MEGA LECTURE

A LEVEL CHEMISTRY

TOPIC 10 –THERMODYNAMICS

TEST

Answer all questions

Max 50 marks

Name ……
Mark ……../50 ……..% Grade ……

1.

2. Calcium fluoride occurs naturally as the mineral fluorite, a very hard crystalline solid that is almost insoluble in water and is used as a gemstone.

Tables 1 and **2** contain thermodynamic data.

Table 1

Process	**$\Delta H^{\ominus}$ / kJ mol^{-1}**
$Ca(s) \rightarrow Ca(g)$	+193
$Ca(g) \rightarrow Ca^{+}(g) + e^{-}$	+590
$Ca^{+}(g) \rightarrow Ca^{2+}(g) + e^{-}$	+1150
$F_2(g) \rightarrow 2F(g)$	+158
$F(g) + e^{-} \rightarrow F^{-}(g)$	–348

Table 2

Name of enthalpy change	**$\Delta H^{\ominus}$ / kJ mol^{-1}**
Enthalpy of lattice dissociation for calcium fluoride	+2602
Enthalpy of lattice dissociation for calcium chloride	+2237
Enthalpy of hydration for F^- ions	–506
Enthalpy of hydration for Cl^- ions	–364
Enthalpy of hydration for Ca^{2+} ions	–1650

(a) Write an equation, including state symbols, for the process that occurs when the calcium fluoride lattice dissociates and for which the enthalpy change is equal to the lattice enthalpy.

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(1)

(b) (i) Define the term *standard enthalpy of formation*.

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(ii) Write an equation, including state symbols, for the process that has an enthalpy change equal to the standard enthalpy of formation of calcium fluoride.

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(iii) Use data from the **Tables 1** and **2** to calculate the standard enthalpy of formation for calcium fluoride.

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(c) Explain why the enthalpy of lattice dissociation for calcium fluoride is greater than that for calcium chloride.

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(d) Calcium chloride dissolves in water. After a certain amount has dissolved, a saturated solution is formed and the following equilibrium is established.

$$CaCl_2(s) \rightleftharpoons Ca^{2+}(aq) + 2Cl^-(aq)$$

(i) Using data from **Table 2**, calculate the enthalpy change for this reaction.

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(ii) Predict whether raising the temperature will increase, decrease or have no effect on the amount of solid calcium chloride that can dissolve in a fixed mass of water.
Explain your prediction.
(If you have been unable to obtain an answer to part (d) (i), you may assume that the enthalpy change = –60 kJ mol^{-1}. This is **not** the correct answer.)

Effect on amount of solid that can dissolve .....

Explanation .....
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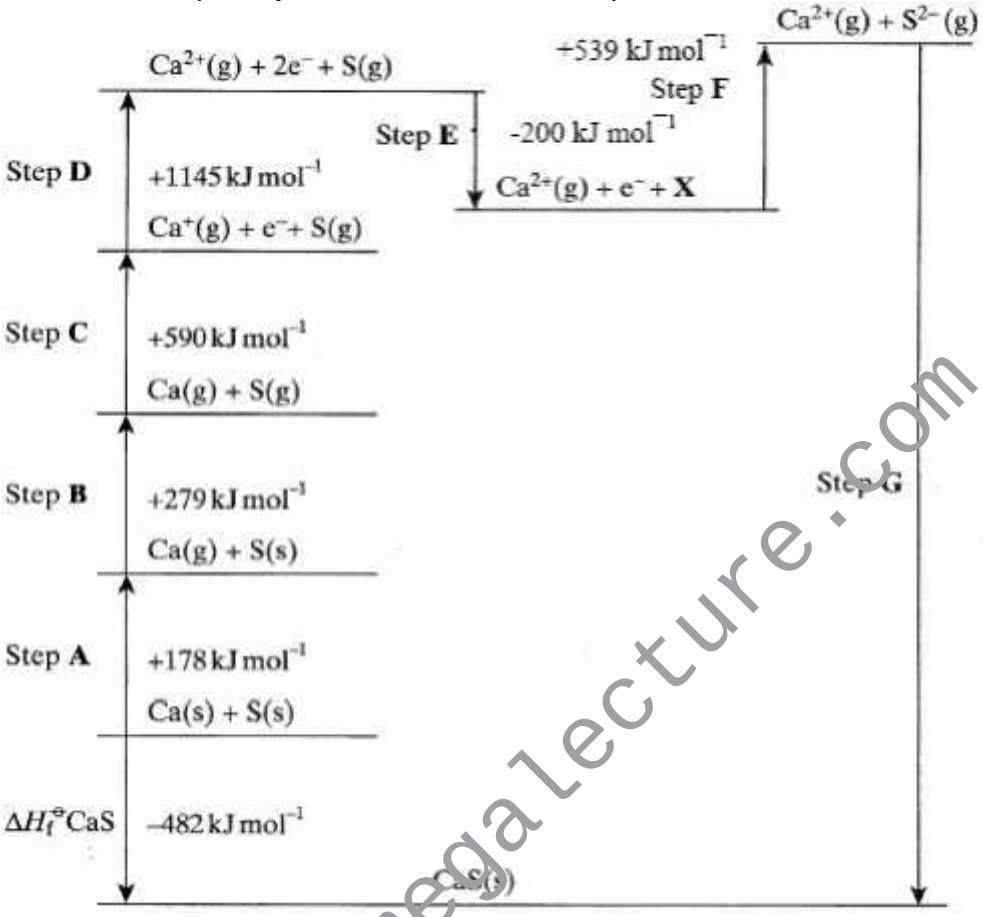
(Total 15 marks)

