1. A student is asked to calculate ΔG at 25 °C for the combustion of butan-1-ol. The teacher provides two pieces of information.

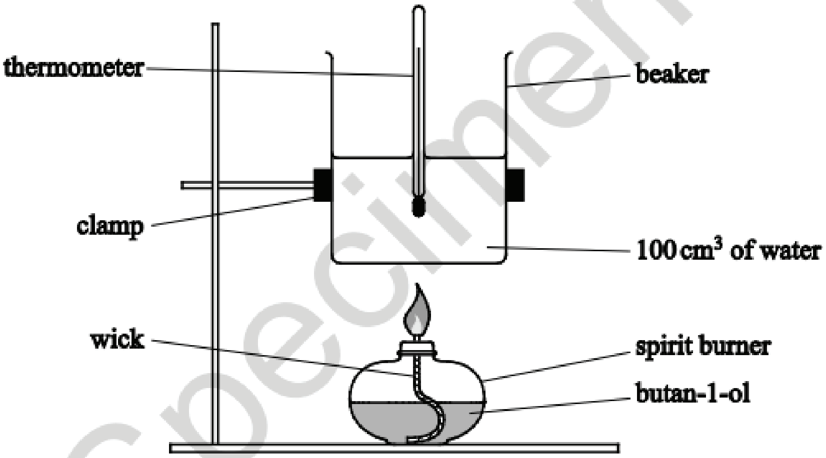
- The equation for the combustion of butan-1-ol.

$$CH_3(CH_2)_3OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l) \qquad \textbf{Equation 2}$$

- Standard entropies of butan-1-ol, oxygen, carbon dioxide and water.

	$CH_3(CH_2)_3OH(l)$	$O_2(g)$	$CO_2(g)$	$H_2O(l)$
S° / J K^{-1} mol^{-1}	228	205	214	70

The student carries out an experiment using the apparatus below and obtains the following results. The specific heat capacity of water is 4.18 J g^{-1} K^{-1}.



Mass of burner and butan-1-ol before burning / g	98.997
Mass of burner and butan-1-ol after burning / g	98.738
Initial temperature / °C	18.5
Maximum temperature reached / °C	39.0

Use the information on the previous page to calculate Δ*G*, in kJ mol^{-1}, for the combustion of butan-1-ol according to **Equation 2** at 25 °C.

Show **all** your working.

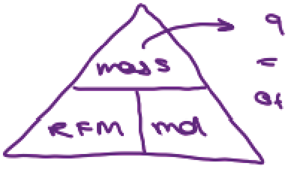
$q = mc\Delta T$

energy

mass of water

specific heat capacity

temperature change



98.997 − 98.738 = 0.259 g of butan-1-ol

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

$q = 100 \times 4.18 \times (39 - 18.5) = 8569\,J \rightarrow 8.569\,kJ$

moles of butan-1-ol: $\frac{0.259}{74} = 3.5 \times 10^{-3}$ mol

$\Delta S = \Delta S_{products} - \Delta S_{reactants}$

$\Delta S = (4 \times 214) + (5 \times 70) - [(228) + (6 \times 205)] = -252\ J K^{-1} mol^{-1}$

$\Delta S = -0.252\ kJ K^{-1} mol^{-1}$

$\Delta H = \frac{8.569}{3.5 \times 10^{-3}} = -2448\ kJ mol^{-1}$

$\Delta G = -2448 - (298 \times -0.252) = -2373\ kJ mol^{-1}$

Δ*G* = ...−2373.......... kJ mol^{-1} **[7]**

2. This question is about free energy changes, ΔG, enthalpy changes, ΔH, and temperature, T.

(a) The Gibbs' equation is shown below.

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

A chemist investigates a reaction to determine how ΔG varies with T.
The results are shown in **Fig. 18.1**.

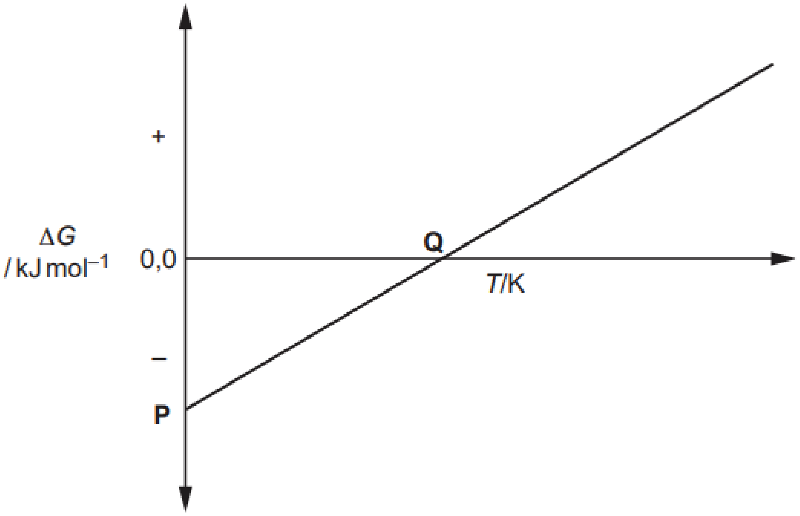


Fig. 18.1

What is significant about the gradient of the line and the values **P** and **Q** shown in **Fig. 18.1**?
Explain your reasoning.

