

2. In an experiment to measure the enthalpy change of combustion of ethanol, a student heated a copper calorimeter containing 100 cm³ of water with a spirit lamp and collected the following data.

Treat as 3 sig figs!

| | | |
|-------------------------------|---------------------------|--|
| Initial temperature of water: | 20.0 °C | } $\Delta T = 55.0 - 20.0 = 35.0\text{ }^\circ\text{C}$ |
| Final temperature of water: | 55.0 °C | |
| Mass of ethanol burned: | 1.78 g | |
| Density of water: | → 1.00 g cm ⁻³ | (∴ 1 cm ³ H ₂ O = 1 g H ₂ O!) |

(i) Use the data to calculate the heat evolved when the ethanol was combusted.

$$q = cm\Delta T$$

$$= (4.18 \text{ J/g}\cdot^\circ\text{C}) (100. \text{ g}) (35.0\text{ }^\circ\text{C}) = \boxed{14600 \text{ J}}$$

↑ value IB uses!

$$\text{or } \boxed{14.6 \text{ kJ}}$$

(ii) Calculate the enthalpy change of combustion per mole of ethanol.

$$1.78 \text{ g} \times \frac{1 \text{ mol}}{46.08 \text{ g}} = 0.0386 \text{ mol}$$

$$8 \times 17 = -14.6 \text{ kJ}$$

$$\Delta H = \frac{-14.6 \text{ kJ}}{0.0386 \text{ mol}} = \boxed{-378 \text{ kJ/mol}}$$

(-378 kJ mol⁻¹)

(iii) Suggest two reasons why the result is not the same as the value in the Data Booklet.

Any two:

- Heat loss by the calorimeter
- Incomplete combustion
- Heat absorbed by the calorimeter itself
- other sensible answers may be accepted.

(Total 6 marks)

3. An experiment was designed to investigate how the enthalpy change for a displacement reaction relates to the reactivities of the metals involved. The following metals in order of decreasing reactivity were available.

| | |
|-----------|----------------|
| Magnesium | most reactive |
| Uranium | |
| Zinc | |
| Iron | |
| Copper | |
| Silver | least reactive |

↓

Excess amounts of each metal were added to 1.00 mol dm⁻³ copper(II) sulfate solution. The temperature change was measured and the enthalpy change calculated.

(i) Suggest a possible hypothesis for the relationship between the enthalpy change of the reaction and the reactivity of the metal.

The more reactive the metal,
the more exothermic the reaction.

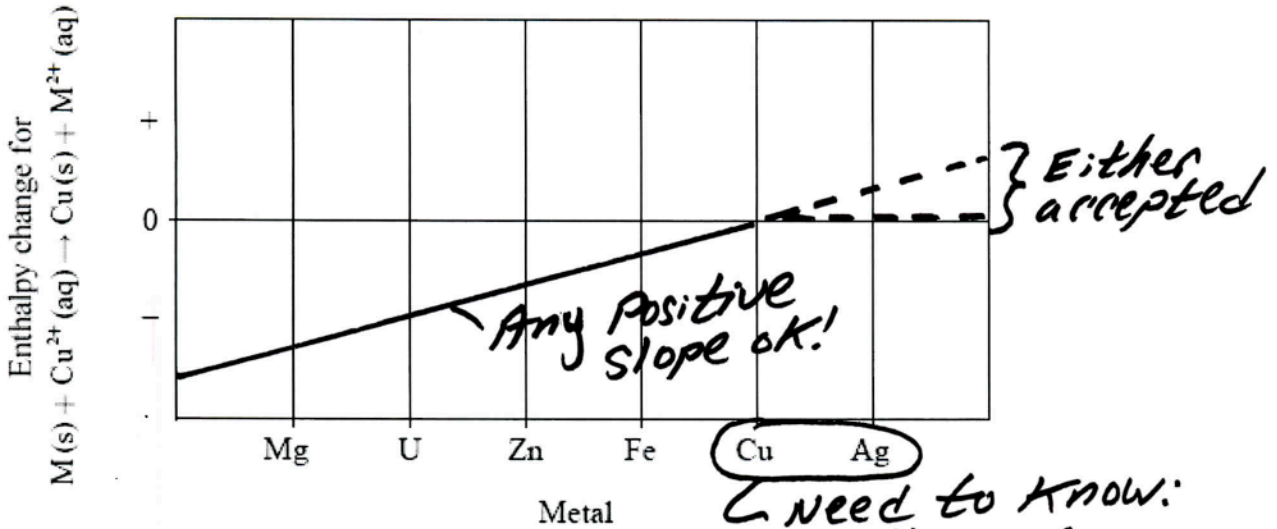
Note: Always exothermic whenever an element is a reactant!

