2024 Y5 CHEMISTRY REVISION LECTURE

Term 4 Week 1

Discussion Questions

1 (a) A sample of chlorodifluoromethane, CHC*l*F₂, undergoes pyrolysis in a closed vessel of volume 18.5 dm³ according to the following equilibrium.

 $2CHC lF_2(g) \Rightarrow C_2F_4(g) + 2HC l(g) \qquad \Delta H = +128 \text{ kJ mol}^{-1}$

At the reaction temperature, the degree of dissociation was found to be 75%.

- (i) Write the expression for the equilibrium constant, K_c , for the reaction. [1]
- (ii) Calculate the *initial* amount of $CHClF_2$ in the sample given that the magnitude of K_c at the reaction temperature was 0.780. [2]
- (iii) Explain how the K_c of the reaction will change when the reaction is carried out at a lower temperature. [2]

2	(a)	Nitrogen monoxide in the air can be converted to nitric acid, which results in acid rain. Both nitrogen monoxide and nitrogen dioxide also cause ozone layer depletion. One way of forming nitrogen monoxide is through the dissociation of nitrogen dioxide.				
			$2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g)$ $\Delta H = +114.2 \text{ kJ mol}^{-1}$			
		At 49	P4 °C, the value of K_p for the above reaction is 36.9 kPa.			
	When nitrogen dioxide is placed in a sealed containe 494 °C, equilibrium is reached when 40% of the original nitrogen dioxide has diss					
		(i)	Given that the standard enthalpy change of formation of NO(g) is +90.3 kJ mol ⁻¹ , calculate the standard enthalpy change of formation of NO ₂ (g).			
			[1]			
		(ii)	Write an expression for the equilibrium constant, K_{p} , for the reaction. [1]			
	 (iii) Calculate the initial partial pressure of nitrogen dioxide and 		Calculatethe initial partial pressure of nitrogen dioxide and			
			• the total pressure at equilibrium. [3]			
		(iv)	State and explain the impact on the equilibrium yield of NO(g) and value of K_p if the process was now conducted at 400 °C. [2]			

(v)	The volume of the sealed container was instantaneously doubled at constant temperature, and the system allowed to reach equilibrium.Explain the effect this will have on the partial pressures of the individual gases.
	[3]
(vi)	Hence, predict how the value of K_{ρ} will change. [1]
	N2018/P2/2

MCQ

1 A gas **P** decomposes to two other gases, **Q** and **R**, according to the equation:

$$\mathbf{2P}(g) \rightleftharpoons \mathbf{3Q}(g) + \mathbf{R}(g)$$

The graph below represents the decomposition of 1.0 mole of \mathbf{P} in the presence of a catalyst at various temperatures.



Which of the following statements about the above system are correct?

- **1** The equilibrium constant, K_p , for the decomposition reaction increases with increasing temperature.
- **2** At T_1 K, the mole fraction of **P** is greater than that of **R** at equilibrium.
- **3** The decomposition reaction is spontaneous at all temperatures.

2 The graph below shows how the fraction of a substance, **X** represented by one of the following compounds in the equilibrium mixture shown below varies with temperature at pressures of **Y** Pa and **Z** Pa.



Identify **X** and the correct relative magnitudes of **Y** and **Z**.

	X	Pressure
Α	N_2	Z > Y
в	O ₂	Y > Z
С	H ₂ O	Y > Z
D	NH_3	Z > Y

Extra Practice Question

One of the uses of natural gas is the reforming of methane to produce pure hydrogen gas which is widely used in the Haber process to manufacture ammonia. This method is also being considered as a means to produce the hydrogen required by fuel cells.

 $CH_4(g) + H_2O(g) \implies 3H_2(g) + CO(g)$

In an industrial reformer, methane and steam are sealed into a vessel at 700 °C. The initial partial pressure of each gas in the vessel is 1 atm. When equilibrium is reached, 40 % of the methane was converted to carbon monoxide.

