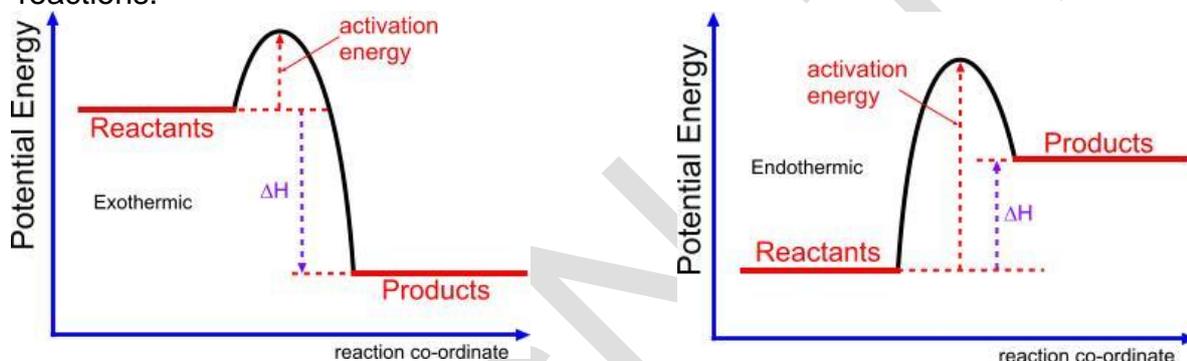


- Understand the difference between heat and temperature.
Heat - a form of energy that flows from something at a higher temperature to something at a lower temperature.
Temperature - a measure of the average kinetic energy of particles.
- Understand the difference between exothermic and endothermic reactions in terms of temperature changes and transfer of heat.
Exothermic – heat energy is transferred from a system (chemical reaction) to the surroundings – the surroundings get hotter.
Endothermic - a system (chemical reaction) takes in heat energy from the surroundings – the surroundings get cooler.
- State the signs of ΔH for exothermic and endothermic reactions
Negative for exothermic and positive for endothermic
- State whether the products are more or less stable than the reactants in an exothermic reaction. Products more stable
- Draw potential energy profiles/energy level diagrams for exothermic and endothermic reactions.

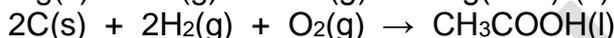
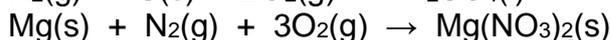
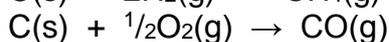
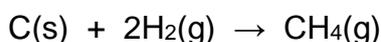


- State whether combustion is an exothermic or endothermic process. Exothermic
- State whether the neutralisation reaction when an acid reacts with an alkali is an exothermic or endothermic process. Exothermic
- Calculate the enthalpy change of combustion of hexane given that when 1.20 g of hexane (C_6H_{14}) is burnt the temperature of 250.0 g of water (specific heat capacity $4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}$) is raised by $56.0 \text{ }^\circ\text{C}$. $-4190 \text{ kJ mol}^{-1}$
- Give two reasons why the enthalpy change in 8 is less exothermic than the literature value.
Incomplete combustion/heat loss to the surroundings
- Describe an experimental procedure which could be used to measure the enthalpy change of a reaction occurring in aqueous solution.

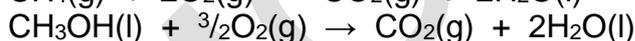
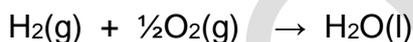
The reagents are measured out and their temperatures measured. They are then mixed together in a polystyrene cup (calorimeter) and the maximum/minimum temperature recorded. The specific heat capacity and density of the final solution is assumed to be the same as water.
- Calculate the enthalpy change of solution (for 1 mol of substance dissolving) given that, when 0.950 g of magnesium chloride is dissolved in 25.0 cm^3 of water the temperature rises by $14 \text{ }^\circ\text{C}$. Explain 2 assumptions that have been made in the calculation.

-146 kJ mol^{-1} Specific heat capacity same as water Density same as water

- 12 Explain Hess's Law.
The enthalpy change accompanying a chemical reaction is independent of the pathway between the initial and final states.
- 13 Define the standard enthalpy change of formation for a substance
The enthalpy change when one mole of the substance is formed from its elements in their standard states under standard conditions.
- 14 Explain what is meant by *standard conditions*
All values for enthalpy changes are quoted for standard conditions.
Pressure: 100 kPa
Temperature is not included in the definition of standard conditions and should always be quoted – if no temperature is given assume 298 K (25 °C)
- 15 Explain what is meant by *standard state*
The pure substance at 100 kPa and a specified temperature (usually 298 K).
- 16 Write equations which represents the standard enthalpy change of formation of
CH₄(g) CO(g) H₂SO₄(l) Mg(NO₃)₂(s) CH₃COOH(l)



