

Unit 2 Review package ANSWER KEY

1	A	12	C	23	B	34	A
2	C	13	C	24	A	35	C
3	D	14	A	25	B	36	B
4	B	15	D	26	B	37	D
5	A	16	C	27	A	38	C
6	C	17	D	28	C	39	D
7	B	18	D	29	A	40	D
8	A	19	D	30	D	41	B
9	D	20	C	31	D	42	B
10	A	21	C	32	D	43	B
11	B	22	A	33	C		

1. Both entropy and enthalpy favour the forward direction, therefore the reaction proceeds 100% to the right (“to completion”), and therefore equilibrium can not be attained.

2. If all three have the same starting molarity, then this means that the reaction must proceed to the right to increase the HI as the mole stoichiometry from the balanced equation suggests it should be $2\text{HI} : 1\text{H}_2 : 1\text{I}_2$

3. at 1000°C , reaction shifts left (**Stress!**) and the $[\text{NH}_3]$ would decrease at 100°C , the **rate** of the reaction would be much too low

4. $K_{eq} = 0.50$

5. This reaction is endothermic since the K is increasing when the Temperature is increased. $K \uparrow T \uparrow K \propto T$

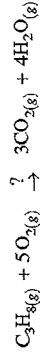
6. $K_{eq} = 0.20$

7. $x = 0.14$; $[\text{H}_2] = 0.76 \text{ M}$

8. $x = 1.8$; $[\text{CO}_2] = 1.8 \text{ M}$ $[\text{CO}] = 2.3 \text{ M}$

1.

Consider the following exothermic reaction:



Explain, in terms of increasing or decreasing entropy and enthalpy, whether or not the reaction will reach equilibrium. (3 marks)

Solution:

For Example:

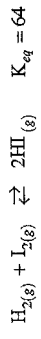
Entropy increases in the forward reaction. ← 1 mark

Enthalpy decreases in the forward reaction. ← 1 mark

Since both favour products, equilibrium will not be attained, or the reaction will go to completion. ← 1 mark

2.

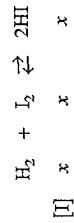
Given the reacting system:



Equal moles of H_2 , I_2 and HI are placed in a 1.0L container. Use calculations to determine the direction the reaction will proceed in order to reach equilibrium. (3 marks)

Solution:

For Example:



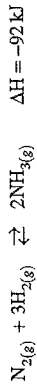
$$\begin{aligned}
 \text{Trial } K_{eq} &= \frac{[HI]^2}{[H_2][I_2]} \\
 &= \frac{(x)^2}{(x)(x)} \\
 &= 1
 \end{aligned}$$

← 2 marks

Since $\text{Trial } K_{eq} < K_{eq}$, equilibrium is established by proceeding to the right. ← 1 mark

3.

Consider the following reaction for the Haber Process for ammonia production:



The system is normally maintained at a temperature of approximately 500°C.

- a) Explain why 1000°C is not used. (1 mark)

Solution:

For Example:

Equilibrium will be shifted to the left, reducing the yield of NH₃. ← 1 mark

- b) Explain why 100°C is not used. (1 mark)

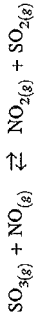
Solution:

For Example:

The rate of the reaction would be too low. ← 1 mark

4.

Consider the following equilibrium:



In an experiment, 0.100 moles of SO₃ and 0.100 moles of NO are placed in a 1.00 L container. When equilibrium is achieved, [NO₂] = 0.0414 mol/L.

(3 marks)

Calculate the K_{eq} value.

Solution:

For Example:

	SO ₃	NO	NO ₂	SO ₂
[I]	0.100	0.100	0	0
[C]	-0.0414	-0.0414	+0.0414	+0.0414
[E]	0.059	0.059	0.0414	0.0414

← 1½ marks

$$K_{eq} = \frac{[\text{NO}_2][\text{SO}_2]}{[\text{SO}_3][\text{NO}]}$$

$$= \frac{(0.0414)(0.0414)}{(0.059)(0.059)}$$

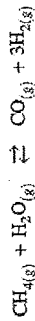
$$= 0.50$$

← ½ mark

← 1 mark

5.