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

$y = c + mx$

gradient = $-\Delta S$

P (y intercept) = ΔH

Q = Temperature where feasibility changes

[4]

(b) Iron can be extracted from its ore Fe_3O_4 using carbon.
Several equilibria are involved including **equilibrium 18.1**, shown below.

equilibrium 18.1 $Fe_3O_4(s) + 4C(s) \rightleftharpoons 3Fe(s) + 4CO(g)$ $\Delta H = +676.4\ kJ\ mol^{-1}$
$\Delta S = +703.1\ J\ K^{-1}\ mol^{-1}$

homogeneous state all species in the same

(i) Why is **equilibrium 18.1** a *heterogeneous* equilibrium?

species in different states / phases

[1]

(ii) Write the expression for K_p for **equilibrium 18.1**.

$$K_p = p(CO_{(g)})^4$$ ← stoichiometry

[1]

(iii) The forward reaction in **equilibrium 18.1** is only feasible at high temperatures.

- Show that the forward reaction is **not** feasible at 25 °C. ← +273 = 298 K
- Calculate the minimum temperature, in K, for the forward reaction to be feasible.

$$\uparrow$$
$$\frac{\Delta H}{\Delta S}$$

J → kJ

$$\Delta G = \Delta H - T\Delta S$$
$$\Delta G = 676.4 - 298 \times 0.7031 = 467\ kJ\ mol^{-1}$$

$\Delta G > 0$ so not feasible

$$\text{min. temp} = \frac{676.4}{0.7031} = 962\ K$$

minimum temperature = 962 K **[3]**

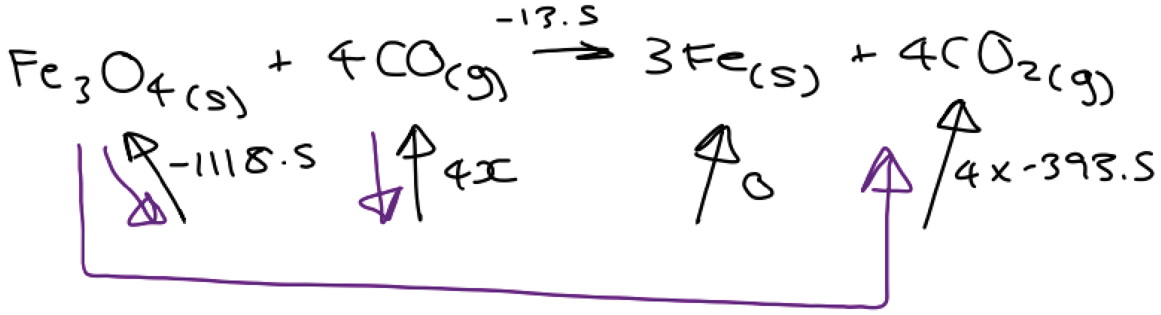
(iv) Another equilibrium involved in the extraction of iron from Fe_3O_4 is shown below.

$$Fe_3O_4(s) + 4CO(g) \rightleftharpoons 3Fe(s) + 4CO_2(g) \qquad \Delta H = -13.5\,kJ\,mol^{-1}$$

Enthalpy changes of formation, $\Delta_f H$, for $Fe_3O_4(s)$ and $CO_2(g)$ are shown in the table.

Compound	**$\Delta_f H$ / kJ mol^{-1}**
$Fe_3O_4(s)$	–1118.5
$CO_2(g)$	–393.5

Calculate the enthalpy change of formation, $\Delta_f H$, for CO(g).



$$(1118.5 - 4x) + (0 + (4 \times -393.5)) = -13.5$$

$$1118.5 - 4x = 1560.5$$

$$-4x = 442$$

$$x = -110.5\ kJmol^{-1}$$

$\Delta_f H$, for CO(g) = ... -110.5 .... kJ mol^{-1} **[3]**

3. The table below shows standard entropies, S°.

Substance	$CO(g)$	$H_2(g)$	$CH_3OH(l)$
S° / J mol^{-1} K^{-1}	197.6	130.6	239.7

What is the entropy change, ΔS°, in J mol^{-1} K^{-1}, for the following reaction?