2. (a) A Born–Haber cycle for the formation of calcium sulphide is shown below. The cycle includes enthalpy changes for all steps except step **G**. (The cycle is not drawn to scale.)



(i) Give the full electronic configuration of the ion S^{2-}

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(ii) Suggest why step **F** is an endothermic process.

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(iii) Name the enthalpy changes in steps **B** and **D**.

Step
B .....

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Step ***D*** .....
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(iv) Explain why the enthalpy change for step **D** is larger than that for step **C**.

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(v) Use the data shown in the cycle to calculate a value for the enthalpy change for step **G**.

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(b) Using a Born–Haber cycle, a value of –905 kJ mol^{-1} was determined for the lattice enthalpy of silver chloride. A value for the lattice enthalpy of silver chloride using the ionic model was –833 kJ mol^{-1}.

Explain what a scientist would be able to deduce from a comparison of these values.

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(c) Some endothermic reactions occur spontaneously at room temperature. Some exothermic reactions do not occur if the reactants are heated together to a very high temperature.

In order to explain the following observations, another factor, the entropy change, ΔS,
must be considered. The equation which relates ΔS to ΔH is given below.

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

(i) Explain why the following reaction occurs at room temperature even though the reaction is endothermic.

$$NaHCO_3(aq) + HCl(aq) \quad NaCl(aq) + H_2O(l) + CO_2(g)$$

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(ii) Explain why the following reaction does not occur at very high temperatures even though the reaction is exothermic.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

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(Total 18 marks)

3. Methanol can be regarded as a carbon-neutral fuel because it can be synthesised from carbon dioxide as shown in the equation below.

$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$$

Standard enthalpy of formation and standard entropy data for the starting materials and products are shown in the following table.

	$CO_2(g)$	$H_2(g)$	$CH_3OH(g)$	$H_2O(g)$
$\Delta H_f^\ominus$ / kJ mol^{-1}	–394	0	–201	–242
$S^\ominus$ / J K^{-1} mol^{-1}	214	131	238	189

(a) Calculate the standard enthalpy change for this reaction.

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(b) Calculate the standard entropy change for this reaction.

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(c) Use your answers to parts (a) and (b) to explain why this reaction is **not** feasible at high temperatures.

Calculate the temperature at which the reaction becomes feasible.

Suggest why the industrial process is carried out at a higher temperature than you have calculated.

(If you have been unable to calculate values for ΔH and ΔS you may assume that they are –61 kJ mol^{-1} and –205 J K^{-1} mol^{-1} respectively. These are **not** the correct values.)

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(d) Write an equation for the complete combustion of methanol. Use your

equation to explain why the combustion reaction in the gas phase is feasible at all temperatures.

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(4)
(Total 16 marks)

4. Which one of the following statements is **not** correct?

A The first ionisation energy of iron is greater than its second ionisation energy.

B The magnitude of the lattice enthalpy of magnesium oxide is greater than that of barium oxide.

C The oxidation state of iron in $[Fe(CN)_6]^{3-}$ is greater than the oxidation state of copper in $[CuCl_2]$

D The boiling point of C_3H_8 is lower than that of CH_3CH_2OH

(Total 1 mark)