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CHEMISTRY

PHYSICAL CHEMISTRY

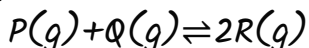
Level & Board	AQA (A-LEVEL)
TOPIC:	CHEMICAL EQUILIBRIA
PAPER TYPE:	SOLUTION -1
TOTAL QUESTIONS	10
TOTAL MARKS	61

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Chemical Equilibria - 2

1.

(a)



According to the balanced equation, for every 1 mol of P or Q that reacts, it produces 2 mol of R. So, if x mol of P reacts, it generates 2x mol of R. So,

At equilibrium: Amount of R (in terms of x) = 2x mol

(1)

(b)

Given the equilibrium constant $K_c = 3.6$ for the reaction:

At equilibrium, let's assume the concentration of R formed is 2x (as previously established) and the concentrations of P and Q that react are (1-x) each.

Using the K_c expression:

$$K_c = \frac{[R]^2}{[P][Q]}$$

$$3.6 = \frac{[2x]^2}{[(1-x)][(1-x)]}$$

$$3.6 = \frac{[2x]^2}{(1-x)^2}$$

$$\sqrt{3.6} = \frac{2x}{1-x}$$

$$\sqrt{3.6} - \sqrt{3.6}x = 2x$$

$$1.9 = 3.9x$$

$$x = 0.49$$

So,

$$[R] = 0.97 \text{ mol dm}^{-3}$$

(3)

2.

The conditions for the industrial synthesis of hydrogen from methane and steam $CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$ at 800°C and low pressure of 300 kPa in the presence of a nickel catalyst are primarily based on equilibrium yield and cost considerations:

Equilibrium Yield:

At higher temperatures, the endothermic reaction ($\Delta H = +206 \text{ kJ mol}^{-1}$) shifts towards the right side, favoring the production of hydrogen (H_2). Increasing the temperature helps counteract the endothermic nature of the reaction, facilitating higher yields of hydrogen. So, 800°C strikes a balance between favorable equilibrium yield and minimizing side reactions.

Cost:

Operating at lower pressures reduces costs associated with the construction and maintenance of high-pressure equipment. Although higher pressures would favor the side with fewer gas molecules (Le Chatelier's principle), increasing the pressure also increases operational expenses.

Therefore, a lower pressure of 300 kPa is chosen to maintain a balance between achieving a reasonable yield of hydrogen and cost-effectiveness.

Catalyst Effect:

The use of a nickel catalyst enables the reaction to proceed at a reasonable rate while lowering the activation energy required for the reaction to occur. This facilitates a higher rate of reaction and helps achieve the desired equilibrium yield of hydrogen without excessively raising the operational temperature or pressure, thereby contributing to cost-effectiveness.

The chosen conditions of temperature, pressure, and catalyst usage aim to maximize the equilibrium yield of hydrogen while considering the practicality and cost-effectiveness of the industrial process. These conditions give a balance between obtaining a high yield of hydrogen and minimizing production costs, ensuring an efficient and economically viable hydrogen production process.

(Total 6 marks)

3. (c)

(Total 1 mark)

4. (a)

A decrease in overall pressure shifts the equilibrium towards the side with more gas molecules to counteract the pressure change. In this reaction, it shifts towards the reactants (SO_2 and O_2) to oppose the pressure decrease, decreasing the yield of SO_3 .

(3)

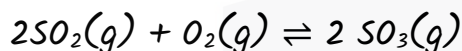
(b)

A 0.460 mol sample of SO_2 is mixed with a 0.250 mol sample of O_2 in a sealed container at a constant temperature.

When equilibrium is reached at a pressure of 215 kPa, the mixture contains 0.180 mol of SO_3

Calculate the partial pressure, in kPa, of SO_2 in this equilibrium mixture.

The balanced chemical equation for the formation of SO_3 from SO_2 and O_2 is:



Initial moles:

- $\text{SO}_2 = 0.460 \text{ mol}$
- $\text{O}_2 = 0.250 \text{ mol}$

Change at equilibrium (based on the stoichiometry of the balanced equation):

- $\text{SO}_3 \text{ formed} = 0.180 \text{ mol}$

According to the balanced equation, 2 moles of SO_2 produce 2 moles of SO_3 . Therefore, the moles of SO_2 consumed will be the same as the moles of SO_3 formed, which is 0.180 mol.