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$$

p − r

$239.7 - (197.6 + (2 \times 130.6))$

$= -219.1$ J mol^{-1} K^{-1}

A –219.1

B –88.5

C +88.5

D +219.1

Your answer A **[1]**

4. A student carries out two experiments in the laboratory based on succinic acid (butanedioic acid), $(CH_2COOH)_2$.

(a) Aqueous succinic acid can be neutralised by aqueous sodium hydroxide, NaOH(aq):

$$(CH_2COOH)_2(aq) + 2NaOH(aq) \rightarrow (CH_2COONa)_2(aq) + 2H_2O(l)$$

This reaction can be used to determine a value for the enthalpy change of neutralisation, $\Delta_{neut}H$.

energy released to produce 1 mol of water there are ② so ÷ by 2 at the end of the calculation

The student follows this method:

- Add 50.0 cm^3 of 0.400 mol dm^{-3} succinic acid to a polystyrene cup.
- Measure out 50.0 cm^3 of 1.00 mol dm^{-3} NaOH(aq), which is in excess.
- Measure the temperature of both solutions.
- Add the NaOH(aq) to the aqueous succinic acid in the polystyrene cup, stir the mixture, and record the maximum temperature.

Temperature readings

Maximum temperature of mixture/°C	26.5
Initial temperature of both solutions/°C	21.5

Calculate a value for the enthalpy change of neutralisation, $\Delta_{neut}H$, in kJ mol^{-1}.

Assume that the density of all solutions and the specific heat capacity, *c*, of the reaction mixture are the same as for water.

n
v×c

$50\times10^{-3} \times 0.4 = 0.02$ mol

$Q = mc\Delta T \leftarrow 26.5 - 21.5 = 5$ °c

$\uparrow$ 100 cm^3 $\nwarrow$ 4.18

$100 \times 4.18 \times 5 = 2090$ J $= 2.090$ kJ

$\frac{2.090}{0.02} = \pm104.5$ kJmol^{-1}

$\frac{104.5}{2} = \pm52.3$ kJmol

enthalpy change of neutralisation is -ve

$\Delta_{neut}H$ = – 52.3 kJ mol^{-1} **[4]**

(b) Succinic acid is esterified by ethanol, C_2H_5OH, in the presence of an acid catalyst to form an equilibrium mixture.

The equilibrium constant, K_c, for this equilibrium can be calculated using the amounts, in moles, of the components in the equilibrium mixture, using **expression 5.1**.

Ester

$$Kc = \frac{\text{Products}}{\text{reactants}} \quad K_c = \frac{n((CH_2COOC_2H_5)_2) \times n(H_2O)^{2}}{n((CH_2COOH)_2) \times n(C_2H_5OH)^{2}}$$

Expression 5.1

Powers tells you the stoichiometry of the equilibrium

A student carries out an experiment to determine the value of K_c for this equilibrium.

- The student mixes together 0.0500 mol of succinic acid and 0.150 mol of ethanol, with a small amount of an acid catalyst.
- The mixture is allowed to reach equilibrium.
- The student determines that 0.0200 mol of succinic acid are present in the equilibrium mixture.

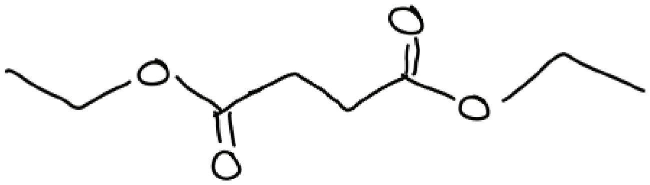
(i) Which technique could be used to determine the equilibrium amount of succinic acid?

titration **[1]**

(ii) Write the equation for the equilibrium reaction that takes place.

$(CH_2COOH)_2 + 2C_2H_5OH \rightleftharpoons (CH_2COOC_2H_5)_2 + 2H_2O$ **[1]**

(iii) Draw the skeletal formula of the ester present in the equilibrium mixture.



[1]

(iv) K_c is the equilibrium constant in terms of equilibrium concentrations.

Why can **expression 5.1** be used to calculate K_c for this equilibrium?

volumes cancel out / same number of moles on each side of the equation **[1]**

(v) Calculate the value of K_c for this reaction.

reactants have opposite change to products

Show your working.