(a) Calculate the value of the equilibrium constant, K_{p} , for the reaction at 700 °C.

[2]

(b) The following data shows how the percentage of H₂ in the equilibrium mixture varies with temperature and the volumes of the reaction vessel in a methane reformer.

Temperature	Percentage of H ₂ in the equilibrium mixture			
/ °C	for vessel with volume			
	X m ³	Y m ³	Z m ³	
550	1.68	0.87	0.009	
650	2.99	1.54	0.016	
750	5.71	3.02	0.032	
850	11.9	6.70	0.077	

Use the data to

- I deduce whether the production of H₂ is an exothermic or endothermic process.
- II arrange, in increasing order, the volumes X, Y and Z.

Explain your answer.

[4]

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1 (a) A sample of chlorodifluoromethane, CHC*l*F₂, undergoes pyrolysis in a closed vessel of volume 18.5 dm³ according to the following equilibrium.

 $2CHC lF_2(g) \Rightarrow C_2F_4(g) + 2HC l(g) \qquad \Delta H = +128 \text{ kJ mol}^{-1}$

At the reaction temperature, the degree of dissociation was found to be 75%.

(i) Write the expression for the equilibrium constant, K_c , for the reaction. [1]

[1] $K_c = \frac{[C_2F_4][HCI]^2}{[CHCIF_2]^2}$

(ii) Calculate the *initial* amount of $CHClF_2$ in the sample given that the magnitude of K_c at the reaction temperature was 0.780. [2]

Let initial amount of	CHC/F ₂ be x					
	2CHC <i>l</i> F ₂ (g)	⇒	C ₂ F ₄ (g)	+	2HC <i>l</i> (g)	
Initial amt / mol	X		0		0	
Change amt / mo	ol –0.75x		+0.375x		+0.75x	
Equil amt / mol	0.25x		0.375x		0.75x	
$\mathbf{K}_{c} = \frac{[C_{2}F_{4}][HCl]^{2}}{[CHClF_{2}]^{2}}$ $(0.375x)(0.75x)^{2}$						
$= \frac{(\frac{18.5}{18.5})(\frac{18.5}{18.5})}{(\frac{0.25x}{18.5})^2}$	= 0.780					

[1] Working to find the 3 equilibrium amounts in ratio.

[1] x = 4.28 mol

(iii) Explain how the K_c of the reaction will change when the reaction is carried out at a lower temperature. [2]

When temperature decreases, by LCP, the position of equilibrium shifts

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left towards the exothermic reaction to produce heat. [1]

Concentration of the products C_2F_4 and HCl decreases and

concentration of the reactants CHC/F₂ increases. K_c will decrease. [1]

2	(a)	Nitro Both of for	Nitrogen monoxide in the air can be converted to nitric acid, which results in acid rain. Both nitrogen monoxide and nitrogen dioxide also cause ozone layer depletion. One way of forming nitrogen monoxide is through the dissociation of nitrogen dioxide.				
			$2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g)$ $\Delta H = +114.2 \text{ kJ mol}^{-1}$				
		At 49	94 °C, the value of K_p for the above reaction is 36.9 kPa.				
		Whe 494 °	n nitrogen dioxide is placed in a sealed container at ² C, equilibrium is reached when 40% of the original nitrogen dioxide has dissociated.				
		(i)	Given that the standard enthalpy change of formation of NO(g) is +90.3 kJ mol ⁻¹ , calculate the standard enthalpy change of formation of NO ₂ (g).				
			[1]				
		$\Delta H = \Delta H_f \text{ (products)} - \Delta H_f \text{ (reactants)}$					
			+114.2 = 2(90.3) + 0 – (2 × Δ H _f NO ₂)				
			$\Delta H_{\rm f} \rm NO_2 = \underline{+33.2 \ \rm kJ \ \rm mol^{-1}} [1]$				
		(ii)	Write an expression for the equilibrium constant, K_{p} , for the reaction. [1]				
			2				
		$K_p =$	$K_p = \frac{P_{NO}^2 P_{O_2}}{P_{NO_2}^2} $ [1]				