- 17 Explain why the standard enthalpy change of formation of O₂(g) is zero.
It is an element in its standard state therefore zero by definition
- 18 Write an equation which represents the standard enthalpy change of combustion of
H₂(g) CH₄(g) CH₃OH(l)



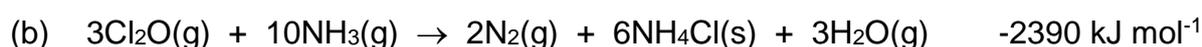
- 19 Calculate the standard enthalpy change of combustion of methanol using standard enthalpies of formation, ΔH_f^\ominus

-740 kJ mol⁻¹

Compound	CH ₃ OH(l)	CO ₂ (g)	H ₂ O(l)
$\Delta H_f^\ominus / \text{kJ mol}^{-1}$	-240	-400	-290

- 20 Calculate the enthalpy changes for the following reactions given enthalpy change of formation data

	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$		$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
SO ₂ (g)	-297	POCl ₃ (l)	-597
PCl ₅ (s)	-444	NH ₃ (g)	-46
SOCl ₂ (l)	-246	NH ₄ Cl(s)	-314
Cl ₂ O(g)	80	H ₂ O(g)	-242



21 Calculate the enthalpy change for a reaction given known enthalpy changes

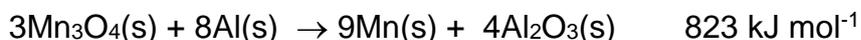
- (a) Given the values in the table, calculate the enthalpy change for:

$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$	$\Delta H = -92 \text{ kJ mol}^{-1}$
$2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$	$\Delta H = -484 \text{ kJ mol}^{-1}$



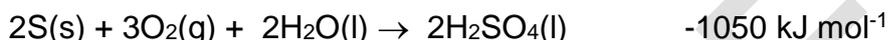
- (b) Given the values in the table, calculate the enthalpy change for:

$3\text{Mn}(\text{s}) + 2\text{O}_2(\text{g}) \rightarrow \text{Mn}_3\text{O}_4(\text{s})$	$\Delta H = -1387 \text{ kJ mol}^{-1}$
$4\text{Al}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{Al}_2\text{O}_3(\text{s})$	$\Delta H = -1669 \text{ kJ mol}^{-1}$



- (c) Given the values in the table, what is the enthalpy change for:

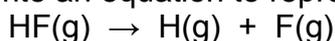
$\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$	$\Delta H = -297 \text{ kJ mol}^{-1}$
$2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g})$	$\Delta H = -196 \text{ kJ mol}^{-1}$
$\text{SO}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_4(\text{l})$	$\Delta H = -130 \text{ kJ mol}^{-1}$



22 Define average bond enthalpy

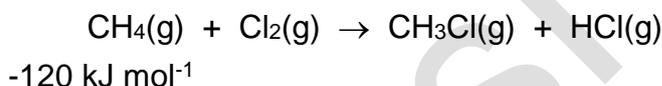
The enthalpy change when one mole of covalent bonds, in a gaseous molecule, are broken under standard conditions. The value is averaged for the same bond in a range of similar compounds.

23 Write an equation to represent the bond enthalpy (energy) of hydrogen fluoride.



24 State whether bond breaking is an exothermic or endothermic process. Endothermic

25 Use the bond energies given below to calculate the enthalpy change for the reaction:



BOND	Bond Energy / kJ mol ⁻¹
H-Cl	430
C-Cl	340
C-H	410
Cl-Cl	240

26 Explain in terms of bond enthalpies why the reaction in 25 is exothermic.

More energy is released when bonds are broken than is required to break them. On average the bonds made are stronger than the bonds broken.

27 Explain why bond enthalpies cannot be used to calculate the enthalpy change for the reaction:



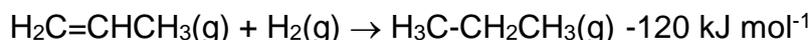
Bond energies can only be used for gaseous molecules and Br₂ is a liquid.

Intermolecular forces must also be taken into account.

28 Explain why calculations involving bond enthalpies may not be as accurate as those involving enthalpy changes of formation.

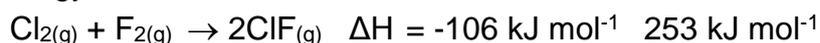
Bond energies are average values but enthalpy changes of formation are specific to a particular compound.