	$(CH_2COOH)_2$	C_2H_5OH	$(CH_2COOC_2H_5)_2$	H_2O
I	0.05	0.15	0	0
C	-0.03	-0.06	+0.03	+0.06
E	0.02	0.09	0.03	0.06

double the change

K_c = 0.67 **[3]**

$$kc = \frac{0.03 \times 0.06^2}{0.02 \times 0.09^2} = 0.67$$

5. Sir Humphry Davy discovered several elements including sodium, potassium, magnesium, calcium and strontium.

(a) Explain which block in the Periodic Table sodium and magnesium belong to.

S blook because the highest energy is in a s-orbital **[1]**

(b) A sample of magnesium, $A_r = 24.305$, is found to consist of three isotopes. The accurate relative isotopic masses and % abundances of two of the isotopes are shown in the table.

Isotope	Relative isotopic mass	% abundance
^{24}Mg	23.985	78.99%
^{25}Mg	24.986	10.00%

Determine the relative isotopic mass of the third isotope of magnesium in the sample.

Give your answer to **5** significant figures.

$$\left(\frac{78.99 \times 23.985}{100}\right) + \left(\frac{10 \times 24.986}{100}\right) + \left(\frac{11.01 \times m}{100}\right)$$

$= 24.305$

$$\frac{11.01 \times m}{100} = 2.8606485$$

$11.01 \times m = 286.06485$

$m = 25.982$ (5 sf.)

relative isotopic mass = 25.982 **[2]**

(c) A student adds an excess of calcium oxide to water in a test tube.
In a separate test tube, the student adds an excess of strontium oxide to water.

(i) Write the equation for the reaction of calcium oxide with water.

State symbols are **not** required.

$CaO + H_2O \rightarrow Ca(OH)_2$ **[1]**

(ii) Suggest the approximate pH of the two solutions formed in the test tubes.

pH with calcium oxide 12

pH with strontium oxide 13

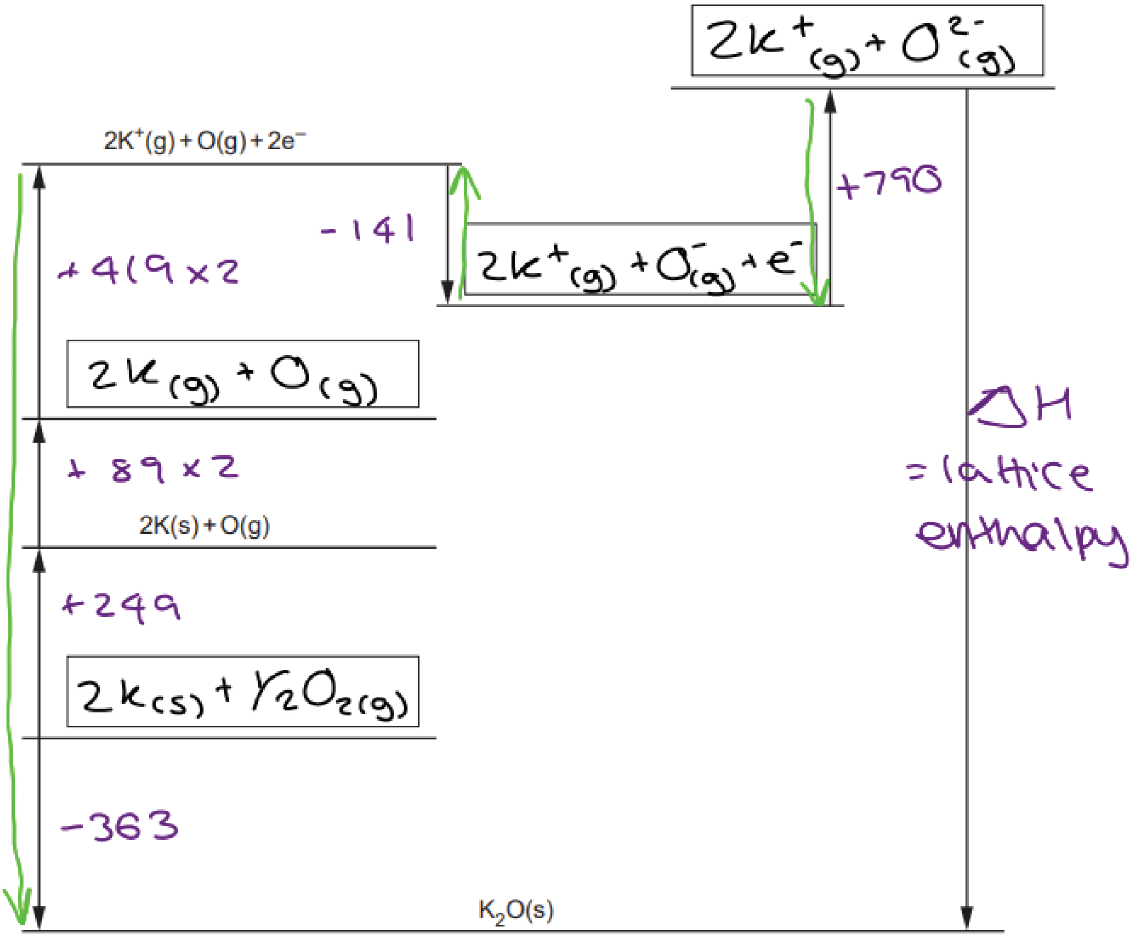
[1]