Amount of SO_2 at equilibrium

$$\begin{aligned} &= \text{Initial moles of } \text{SO}_2 - \text{Moles of } \text{SO}_3 \text{ formed} \\ &= 0.46 \text{ mol} - 0.18 \text{ mol} \\ &= 0.28 \text{ mol.} \end{aligned}$$

Amount of O_2 at equilibrium

$$\begin{aligned} &= \text{Initial moles of } \text{O}_2 - \text{Moles of } \text{O}_2 \text{ used} \\ &= 0.25 \text{ mol} - (0.5 \times \text{Moles of } \text{SO}_3 \text{ formed}) \\ &= 0.25 \text{ mol} - 0.09 \text{ mol} \\ &= 0.16 \text{ mol.} \end{aligned}$$

Total amount of gas at equilibrium

$$\begin{aligned} &= \text{Moles of } \text{SO}_2 + \text{Moles of } \text{O}_2 + \text{Moles of } \text{SO}_3 \\ &= 0.28 \text{ mol} + 0.16 \text{ mol} + 0.18 \text{ mol} \\ &= 0.62 \text{ mol.} \end{aligned}$$

Partial pressure of $\text{SO}_2 = \text{Total pressure} \times (\text{Moles of } \text{SO}_2 / \text{Total moles of gas})$

$$\begin{aligned} &= 215 \text{ kPa} \times (0.28 \text{ mol} / 0.62 \text{ mol}) \\ &= 215 \text{ kPa} \times 0.4516 \\ &= 97.1 \text{ kPa.} \end{aligned}$$

(4)

(c)

$$K_p = \frac{(p_{\text{SO}_3})^3}{(p_{\text{O}_2})(p_{\text{SO}_2})^2}$$

(1)

(d) (c)

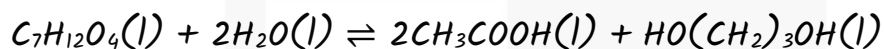
(1)

5. (c)

(Total 1 mark)

6. (a)

Using balance chemical equation:



Amount in the mixture / mol				
	Diester	Water	Acid	Diol
At the start	1	1	0	0
At equilibrium	1-x/2	1-x	x	x/2

(3)

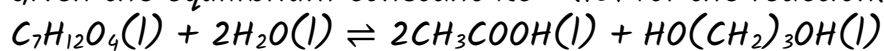
(b)

Deduce the structure of the diester :



(1)

(c)

Given the equilibrium constant $K_c=0.161$ for the reaction:

Given the amounts of substances at equilibrium:

- Amount of diester ($C_7H_{12}O_4$) = 0.971 mol
 - Amount of acid (CH_3COOH) = 0.452 mol
 - Amount of diol ($HO(CH_2)_3OH$) = 0.273 mol
- The amount of water (H_2O) at equilibrium
Using the equilibrium constant expression:

$$K_c = \frac{[HO(CH_2)_3OH][CH_3COOH]^2}{[C_7H_{12}O_4] \cdot [H_2O]^2}$$

Now, put in the given values:

$$[H_2O]^2 = \frac{(0.273)(0.452)^2}{(0.971)(0.161)}$$

$$[H_2O]^2 = 0.357$$

Taking the square root of both sides to solve for $[H_2O]$:

$$[H_2O] = 0.598$$

7. (A)

(Total 1 mark)

8.
(a)

The nearest minute, when equilibrium is first established is 3 minutes as At equilibrium, $rate_{forward} = rate_{back}$ so concentrations (of O_2 and SO_3) remain constant

CHEMISTRY ONLINE (2)

(b)



The higher temperature (T_2) caused the equilibrium to shift to the right (towards the product side) in an endothermic direction, leading to a decrease in the concentration of sulfur trioxide (SO_3) in the new equilibrium.

(2)

9.

(a)

A catalyst has no effect on the yield of methanol in this equilibrium as it increases the rate of both forward and reverse reactions equally without altering the final amount of methanol produced.

