J K}^{-1} \text{ mol}^{-1}$$

(c) Calculate the standard free energy change for the combustion of hex-1-ene.

$$\Delta G^\ominus = \Delta H - T\Delta S = -3773 - (298 \times 0.188)$$

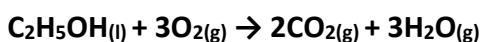
$$\Delta G^\ominus = -3829 \text{ kJ mol}^{-1} / -3830 \text{ kJ mol}^{-1} \text{ (3 SF)}$$

(d) State and explain whether or not the combustion of hex-1-ene is spontaneous at 25°C. **Reaction is spontaneous (ΔG^\ominus negative)**

Standard Gibbs free energy (ΔG^\ominus)

$$\Delta G^\ominus = \sum \Delta G_f^\ominus (\text{products}) - \sum \Delta G_f^\ominus (\text{reactants})$$

Example: Calculate the ΔG^\ominus for the following reaction. State and explain if the reaction is spontaneous at 298 K.



Compound	$\Delta G_f^\ominus / \text{kJ mol}^{-1}$
$\text{C}_2\text{H}_5\text{OH}(\text{l})$	-175
$\text{CO}_{2(\text{g})}$	-394
$\text{H}_2\text{O}(\text{g})$	-229
$\text{O}_{2(\text{g})}$	0

$$\Delta G^\ominus = (3 \times -229 + 2 \times -394) - (-175)$$

$$\Delta G^\ominus = -1300 \text{ kJ mol}^{-1}$$

The effect of ΔH , ΔS and T on the spontaneity of a chemical reaction

Exothermic reactions

ΔH	ΔS	T	ΔG	spontaneity
negative < 0	positive > 0	any temperature		

- ΔH is negative and ΔS is positive, therefore ΔG will be negative at any temperature.
- An exothermic reaction with an increase in entropy will be spontaneous at any temperature.

ΔH	ΔS	T	ΔG	spontaneity
negative < 0	negative < 0	high		

ΔH	ΔS	T	ΔG	spontaneity
negative < 0	negative < 0	low		

- ΔG will be negative only at low temperatures; at high T , ($T\Delta S$) makes ΔG positive.

Endothermic reactions

ΔH	ΔS	T	ΔG	spontaneity
positive > 0	negative < 0	any temperature		

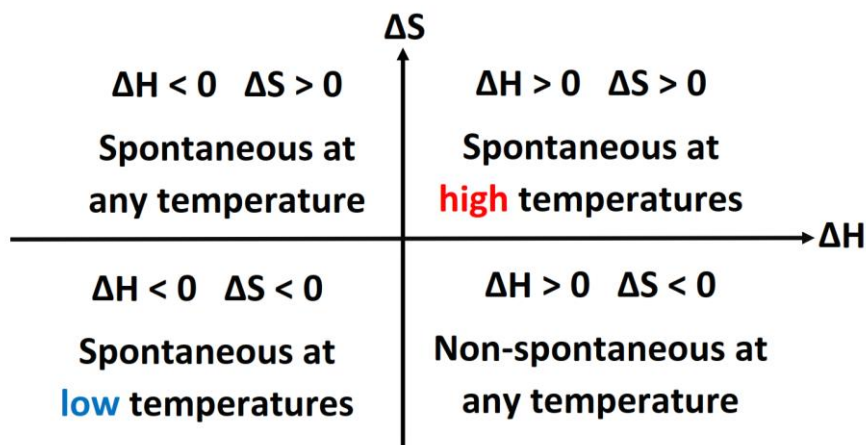
- ΔH is positive and ΔS is negative, therefore ΔG will be positive at any temperature.
- An endothermic reaction with a decrease in entropy will be non-spontaneous at any temperature.

ΔH	ΔS	T	ΔG	spontaneity
positive > 0	positive > 0	high		

ΔH	ΔS	T	ΔG	spontaneity
positive > 0	positive > 0	low		

- ΔG will be negative only at high temperatures ($T\Delta S > \Delta H$); at low temperatures $\Delta H > T\Delta S$ ($+\Delta G$).

Summary



Exercises:

- Some words used in chemistry can have a specific meaning which is different to their meaning in everyday English. State what the term *spontaneous* means when used in a chemistry context.

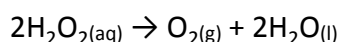
A reaction that occurs without adding energy beyond that required to overcome the energy barrier.

- Under what circumstances is a reaction spontaneous at all temperatures?

	ΔH^\ominus	ΔS^\ominus
A.	+	+
B.	+	–
C.	–	–
D.	–	+

D

- When hydrogen peroxide decomposes, the temperature of the reaction mixture increases.



What are the signs of ΔH , ΔS and ΔG for this reaction? **B**

	ΔH	ΔS	ΔG
A.	–	–	–
B.	–	+	–
C.	+	+	–
D.	–	+	+

4. ΔG^\ominus calculations predict that a reaction is always spontaneous for which of the following combinations of ΔH^\ominus and ΔS^\ominus ?

- A. $+\Delta H^\ominus$ and $+\Delta S^\ominus$
- B. $+\Delta H^\ominus$ and $-\Delta S^\ominus$
- C. $-\Delta H^\ominus$ and $-\Delta S^\ominus$
- D. $-\Delta H^\ominus$ and $+\Delta S^\ominus$

D

5. What are the signs of ΔH^\ominus and ΔS^\ominus for a reaction that is non-spontaneous at low temperature but spontaneous at high temperature?

	ΔH^\ominus	ΔS^\ominus
A.	–	–
B.	+	–
C.	–	+
D.	+	+

D