(d) The table below shows enthalpy changes involving potassium, oxygen and potassium oxide, K_2O.

	Enthalpy change / kJ mol^{-1}
formation of potassium oxide	–363
1st electron affinity of oxygen	–141
2nd electron affinity of oxygen	+790
1st ionisation energy of potassium	+419
atomisation of oxygen	+249
atomisation of potassium	+89

(i) The incomplete Born–Haber cycle below can be used to determine the lattice enthalpy of potassium oxide.

In the boxes, complete the species present in the cycle.
Include state symbols for the species.



[4]

(ii) Calculate the lattice enthalpy of potassium oxide.

$-790 + 141 - (2 \times 419) - (89 \times 2) - 249$
$- 363 = \quad -2277$ kJmol^{-1}

lattice enthalpy = -2277 kJ mol^{-1} **[2]**

(e) A similar Born–Haber cycle to potassium oxide in **(d)** can be constructed for sodium oxide.

(i) The first ionisation energy of sodium is more endothermic than that of potassium.

Explain why.

Sodium has a smaller atomic radius so sodium' nuclear attraction increases

[2]

(ii) The lattice enthalpy of sodium oxide is more exothermic than that of potassium oxide.

Explain why.

For sodium ions the ionic radius is smaller so, Na^+ has a stronger attraction to O^{2-}.

[2]

6. Sulfuric acid is an important chemical used to make detergents, fertilisers and dyes. It is manufactured in a multi-step process.

(a) In the first step of the manufacture of sulfuric acid, sulfur dioxide, SO_2, can be made from the combustion of hydrogen sulfide, H_2S, shown in **Reaction 1**.

$2H_2S(g) + 3O_2(g) \rightarrow 2SO_2(g) + 2H_2O(l)$ $\Delta_r H = -1125\,kJ\,mol^{-1}$ **Reaction 1**

(i) Explain why the enthalpy change for **Reaction 1** has a negative value.

Use ideas about enthalpy changes associated with bond breaking and bond making.

More energy released by forming bonds than required when breaking bonds

[1]

(ii) Some standard entropy values are given below.

Substance	$H_2S(g)$	$O_2(g)$	$SO_2(g)$	$H_2O(l)$
$S^\ominus / J\,K^{-1}\,mol^{-1}$	206	205	248	70

Using calculations, explain whether **Reaction 1** is feasible at 20 °C.

Calculations

$\Delta S = ((2 \times 248) + (2 \times 70)) - ((2 \times 206) + (3 \times 205))$

$\Delta S = -391\ JK^{-1}mol^{-1}$

$= -0.391\ kJK^{-1}mol^{-1}$

$\Delta G = -1125 - ((273 + 20) \times -0.391)$

$\Delta G = -1010\ kJmol^{-1}$

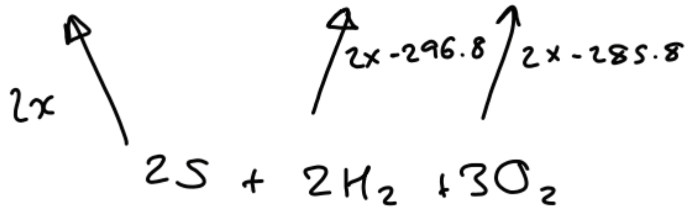
Explanation for feasible or non feasible feasible because $\Delta G < 0$

[4]

(iii) Calculate the standard enthalpy change of formation, $\Delta_f H^\ominus$, of hydrogen sulfide using the enthalpy change for **Reaction 1**, and the standard enthalpy changes of combustion below.

Substance	$\Delta_c H^\ominus$ / kJ mol^{-1}
S(s)	−296.8
H_2(g)	−285.8

$2H_2S(g) + 3O_2(g) \xrightarrow{-1125} 2SO_2(g) + 2H_2O(l)$ $\Delta_r H = -1125\ \text{kJ mol}^{-1}$ **Reaction 1**



$(-296.8 \times 2) + (-285.8 \times 2) + 1125 = -40.2 = 2x$

$x = -\frac{40.2}{2} = -20.1$

$\Delta_f H^\ominus$ of hydrogen sulfide = −20.1 kJ mol^{-1} **[3]**

(b) The second step in the manufacture of sulfuric acid is the conversion of SO_2 into sulfur trioxide, SO_3, using **Equilibrium 1**.

exothermic

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ $\Delta H = -197\ kJ\ mol^{-1}$ **Equilibrium 1**

An industrial chemist carries out some research into **Equilibrium 1**.