(1)

2

(cont.)

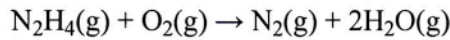
(ii) Sketch a graph on the diagram below to illustrate your hypothesis.



Need to know:
 $Cu + Cu^{2+} \rightarrow \text{NO RXN}$ (1)
 (Total 2 marks)
 $*Ag + Cu^{2+} \rightarrow \text{NO RXN}$

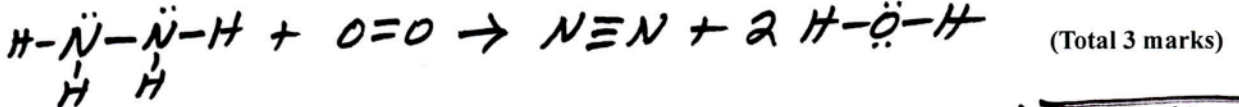
4. Hydrazine is a valuable rocket fuel.

The equation for the reaction between hydrazine and oxygen is given below.



*recall the $AgNO_3 + Cu$
 L6!
 $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$
 $\therefore Ag + Cu^{2+} \rightarrow \text{NO RXN!}$

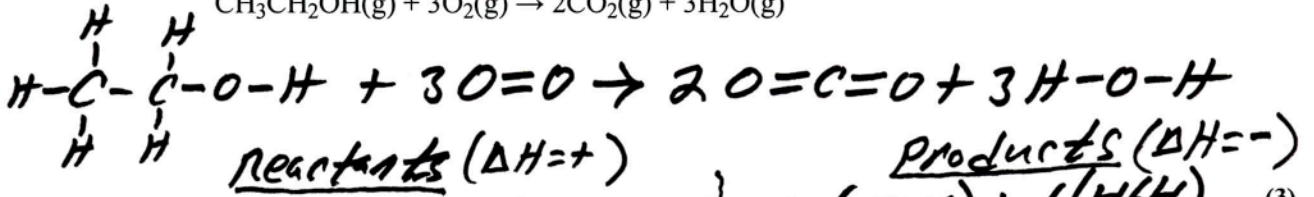
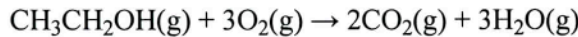
Use the bond enthalpy values from Table 10 of the Data Booklet to determine the enthalpy change for this reaction.



$$\underbrace{4(391) + (158) + (498)}_{\Delta H = +} + \underbrace{(-945) + 4(-464)}_{\Delta H = -} = \boxed{-581 \text{ kJ}}$$

5. In some countries, ethanol is mixed with gasoline (petrol) to produce a fuel for cars called gasohol.

(i) Use the information from Table 10 of the Data Booklet to determine the standard enthalpy change for the complete combustion of ethanol.



$$\underbrace{5(413) + (347) + (358) + (464)}_{\substack{\text{C-H} \\ \text{C-O} \\ \text{C-O} \\ \text{O-H}}} + \underbrace{4(-746) + 6(464)}_{\substack{\text{C=O} \\ \text{H-O}}} + 3(498)_{\text{O=O}} = \boxed{-1040 \text{ kJ}}$$

3

(cont.)

- (ii) The standard enthalpy change for the complete combustion of octane, C_8H_{18} , is $-5471 \text{ kJ mol}^{-1}$. Calculate the amount of energy produced in kJ when 1 g of ethanol and 1 g of octane is burned completely in air.

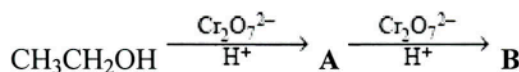
Treat as ∞ sig figs!

Ethanol: $1g \times \frac{1 \text{ mol}}{46.08g} \times \frac{1040 \text{ kJ}}{1 \text{ mol}} = \boxed{22.57 \text{ kJ}}$

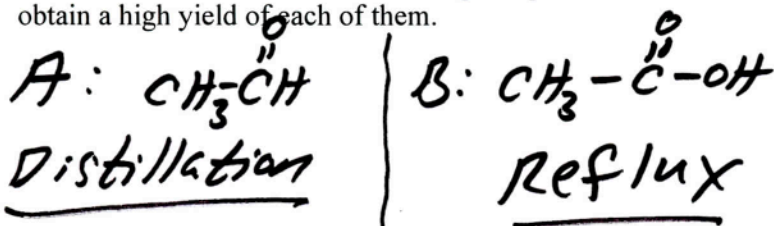
Octane: $1g \times \frac{1 \text{ mol}}{114.26g} \times \frac{5471 \text{ kJ}}{1 \text{ mol}} = \boxed{47.88 \text{ kJ}}$

(2)
Do Not
Add
together!