29 Use bond enthalpies to work out the enthalpy change for the reaction:



BOND	Bond Energy / kJ mol ⁻¹
C-H	410
C-C	350
H-H	440
C=C	610

30 Using the enthalpy change given and the appropriate bond energies work out the Cl-F bond energy in ClF.



BOND	Bond Energy / kJ mol ⁻¹
Cl-Cl	242
F-F	158

31 Explain in terms of bond strengths whether O_2 or O_3 absorbs longer wavelength UV radiation. O_2 bond order is 2, O_3 bond order is 1.5
Weaker bonds in O_3 therefore longer wavelength (lower energy) radiation required to dissociate it

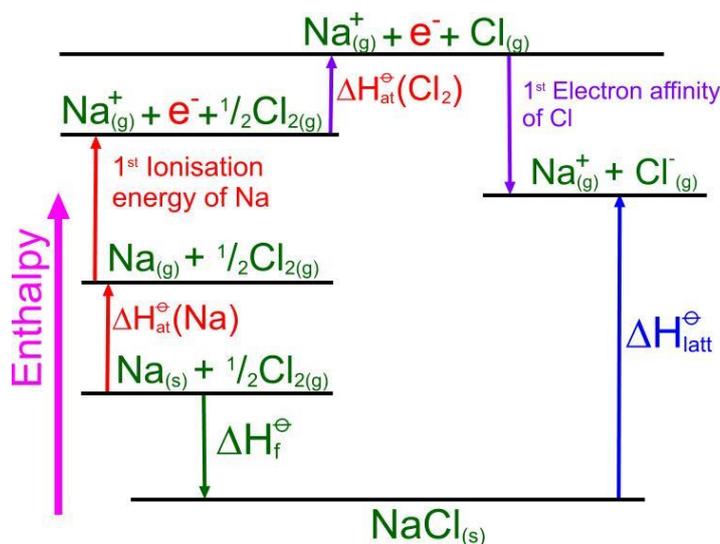
32 Write equations representing the following enthalpy changes:

- The first ionisation energy of magnesium $Mg(g) \rightarrow Mg^+(g) + e^-$
- The atomisation energy of chlorine $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$
- The second electron affinity of oxygen $O^-(g) + e^- \rightarrow O^{2-}(g)$
- The lattice enthalpy of magnesium chloride $MgCl_2(s) \rightarrow Mg^{2+}(g) + 2Cl^-(g)$
- The hydration enthalpy of chloride ions $Cl^-(g) \rightarrow Cl^-(aq)$
- The enthalpy change of solution of sodium carbonate $Na_2CO_3(s) \rightarrow 2Na^+(aq) + CO_3^{2-}(aq)$

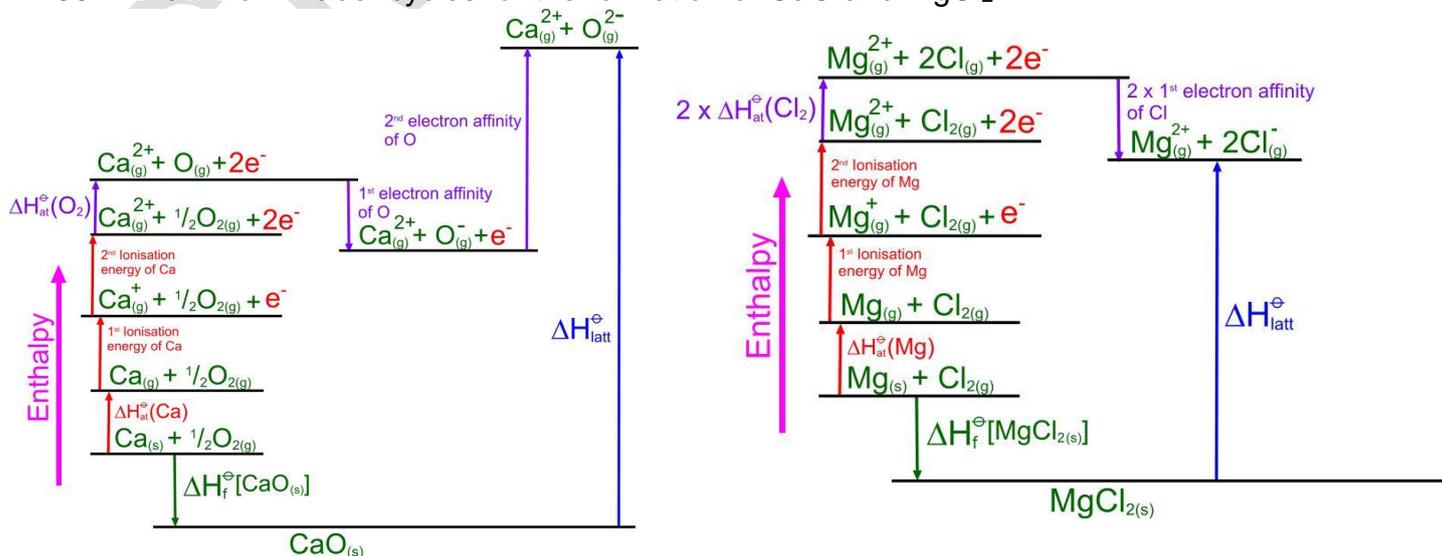
33 Work out enthalpy changes of atomisation from bond enthalpies. The F-F bond enthalpy is 160 kJ mol^{-1} , what is the value of the enthalpy change of atomisation of fluorine? 80 kJ mol^{-1}