- The chemist fills a 10.2 dm^3 container with $SO_2(g)$ at RTP, and then adds 12.0 g of $O_2(g)$.
- The chemist adds the vanadium(V) oxide catalyst, and heats the mixture. The mixture is allowed to reach equilibrium at a pressure of 2.50 atm and a temperature of 1000 K.
- A sample of the equilibrium mixture is analysed, and found to contain 0.350 mol of SO_3.

(i) Write an expression for K_p for **Equilibrium 1**.

Include the units.

$$\frac{atm^2}{atm^3} = \frac{1}{atm}$$

$$K_p = \frac{p(SO_3)^2}{p(SO_2)^2 \times p(O_2)}$$

units = atm^{-1} **[2]**

(ii) Determine the value of K_p for **Equilibrium 1** at 1000 K.

Show all your working.

Give your answer to **3** significant figures.

RTP

$$\frac{\text{volume}(dm^3)}{24} = \text{mol}$$

$$\frac{10.2}{24} = 0.425 \text{ mol of } SO_2$$

$$\frac{12}{32} = 0.375 \text{ mol of } O_2$$

partial pressure = mol fraction × total pressure

	SO_2	O_2	SO_3
I	0.425	0.375	0
C	−0.350	$-\frac{0.35}{2}$	+0.350
E	0.075	0.200	0.350

total moles = 0.625 mol

$$p(SO_2) = \frac{0.075}{0.625} \times 2.5 = 0.3\ atm$$

$$p(O_2) = \frac{0.2}{0.625} \times 2.5 = 0.8\ atm$$

$$p(SO_3) = \frac{0.35}{0.625} \times 2.5 = 1.4\ atm$$

$$K_p = \frac{(1.4)^2}{(0.3)^2 \times (0.8)}$$

$$K_p = 27.2\ atm^{-1}$$

K_p = 27.2 **[5]**

(iii) The chemist repeats the experiment in **(b)** at a different temperature.

The chemist finds that the value of K_p is greater than the answer to **(b)(ii)**.

Explain whether the temperature in the second experiment is higher or lower than 1000 K.

Greater Kp value means equilibrium position shifted to the right so lower temperature because forward reaction is exothermic **[2]**

(iv) Explain the significance of the expression: $K_p \gg 1$.

equilibrium position far to the right **[1]**

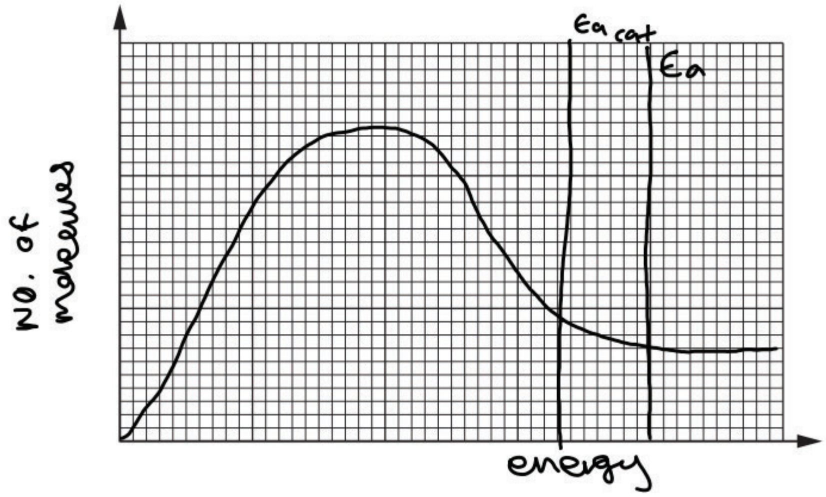
(c) Vanadium(V) oxide, $V_2O_5(s)$, is used as a catalyst in **equilibrium 1**.

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H = -197\,kJ\,mol^{-1}$ **Equilibrium 1**

(i) Explain how the presence of $V_2O_5(s)$ increases the rate of reaction.

Include a labelled sketch of the Boltzmann distribution, on the grid below.

Label the axes.



more molecules collide with energy above activation energy (with catalyst).

[4]

(ii) Explain whether vanadium(V) oxide is acting as a homogeneous or heterogeneous catalyst.