- (iii) Ethanol can be oxidized using acidified potassium dichromate, $K_2Cr_2O_7$, to form two different organic products.



State the structural formulas of the organic products A and B and describe the conditions required to obtain a high yield of each of them.

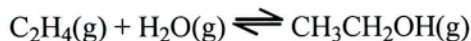


Sorry - (3)
Not
energetic
questions! (2)

- (iv) Deduce and explain whether ethanol or A has the higher boiling point.

Ethanol has the higher boiling point due to hydrogen bonding.

- (v) Ethene can be converted into ethanol by direct hydration in the presence of a catalyst according to the following equation.



For this reaction identify the catalyst used and state one use of the ethanol formed other than as a fuel.

H_2SO_4 (or concentrated H_3PO_4)

uses: - Dyes
- Drugs
- cosmetics
- solvent
- Disinfectant

(2)

(Total 12 marks)

HL

1. 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.

(Total 1 mark)

2. Propene can be hydrogenated in the presence of a nickel catalyst to form propane. Use the data below to answer the questions that follow.

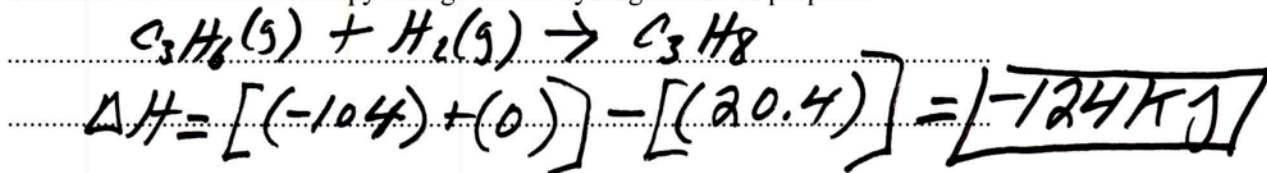
| Compound | Formula | $\Delta H_f^\ominus / \text{kJ mol}^{-1}$ | $S^\ominus / \text{J K}^{-1} \text{mol}^{-1}$ |
|----------|----------------------------------|---|---|
| hydrogen | $\text{H}_2(\text{g})$ | 0 | + 131 |
| propane | $\text{C}_3\text{H}_8(\text{g})$ | - 104 | + 270 |
| propene | $\text{C}_3\text{H}_6(\text{g})$ | + 20.4 | + 267 |

- (i) Outline why the value for the standard enthalpy change of formation of hydrogen is zero.

By definition, ΔH_f^\ominus is the enthalpy associated with the formation of one mole of a compound. $\text{H}_2(\text{g})$ is not a compound.

(1)

- (ii) Calculate the standard enthalpy change for the hydrogenation of propene.



(2)

- (iii) Calculate the standard entropy change for the hydrogenation of propene.

$$\Delta S = [(270)] - [(267) + (131)] = \boxed{-128 \text{ J/K}}$$

units must agree!

(2)

- (iv) Determine the value of ΔG^\ominus for the hydrogenation of propene at 298 K.

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

$$= (-124 \text{ kJ}) - 298 \text{ K} (-0.128 \text{ kJ/K})$$

$$= \boxed{-86.3 \text{ kJ}}$$

(2)

(cont.)