(iii)	Calculate the initial partia the total pressu 	I pressure of ure at equilibri	nitrogen dioxide um.	and	[3]
Let t	he initial pressure of NO	2 be <i>x</i> kPa. 2 NO₂ (g) 、	<u>2 NO (g)</u> -	+ O ₂ (g)	
Initia	al pressure (kPa)	x	0	0	
Cha	nge in Pressure (kPa)	-0.40 <i>x</i>	+0.40 <i>x</i>	+0.20 <i>x</i>	
Equi	librium pressure (kPa)	0.60 <i>x</i>	0.40 <i>x</i>	0.20x [1]	
		$K_{p} = \frac{(0.40x)}{(0.6x)}$	$\frac{x^2(0.20x)}{(0x)^2} = 36.9$		
		<i>x</i> = 415 kPa	(or 415125 Pa)	
Hen	ce, initial pressure of NC	0₂ = 415 kPa	[1]		
Equi	Equilibrium pressure = $0.60x + 0.40x + 0.20x$ Or (249075 + 166050 + 83025)				
= 1.2 <i>x</i>					
	= 1.2 x	415			
	= 498 k	(Pa [1]			

(iv)	State and explain the impact on the equilibrium yield of NO(g) and value of K_p if the process was now conducted at 400 °C. [2]
	$2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g) \qquad \Box \Delta H = +114.2 \text{ kJ mol}^{-1}$
	When temperature is decreased (to 400 °C), <u>backward exothermic</u> reaction is favoured, shifting the equilibrium position to the <u>left</u> to increase the temperature by <u>releasing heat</u> . [1]Equilibrium yield of NO will <u>decrease</u> .
	Value of K_p will <u>decrease</u> as <u>partial pressure of NO and O₂ decreased and</u> partial pressure of NO ₂ has increased. [1]

(v)	The volume of the sealed container was instantaneously doubled at constant temperature, and the system allowed to reach equilibrium.
	Explain the effect this will have on the partial pressures of the individual gases.
	[3]
	When the volume was doubled, the total_pressure was <u>decreased</u> (<u>halved</u>) or The <u>partial pressures of the 3 gases would also be halved/decreased</u> initially. [1]
	To increase the total pressure / increase the total number of gaseous particles in the system, by Le Chatelier's Principle, the position of equilibrium would shift to the right. [1]
	The partial pressure of NO ₂ would then <u>decrease</u> further_while that of NO and O ₂ would <u>increase</u> until equilibrium is re-established. [1]
(vi)	Hence, predict how the value of K_{ρ} will change. [1]
	N2018/P2/2
	K _p will not change. [1]
	[Total: 15]

Practice Question

3 One of the uses of natural gas is the reforming of methane to produce pure hydrogen gas which is widely used in the Haber process to manufacture ammonia. This method is also being considered as a means to produce the hydrogen required by fuel cells.

 $CH_4(g) + H_2O(g) \implies 3H_2(g) + CO(g)$

In an industrial reformer, methane and steam are sealed into a vessel at 700 °C. The initial partial pressure of each gas in the vessel is 1 atm. When equilibrium is reached, 40 % of the methane was converted to carbon monoxide.

(a) Calculate the value of the equilibrium constant, $K_{\rm p}$, for the reaction at 700 °C.

	CH ₄	⊦ H₂O	→ 3 H ₂	+ CO
Initial P/atm	1.0	1.0	0.0	0.0
Eqm P/atm	0.60	0.60	3(0.40	0) 0.40
[1]				
$K_{\rm p} = \frac{(1.2)}{(1.2)}$	$\frac{)^{3}(0.40)}{0.6)^{2}}$	= <u>1.92 atm</u>	² [1]	

(b) The following data shows how the percentage of H_2 in the equilibrium mixture varies with temperature and the volumes of the reaction vessel in a methane reformer.