Heterogenous because catalyst is in a different phase / state. **[1]**

7. This question is about copper(II) sulfate, $CuSO_4$, and sodium thiosulfate, $Na_2S_2O_3$.

(a) The enthalpy change of reaction, $\Delta_r H$, for converting anhydrous copper(II) sulfate to hydrated copper(II) sulfate is difficult to measure directly by experiment.

$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4{\bullet}5H_2O(s)$ $\Delta_r H$ **reaction 5.1**

The enthalpy changes of solution of anhydrous and hydrated copper(II) sulfate can be measured by experiment. The reactions are shown below.

In the equations, 'aq' represents an excess of water.

$CuSO_4(s) + aq \rightarrow Cu^{2+}(aq) + SO_4^{2-}(aq)$ $\Delta_{sol}H(CuSO_4(s))$ **reaction 5.2**

$CuSO_4{\bullet}5H_2O(s) + aq \rightarrow Cu^{2+}(aq) + SO_4^{2-}(aq)$ $\Delta_{sol}H(CuSO_4{\bullet}5H_2O(s))$ **reaction 5.3**

Experiment 1
A student carries out an experiment to find $\Delta_{sol}H(CuSO_4(s))$ for **reaction 5.2**.

Student's method
- Weigh a bottle containing $CuSO_4(s)$ and weigh a polystyrene cup.
- Add about 50 cm^3 of water to the polystyrene cup and measure its temperature.
- Add the $CuSO_4(s)$, stir the mixture, and measure the final temperature.
- Weigh the empty bottle and weigh the polystyrene cup with final solution.

Mass readings

Mass of bottle + $CuSO_4(s)$ / g	28.04
Mass of empty bottle / g	20.06
Mass of polystyrene cup / g	23.43
Mass of polystyrene cup + final solution / g	74.13

Temperature readings

Initial temperature of water / °C	20.5
Temperature of final solution / °C	34.0

Experiment 2
The student carries out a second experiment with $CuSO_4{\bullet}5H_2O$ (**reaction 5.3**). The student uses the same method as in **Experiment 1**.

The student calculates $\Delta_{sol}H(CuSO_4{\bullet}5H_2O(s))$ as +8.43 kJ mol^{-1}.

(i)* Calculate $\Delta_{sol}H(CuSO_4(s))$ for **reaction 5.2** and determine the enthalpy change of **reaction 5.1**, Δ_rH.

Assume that the specific heat capacity, *c*, of the solution is the same as for water.

Show your working, including an energy cycle linking the enthalpy changes. **[6]**

$E = mc\Delta T$

$74.13 - 23.43 = 50.7$ g (mass of solution)

$\Delta T = 34.0 - 20.5 = 13.5$ °C

$E = 50.7 \times 4.18 \times 13.5 = 2861$ J $= 2.861$ kJ

$28.04 - 20.06 = 7.98$ g

$$\frac{7.98}{(63.5+32+(16\times4))} = 0.05 \text{ mol of } CuSO_4$$

$$\frac{2.861}{0.05} = -57.22 \text{ kJmol}^{-1} \; \Delta_{sol}H(CuSO_{4}(s))$$

$-57.22 - 8.43 = -65.65$ kJmol^{-1}

$= \Delta_rH$

Additional answer space if required

(ii) The thermometer had an uncertainty in each temperature reading of ±0.1 °C.

used ×2 in temp. change

The student calculates a 20% uncertainty in the temperature change in **Experiment 2**.

Calculate the temperature change in **Experiment 2**.

$$(0.1 \times 2) \times \frac{100}{20} = 1.0°C$$

temperature change = 1.0°C °C **[1]**

(b) The standard enthalpy change of reaction, $\Delta_r H^\ominus$, and the standard free energy change, $\Delta G^\ominus$, for converting anhydrous sodium thiosulfate to hydrated sodium thiosulfate are shown below.

$Na_2S_2O_3(s) + 5H_2O(l) \rightarrow Na_2S_2O_3{\bullet}5H_2O(s)$ $\Delta_r H^\ominus = -55.8\,kJ\,mol^{-1}$

$\Delta G^\ominus = -16.1\,kJ\,mol^{-1}$

Standard entropies are given in the table.

Compound	**$S^\ominus / J\,K^{-1}\,mol^{-1}$**
$Na_2S_2O_3{\bullet}5H_2O(s)$	372.4
$H_2O(l)$	69.9

Determine the **standard** entropy, $S^\ominus$, of anhydrous sodium thiosulfate, $Na_2S_2O_3(s)$.

Give your answer to **3** significant figures.