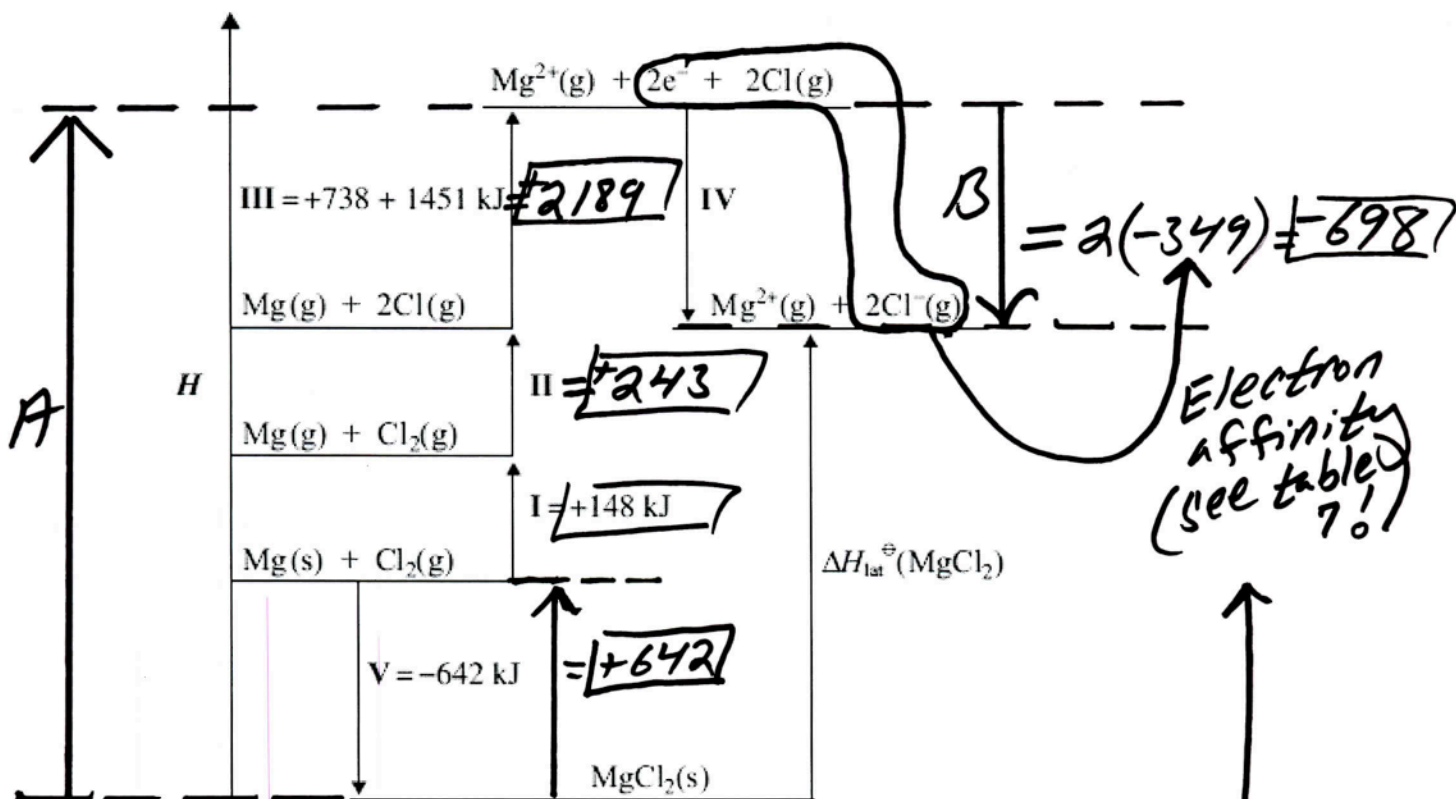
- (v) At 298 K the hydrogenation of propene is a spontaneous process. Determine the temperature above which propane will spontaneously decompose into propene and hydrogen.

When $\Delta G = 0$

$$T = \frac{\Delta H}{\Delta S} = \frac{-124 \text{ kJ}}{-0.128 \text{ kJ/K}} = \boxed{972 \text{ K}}$$

(2)
(Total 9 marks)

3. The lattice enthalpy of magnesium chloride can be calculated from the Born-Haber cycle shown below.



- (i) Identify the enthalpy changes labelled by I and V in the cycle.



(2)

- (ii) Use the ionization energies given in the cycle above and further data from the Data Booklet to calculate a value for the lattice enthalpy of magnesium chloride.

$$\Delta H_{\text{lat}} = \underbrace{642 + 148 + 243 + 2189}_{A, \Delta H = +} - \underbrace{698}_{B, \Delta H = -} = \boxed{2524 \text{ kJ}}$$

(4)

(cont.)

- (iii) The theoretically calculated value for the lattice enthalpy of magnesium chloride is +2326 kJ. Explain the difference between the theoretically calculated value and the experimental value.

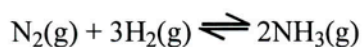
Theoretical values assume perfectly spherical ions with only electrostatic forces between them. There is, however, some covalent character. (Please read p182 of your text!) (2)

- (iv) The experimental lattice enthalpy of magnesium oxide is given in Table 13 of the Data Booklet. Explain why magnesium oxide has a higher lattice enthalpy than magnesium chloride.

Oxide has both a greater charge (+2) and a smaller radius than chloride (-1).