Temperature	Percentage of H ₂ in the equilibrium mixture				
/ °C	for vessel with volume				
	X m ³	Y m ³	Z m ³		
550	1.68	0.87	0.009		
650	2.99	1.54	0.016		
750	5.71	3.02	0.032		
850	11.9	6.70	0.077		

Use the data to

I deduce whether the production of H₂ is an exothermic or endothermic process

II arrange, in increasing order, the volumes **X**, **Y** and **Z**. Explain your answer.

(b) When temperature increases (from 550°C to 850°C),

Percentage of H_2 increases (from 1.68% to 11.9%), thus <u>forward reaction is</u> favoured to absorb heat. [1]

By L.C.P, endothermic reaction is favoured. Hence, production of H_2 is <u>endothermic</u>. [1]

When volume changes from Z to X,

Percentage of H_2 increases (from 0.009% to 1.68%) thus forward reaction is favoured

When equilibrium shifts to the right, the number of moles of gas increases.

By L.C.P. pressure has decreased / volume has increased. [1]

Hence, increasing order: **Z** < **Y** < **X** [1]

MCQ

1 A gas P decomposes to two other gases, Q and R, according to the equation:

$$\mathbf{2P}(g) \rightleftharpoons \mathbf{3Q}(g) + \mathbf{R}(g)$$

The graph below represents the decomposition of 1.0 mole of \mathbf{P} in the presence of a catalyst at various temperatures.



Which of the following statements about the above system are correct?

- **1** The equilibrium constant, K_p , for the decomposition reaction increases with increasing temperature.
- **2** At T_1 K, the mole fraction of **P** is greater than that of **R** at equilibrium.
- **3** The decomposition reaction is spontaneous at all temperatures.

Gradient at t=0 gives the initial rate of the decomposition. Gradient at t=0 at T_1 is steeper than T_3 . Initial rate is faster at T_1 . T_1 is higher temperature than T_3 as rate increases at higher temperature.

Statement 1 (False)

At T_1 , amount of P present at equilibrium is more than that at T_3 , meaning at higher temperature, more P remains at equilibrium. Position of equilibrium is favours left hand side at higher temperature. Therefore, at higher temperature, Kp **decreases** since position of equilibrium favours left hand side (endothermic).

Statement 2 (True)

At T_1 , 0.4 mol of P is left, 0.6 mol of it decomposed to form 0.3 mol of R at equilibrium. Amount of P at equilibrium (hence mole fraction of P) is larger than R at T1 at equilibrium.

Statement 3 (True)

Forward reaction is exothermic. ΔH° is negative. Since there are more moles of gaseous products than reactant, ΔS° is positive. Therefore, ΔG° is always negative (spontaneous) at all temperatures.

$$\Delta G^{\ominus} = \Delta H^{\ominus} - T \Delta S^{\ominus}$$

2 The graph below shows how the fraction of a substance, **X** represented by one of the following compounds in the equilibrium mixture shown below varies with temperature at pressures of **Y** Pa and **Z** Pa.



Identify **X** and the correct relative magnitudes of **Y** and **Z**.

	Х	Pressure
Α	N ₂	Z > Y
В	O ₂	Y > Z
С	H ₂ O	Y > Z
D	NH_3	Z > Y

Ans: A

Looking at one graph (say, at constant pressure **Z** Pa), fraction of X decreases at higher temperatures. As the reaction is exothermic, by LCP, when **temperature increases**, position of equilibrium (POE) shifts to the **left** to favour the endothermic reaction to absorb heat. More reactants and less products will be formed at new equilibrium. Therefore, **X** refers to the "products" on the RHS of the chemical equation. Options B and D can be eliminated.

By maintaining a constant temperature (draw a vertical line at any temperature cutting through both graphs), fraction of **X** (gaseous particles at the RHS) at equilibrium increases at **Y** Pa. Since the RHS, with more moles of gaseous particles is favoured **Y** Pa. Pressure of the system must be **reduced** as POE shifts right to increase the pressure. **Y** Pa is **lower** than **Z** Pa.