## **Stage 2 Chemistry**

## **Birdwood**

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

**Chemical Industry and Metal Production**

**Review Paper 17**

**DUE DATE:** Ref: ESSENTIALS pages 179 - 196

**Question 1**

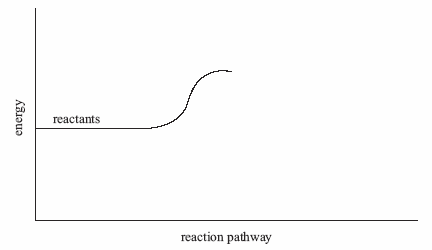
The conversion of nitrogen and hydrogen to ammonia in the Haber Process is represented by the following equation:

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

A catalyst and a relatively low temperature are used in this reaction to give the best yield.

a) State the *purpose* of the catalyst in this reaction.

lowers activation energy to speed up the reaction

b) The beginning of the energy profile

No catalyst

diagram for the reaction is shown opposite.

Complete the energy profile diagram,

clearly identifying the ΔH and the

Activation energy

activation energy.

c) On the energy profile diagram, draw the

ΔH

reaction pathway of this reaction

without the catalyst.

d) A compromise between yield and rate of reaction often determines the optimum conditions for an industrial process.

i State and explain the effect of low temperature on the yield of ammonia.

Low temperature favours exothermic reaction, according to Le Chatelier. Would increase the yield of ammonia

ii In terms of Collision Theory, state and explain the effect low temperature has on the rate of this

reaction.

Low temp reduces energy of the particles, leading to fewer collisions due to less movement and fewer successful collisions due to not having the required activation energy.

The table below gives the percentages of ammonia in the equilibrium mixture at various pressures.

|  |  |
| --- | --- |
| **Pressure (atmospheres)** | **Yield of ammonia at 5500C with a catalyst** |
| 1  100  1000 | negligible  7%  41% |

In industry, a pressure of 250 atmospheres is used in the above equilibrium.

e) State one reason why a pressure of 250 atmospheres is used and not 1000 atmospheres.

Very costly to maintain a pressure of 1000 atm.

In one experiment conducted at 5500C, the concentrations of all three gases in the system were determined. The results are shown in the table below:

|  |  |
| --- | --- |
| **Gas** | **Concentration**  **(mol L-1)** |
| N2  H2  NH3 | 0.5  0.3  0.3 |

f) Write a Kc expression for this reaction.

[NH3]2 / [H2]3[N2]

g) The value for Kc for this reaction at 5500C is **0.35**.

Using your Kc expression, *calculate the value of Kc*, using the values in the above table, to explain why the

system is not at equilibrium.

Kc = 6.7, which is not 0.35, therefore the reaction is not at equilibrium

(16 marks)

**Question 2**

In the production of zinc, the zinc sulfide ore is roasted in air to form zinc oxide.

Sulfur dioxide gas is the other product, which is used in the manufacture of sulfuric acid. The sulfuric acid is then reacted with zinc oxide in a leaching process to produce a solution of zinc sulfate.

The zinc sulfate solution must be purified before electrolysis.

a) Write an equation to represent the roasting stage.

ZnS + O2 ->ZnO + SO2

*The next stage is the leaching of zinc oxide with sulfuric acid.*

b) Write an equation to represent this stage.

ZnO + H2SO4 -> ZnSO4 + H2O

*Zinc sulfate is the electrolysed to produce zinc metal.*

c) Write a half-equation to represent the formation of zinc.

Zn2+ + 2e- -> Zn

d) Does this half-equation represent oxidation or reduction?

reduction

e) Will this occur at the positive or negative electrode of the electrolysis unit?

negative

*The purification process is essential in the production of zinc.*

f) Describe the process that purifies the Zn2+(aq).

Add Zn metal to the aqueous salts. Zn will replace any less active metal in solution.

g) State how the final product (zinc) after electrolysis, would be different if purification is not done.

Least active metal is reduced, you would have a mix of copper, silver, iron etc with the zinc

*Environmentalists and conservationists will often view mining and metal production in a critical way.*

h) Describe *one harmful effect* zinc mining or zinc metal production may have on the environment and

*one method to reduce this effect.* Mining is always an issue, plus the production of SO2 and H2SO4

(12 marks)

**Question 3**

Magnesium is produced by a series of processes as shown in the flowchart below:

CaCO3(S)

heater

CaO(s)

slaker

H2O(l)

Ca(OH)2(aq)

precipitator

sea water

Mg(OH)2(s)

neutraliser

HCl

MgCl2(aq)

acid

plant

solar evaporators

driers

MgCl2(s)

electrolysis unit

Cl2

**Magnesium**

a) Name the *three* raw materials used in this process. CaCO3, sea water, H2O

b) Name the *by-product* obtained from the electrolysis unit. Cl2

c) What use is made of the by-product and how might this increase the profitability of the company? make HCl

d) From the flowchart, identify *two* stages at which large amounts of energy are needed. heater, electrolysis

e) Complete the chemical equation for the reaction occurring in the precipitator:

Ca(OH)2(aq) + Mg2+(aq) Mg(OH)2 + Ca2+

f) Name the other substance produced when calcium carbonate is converted into calcium oxide in the heater. Carbon dioxide

g) Why might this part of the process be of concern to environmentalists? CO2 leads to climate change

In the electrolysis unit, *molten* magnesium chloride is electrolysed and magnesium is produced at one of the electrodes.

h) Explain why all of the water must be removed from the magnesium chloride before it is electrolysed. Otherwise water may be reduced instead (water is less reactive than mg)

i) Write a half equation for the electrode reaction in which magnesium is produced. Mg2+ + 2e- -> Mg

j) Name the electrode at which the magnesium is produced. Cathode

(16 marks)

**Question 4**

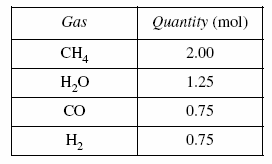
a) What changes will always shift this equilibrium reaction to the right?

2HI(*g*) → H2(*g*) + I2(*g*) Δ*H* = –52 kJ

lower temp

increase amount of HI

remove H2 or I2

b) Consider the following mixture of gases in a closed 5.0 L vessel at 730°C.

The following reaction occurs:

CH4(*g*) +H2O(*g*) → CO(*g*) +3H2(*g*) ∆*H*  = -206 kJ

The equilibrium constant, K, is 0.26 at 730°C.

(i) Determine whether the system is at equilibrium.

(0.15)(0.15)3 / (0.4)(0.25) = 0.00506

not at equilibrium

(ii) Explain how conditions in this reaction could be adjusted to increase the quantity of products.

decrease temp

decrease pressure

increase amount of CH4 or H2O

remove CO or H2

because le chat

(9 marks)

**TOTAL MARK = 53**