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

$-16.1 = -55.8 - 298(\Delta S)$

$39.7 = -298(\Delta S)$

$-0.133 = \Delta S = -133\ JK^{-1}mol^{-1}$

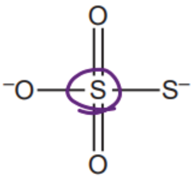
$kJK^{-1}mol^{-1}$

$\Delta S = 372.4 - (5 \times 69.9) = 22.9\ JK^{-1}mol^{-1}$

$\Delta S = 22.9 - (-133) = 156\ JK^{-1}mol^{-1}$

$S^\ominus$ = 156 $JK^{-1}mol^{-1}$ **[4]**

(c) Sodium thiosulfate contains the thiosulfate ion, $S_2O_3^{2-}$.
The displayed formula of $S_2O_3^{2-}$ can be shown as below.



4bp : 0 lp

thiosulfate ion

(i) Predict the O–S–S bond angle and name of the shape of the thiosulfate ion.

Bond angle 109.5°

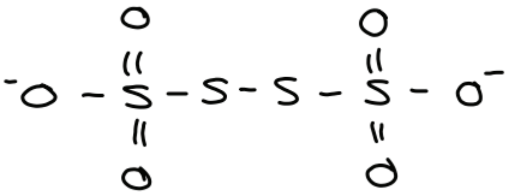
Name of shape tetrahedral

[1]

(ii) In some of its reactions, the thiosulfate ion forms the tetrathionate ion, $S_4O_6^{2-}$.

The $S_4O_6^{2-}$ ion is a 'dimer' of $S_2O_3^{2-}$.

Draw a displayed formula for the $S_4O_6^{2-}$ ion.



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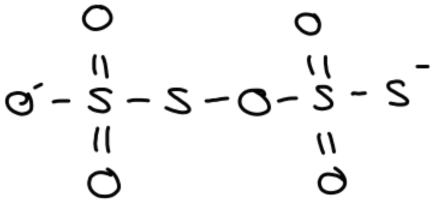
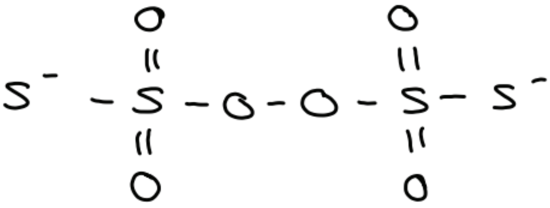


[1]



8. Methanol, CH_3OH, can be made industrially by the reaction of carbon monoxide with hydrogen, as shown in **equilibrium 1**.

$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$ $\Delta H = -91\ kJ\ mol^{-1}$ **Equilibrium 1**

(a) Predict the conditions of pressure and temperature that would give the maximum equilibrium yield of CH_3OH in **equilibrium 1**.

want to favour forwards reaction

Explain your answer.

Right hand side has fewer gaseous moles

So high pressure

Forwards reaction was exothermic so low

temperature

[3]

(b) A catalyst is used in the production of methanol in **equilibrium 1**.

State **two** ways that the use of catalysts helps chemical companies to make their processes more sustainable and less harmful to the environment.

1 lower energy demand

2 less CO_2 emissions

[2]

(c) Standard entropy values are given below.

Substance	CO(g)	H_2(g)	CH_3OH(g)
$S^\circ / J K^{-1} mol^{-1}$	198	131	238

A chemist proposed producing methanol at 525 K using **equilibrium 1**.

Explain, with a calculation, whether the production of methanol is feasible at 525 K.

$\Delta S = 238 - (198 + (2 \times 131)) = -222 \ JK^{-1} mol^{-1}$

$= -0.222 \ kJK^{-1} mol^{-1}$

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

$\Delta G = -91 - (525 \times -0.222) = 25.55$

$kJmol^{-1}$

not feasible as $\Delta G > 0$

[5]

(d) At 298 K, the free energy change, Δ*G*, for the production of methanol in **equilibrium 1** is $-2.48 \times 10^4 \ J mol^{-1}$.

Δ*G* is linked to K_p by the relationship: $\Delta G = -RT \ln K_p$.

R = gas constant
T = temperature in K.

Calculate K_p for **equilibrium 1** at 298 K.

Give your answer to **3** significant figures.

$\ln kp = \frac{\Delta G}{-RT}$

$\ln kp = \frac{-2.48 \times 10^4}{-8.314 \times 298}$

$= 10.01$

$Kp = e^{-10.01} = 2.22 \times 10^4$

$kp = \frac{p(CH_3OH)}{p(CO) \times p(H_2)^2}$

K_p = 2.22×10^4 units atm^{-2} **[3]**