(2)
(Total 10 marks)

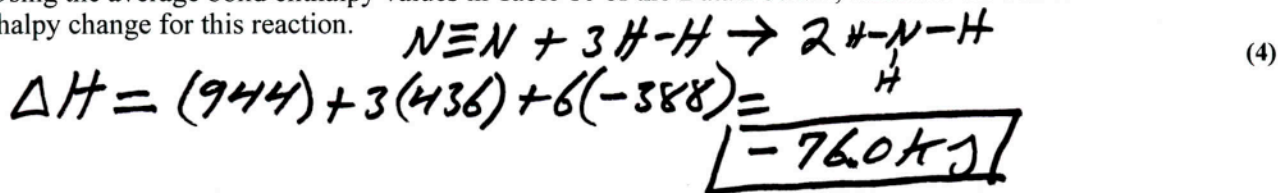
4. Consider the following reaction:



- (i) Suggest why this reaction is important for humanity.

Ammonia is used in fertilizers that increase crop yields; production of explosives for mining, tunnels, demolition. (1)

- (ii) Using the average bond enthalpy values in Table 10 of the Data Booklet, calculate the standard enthalpy change for this reaction.



- (iii) 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 $\text{J K}^{-1} \text{mol}^{-1}$ respectively. Calculate ΔS° for the reaction and explain the sign of ΔS° .

$$\Delta S = [2(193)] - [(192) + 3(131)] = \boxed{-199 \text{ J/K}}$$

There is a decrease in entropy ($-\Delta S$) because four moles of gas ($\text{N}_2 + 3\text{H}_2$) react to form two moles of gas (2NH_3). (2)

(cont.)

- (iv) Calculate ΔG^\ominus for the reaction at 238 K. State and explain whether the reaction is spontaneous.

$$\Delta G = \Delta H - T\Delta S$$
$$= -76.0 \text{ kJ} - 238 \text{ K}(-0.199 \text{ kJ/K}) = \boxed{-16.7 \text{ kJ}}$$

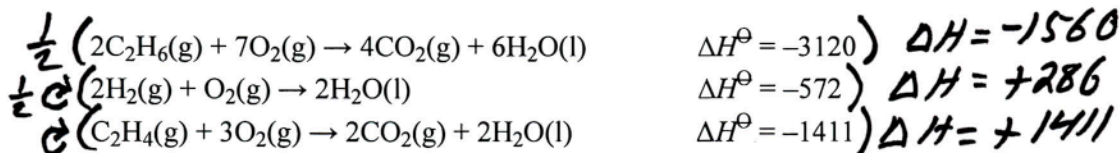
ΔG is negative, so the reaction is spontaneous. (3)

- (v) If ammonia was produced as a liquid and not as a gas, state and explain the effect this would have on the value of ΔH^\ominus for the reaction.

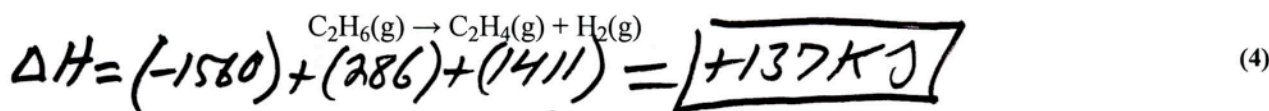
more energy is released in the formation of a liquid (bonds form) than a gas, so ΔH_f^\ominus for $\text{NH}_3(\text{l})$ would have a greater negative value than $\text{NH}_3(\text{g})$, therefore $\Delta H_{\text{rxn}}^\ominus$ would be more negative. (2) (Total 12 marks)

5. (i) The standard enthalpy change of three combustion reactions is given below in kJ.

Hess's Law!



Based on the above information, calculate the standard change in enthalpy, ΔH^\ominus , for the following reaction.



- (ii) Predict, stating a reason, whether the sign of ΔS^\ominus for the above reaction would be positive or negative. (2)

ΔS would be positive,
1 mol of gas \rightarrow 2 mol of gas

- (iii) Discuss why the above reaction is non-spontaneous at low temperature but becomes spontaneous at high temperatures. (2)

At low temperatures $\Delta H > T\Delta S$,
therefore ΔG will be *positive.
At high temperatures $\Delta H < T\Delta S$,
and ΔG will be *negative.

$$*\Delta G = (+) - T(+)$$

(iv) Using bond enthalpy values, calculate ΔH^\ominus for the following reaction.



(3)

$$\Delta H = 6(\underset{\text{C-H}}{413}) + (\underset{\text{C-C}}{347}) + 4(\underset{\text{C-H}}{-413}) + (\underset{\text{C=C}}{612}) + (\underset{\text{H-H}}{-436}) = \boxed{+125 \text{ kJ}}$$

(v) Suggest with a reason, why the values obtained in parts (i) and (iv) are different.

Bond enthalpies are average values and does not take into account other bonding on an atom.

(1)
(Total 12 marks)