(2)

(b)

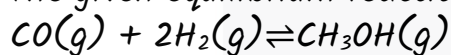


$$K_c = \frac{[\text{CH}_3\text{OH}]}{[\text{CO}][\text{H}_2]^2}$$

(1)

(c)

The given equilibrium reaction is:



At equilibrium:

- Moles of CO = 0.340 mol
- Moles of H₂ = 0.190 mol
- Moles of CH₃OH = 0.0610 mol

$$[\text{CO}] = 0.340 \text{ mol} / 0.250 \text{ dm}^3$$

$$= 1.36 \text{ mol dm}^{-3}$$

$$[\text{H}_2] = 0.190 \text{ mol} / 0.250 \text{ dm}^3$$

$$= 0.76 \text{ mol dm}^{-3}$$

$$[\text{CH}_3\text{OH}] = 0.0610 \text{ mol} / 0.250 \text{ dm}^3$$

$$= 0.244 \text{ mol dm}^{-3}$$

$$K_c = \frac{[\text{CH}_3\text{OH}]}{[\text{CO}][\text{H}_2]^2}$$

$$K_c = \frac{0.244}{(1.36)(0.76)^2}$$

$$= 0.311 \text{ mol}^2 \text{ dm}^{-6}$$

(3)

(d)

$$K_{c\text{reverse}} = 1 / K_{c\text{forward}}$$

$$K_{c\text{reverse}} = 1 / 0.311$$

$$K_{c\text{reverse}} = 3.22 \text{ mol}^2 \text{ dm}^{-6}$$

(2)

10.

(a)

Given:

Initial amount of A = 6.4×10^{-3} molAmount of A at equilibrium = 3.9×10^{-3} mol

Using the given expression for the equilibrium amount of A:

The balanced equation is $2A + B \rightleftharpoons 3C + D$, which means that 2 moles of A react with 1 mole of B to form 3 moles of C and 1 mole of D.

$$K_c = \frac{[D][C]^3}{[B][A]^2}$$

Equilibrium amount of A = $6.4 \times 10^{-3} - 2x$

$$3.9 \times 10^{-3} = 6.4 \times 10^{-3} - 2x$$

$$2x = 6.4 \times 10^{-3} - 3.9 \times 10^{-3}$$

$$2x = 2.5 \times 10^{-3}$$

$$x = 1.25 \times 10^{-3}$$

Now, let's calculate the amounts of substances at equilibrium:

- Moles of B (x) = $9.5 \times 10^{-3} - x$
 $= 9.5 \times 10^{-3} - 1.25 \times 10^{-3}$
 $= 8.25 \times 10^{-3}$ mol
- Moles of C (using 3x) = $2.8 \times 10^{-2} + 3x$
 $= 2.8 \times 10^{-2} + 3 \times 1.25 \times 10^{-3}$
 $= 0.028 + 0.00375$
 $= 0.03175$ mol
- Moles of D (x) = 1.25×10^{-3}
 $x = 1.25 \times 10^{-3}$ mol

(5)

(b)

(K_c) for this equilibrium and its units as

$$K_c = \frac{[D][C]^3}{[B][A]^2}$$

$$= \text{mol dm}^{-3}$$

(2)

(c)

The equilibrium reaction is: $2A + B \rightleftharpoons 3C + D$

Given:

Moles of B at equilibrium = 0.21 mol

Moles of C at equilibrium = 1.05 mol
 Moles of D at equilibrium = 0.076 mol
 Total volume of solution = $5.00 \times 10^2 \text{ cm}^3$
 $= 5.00 \times 10^{-1} \text{ dm}^3$

Equilibrium constant (K_c) = 116

Let's denote the moles of A as x.

Moles of A = 2x

Now, let's construct an expression for K_c using the given concentrations:

$$K_c = \frac{[D][C]^3}{[B][A]^2}$$

$$K_c = 116$$

$$[B] = 0.21 / 0.5 \text{ mol dm}^{-3}$$

$$[C] = 1.05 / 0.5 \text{ mol dm}^{-3}$$

$$[D] = 0.076 / 0.5 \text{ mol dm}^{-3}$$

$$116 = \frac{(0.076/0.5)(1.05/0.5)^3}{(0.21/0.5)(2x)^2}$$

$$4x^2 = \frac{(0.076/0.5)(1.05/0.5)^3}{116(0.21/0.5)}$$

$$4x^2 = \frac{0.152 \times 74.088}{116 \times 0.42}$$

$$4x^2 = \frac{11.26}{48.72}$$

$$x^2 = \frac{11.26}{194.88}$$

$$x = 0.21$$

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(3)

(d)

Adding more water to the equilibrium mixture lowers all substance concentrations.

As $K_c = \text{mole ratio} \times 1/V$

According to Le Chatelier's Principle, the equilibrium shifts to counteract this by favoring the side with more moles. Since water addition increases volume, the equilibrium moves to decrease mole ratios, causing the amount of A to decrease.



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- Founder & CEO of Chemistry Online Tuition Ltd.
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- Chemistry, Physics, Math's and Biology Tutor

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