34 Draw a Born-Haber cycle for the formation of sodium chloride and the data in the table to calculate the standard enthalpy change of formation of sodium chloride. -430 kJ mol^{-1}

$\Delta H_{\text{at}}(\text{Na})$	100 kJ mol^{-1}
Cl-Cl bond enthalpy	240 kJ mol^{-1}
IE (Na)	500 kJ mol^{-1}
EA (Cl)	-350 kJ mol^{-1}
$\Delta H_{\text{latt}}(\text{NaCl})$	800 kJ mol^{-1}



35 Draw Born Haber cycles for the formation of CaO and $MgCl_2$



- 36 Calculate the enthalpy change of solution of magnesium chloride from hydration enthalpies and the lattice enthalpy
-140 kJ mol⁻¹

$\Delta H_{\text{hyd}}(\text{Mg}^{2+})$	-1960 kJ mol ⁻¹
$\Delta H_{\text{hyd}}(\text{Cl}^-)$	-360 kJ mol ⁻¹
$\Delta H_{\text{latt}}(\text{MgCl}_2)$	2540 kJ mol ⁻¹

- 37 Understand the factors that affect the values of lattice enthalpy and enthalpy changes of hydration.

(a) Place the following compounds in order of increasing lattice enthalpy

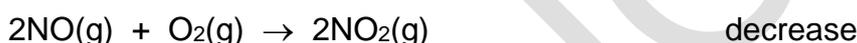
KF MgCl₂ MgF₂ CaO CaS KI
(lowest) KI KF MgCl₂ MgF₂ CaS CaO (highest)

(b) For each of the following pairs state which has the **more exothermic** enthalpy of hydration.

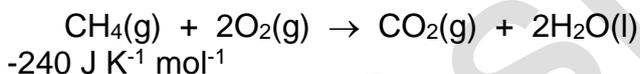
- Na⁺ and Mg²⁺
- Na⁺ and K⁺
- CO₃²⁻ and Cl⁻

- 38 Explain what is meant by *entropy* in terms of the distribution of energy.
Entropy is a measure of how the available energy is distributed among the particles.

- 39 State and explain whether each of the reactions below involves an increase or decrease in entropy and state the sign of ΔS for each reaction.



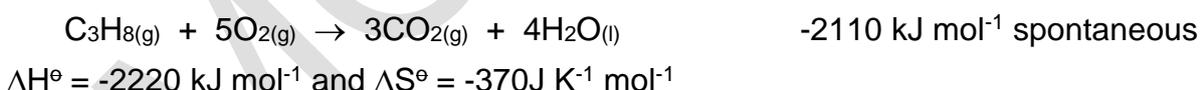
- 40 Use the absolute entropies given below to calculate the standard entropy change for



Substance	S ^o / J K ⁻¹ mol ⁻¹
CH ₄ (g)	185
O ₂ (g)	205
CO ₂ (g)	215
H ₂ O(l)	70

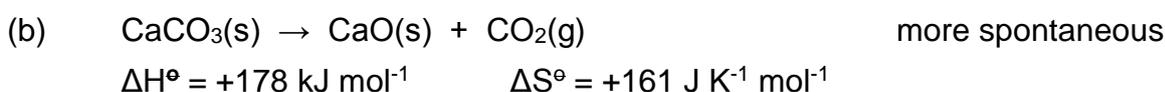
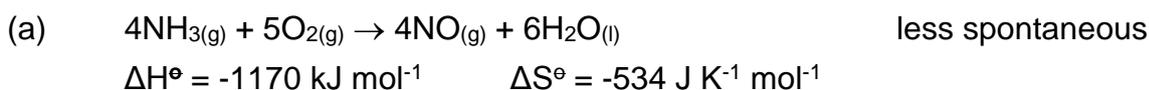
- 41 State whether ΔG is positive or negative for a spontaneous reaction. negative

- 42 Calculate ΔG° for the following reaction and state whether it is spontaneous or not at 298 K



- 43 Explain whether a reaction for which ΔS is positive becomes more or less spontaneous as the temperature increases. More spontaneous

- 44 Explain whether each of the following reactions becomes more or less spontaneous as the temperature increases:



- 45 Work out the temperature (°C) at which the reaction below becomes spontaneous

