## **Stage 2 Chemistry**

## **Birdwood**

HIGH SCHOOL **Topic 3: Using and Controlling Reactions**

**Chemical equilibrium: calculations and Le Chatelier**

**Review Paper 16**

**DUE DATE:** Ref: ESSENTIALS pages 170 179

**Question 1**

Hexadecane C16H34 is a constituent of diesel. Hexadecane has a density of 0.775 g mL-1 and an enthalpy of combustion of 33400 kJ L-1.

1. Calculate the energy density of hexadecane in kJ g-1.

0.775 g/mL = 775 g/L

33400/775 = 43 kJ/g

1. Hence calculate the energy obtained from the complete combustion of 1 kg of hexadecane.

43 x 1000 = 43 000 kJ

c) Write the thermochemical equation for the complete combustion of hexadecane. (8 marks)

C16H34(l) + O2(g) -> CO2(g) + H2O(g) ΔH = -43000kJ

**Question 2**

Hydrogen cyanide, HCN, can be made by reacting carbon monoxide and ammonia, in a closed system as shown by the reaction:

2CO(g) + NH3(g) HCN(g) + CO2(g) + H2(g)

The system was allowed to reach equilibrium at 8000C and analysis found that the concentrations of each substance at equilibrium was;

[CO(g)] = 0.050 M, [NH3(g)] = 0.025 M, [HCN(g)] = 0.084 M, [CO2(g)] = 0.084 M, [H2(g)] = 0.084 M

1. i State *two* essential conditions for a chemical system to establish equilibrium.

Closed system

Constant temperature

ii If the masses of reactants and products in this type of system are held inside the vessel, when

equilibrium is reached, what can be said about the *concentration of reactants and products* at

equilibrium?

Concentration of reactants and products remain constant (note: Creactant ≠ Cproduct)

iii State one more characteristic of a system at equilibrium.

All observable quantities remain constant

1. Explain why the temperature must be stated when quoting an equilibrium constant?

Equilibrium position will change at different temperatures, ie Kc is different for each temperature.

1. Write the Kc expression for the above reaction.
2. Hence calculate the value of the equilibrium constant, Kc, at 8000C.

Kc = 9.48

1. Using what you found in part d), state the value of Kc for the reaction below:

HCN(g) + CO2(g) + H2(g) 2CO(g) + NH3(g)

Assuming the same temp.

Kc = 1/9.48

Kc = 0.105 (9 marks)

**Question 3**

In the commercial production of ammonia by the Haber process, the equilibrium reaction is represented by:

N2(g) + 3H2(g) 2NH3(g) ΔH = -92 kJmol-1

1. Calculate the energy released when 500 kg of nitrogen completely reacts with an excess of hydrogen.

n = m/M

n = 500 000 / 28.02

n = 17 844.4 mol

ΔH = 17 844.4 x 92

ΔH = 1 641 684 kJ

1. Explain using Le Chatelier’s Principle, *the effect on the position of equilibrium* by each of the following conditions:

i A relatively low temperature is used.

According to Le Chat low temperature will favour the exothermic reaction.

Equilibrium will be reestablished to fovour the forwards reaction

More ammonia will be formed

ii A high pressure of 500 atmospheres is used.

High pressure will favour the reaction producing fewer molecules

Forwards reaction favoured, more ammonia produced

iii Ammonia is removed as soon as it forms, by liquefying it. (6 marks)

As the products are being removed the backwards reaction will not occur, therefore the forwards reaction is favoured and more ammonia is produced

**Question 4**

A gaseous mixture of nitrogen dioxide and dinitrogen tetroxide, N2O4, are allowed to come to equilibrium at 300C in a 2.0 L container. The reaction is:

2NO2(g) N2O4(g)

The initial amounts of both gases was 0.50 mole.

At equilibrium, the amount of dinitrogen tetroxide is 0.7188 mole.

This information can be represented in the table below:

|  |  |  |
| --- | --- | --- |
|  | **NO2(g)** | **N2O4(g)** |
| initial no. of moles | 0.50 | 0.50 |
| moles used/formed | -0.4376 | +0.2188 |
| moles at equilibrium | ***0.0624*** | 0.7188 |
| concentration at equilibrium | 0.0312 | 0.3594 |

1. Use the table to find the value of ***x***, the number of moles of nitrogen dioxide, *at equilibrium*.
2. Complete the last row of the table and hence calculate the value of Kc at 300C.

Kc = 369

1. Sketch a graph to show the *concentrations* of each gas, from time = 0 until, at t = 10 min when equilibrium is established and continue your graphs to show what happens to the concentrations until

t = 15 min. *Show clearly on your graph where equilibrium was reached*.

1. At t = 15 min, there is a sudden injection of nitrogen dioxide. Extend your graphs of concentration

vs time in part c) to show what will happen as the system re-establishes equilibrium.

(9 marks)

0.5

15

10

**TOTAL MARK = 32**