Consider the following equilibrium:



K_{eq}	Temperature
1.78×10^{-3}	800°C
4.68×10^{-2}	1000°C

Is the forward reaction in this equilibrium exothermic or endothermic? Explain your answer.

Solution:

For Example:

This equilibrium is endothermic.

Since K_{eq} increases as a result of a temperature increase, equilibrium has shifted to the right.

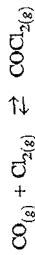
← 1 mark

← 2 marks

(3 marks)

6.

Consider the following equilibrium:



At equilibrium, the system contains 2.00 mol CO, 1.00 mol Cl_2 and 0.200 mol COCl_2 in a 2.0 L container. Calculate the value of K_{eq} . (3 marks)

Solution:

For Example:

$$K_{eq} = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]}$$

← 1/2 mark

$$= \frac{(0.200 \text{ mol}/2.0 \text{ L})}{(2.00 \text{ mol}/2.0 \text{ L})(1.00 \text{ mol}/2.0 \text{ L})}$$

← 1 1/2 mark

$$= \frac{(0.100)}{(1.00)(0.500)}$$

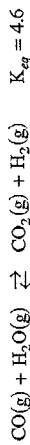
← 1 mark

$$= 0.20$$

(Deduct 1/2 mark for incorrect significant figures.)

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Consider the following equilibrium:



Initially, 0.50 mol CO , 0.50 mol H_2O , 0.62 mol CO_2 and 0.62 mol H_2 are placed in a 1.0 L container, and the reaction proceeds towards products. Calculate the equilibrium $[\text{H}_2]$. (3 marks)

Solution:

For Example:

	CO(g)	$+$	$\text{H}_2\text{O(g)}$	\rightleftharpoons	$\text{CO}_2\text{(g)}$	$+$	$\text{H}_2\text{(g)}$
[I]	0.50		0.50		0.62		0.62
[C]	$-x$		$-x$		$+x$		$+x$
[E]	$0.50 - x$		$0.50 - x$		$0.62 + x$		$0.62 + x$

← 3 marks

$$K_{eq} = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}$$

$$\sqrt{4.6} = \sqrt{\frac{(0.62 + x)^2}{(0.50 - x)^2}}$$

$$x = 0.14$$

$$[\text{H}_2] = (0.62 + x) = 0.76 \text{ M}$$

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Consider the equilibrium: $\text{CO}_2\text{(g)} + \text{H}_2\text{(g)} \rightleftharpoons \text{CO(g)} + \text{H}_2\text{O(g)}$ $K_{eq} = 1.60$

Initially, 8.2 mol of CO and 8.2 mol of H_2O are placed in a 2.0 L container and allowed to react. Calculate the equilibrium concentrations of CO_2 and CO . (4 marks)

Solution:

For Example:

$$[\text{CO}] = [\text{H}_2\text{O}] = \frac{8.2 \text{ mol}}{2.0 \text{ L}} = 4.1 \text{ M}$$

	$\text{CO}_2\text{(g)}$	$+$	$\text{H}_2\text{(g)}$	\rightleftharpoons	CO(g)	$+$	$\text{H}_2\text{O(g)}$
[I]	0		0		4.1		4.1
[C]	$+x$		$+x$		$-x$		$-x$
[E]	x		x		$4.1 - x$		$4.1 - x$

← 1 mark

$$K_{eq} = \frac{[\text{CO}][\text{H}_2\text{O}]}{[\text{CO}_2][\text{H}_2]}$$

$$\sqrt{1.60} = \sqrt{\frac{(4.1 - x)^2}{x^2}}$$

$$x = 1.8$$

$$[\text{CO}_2] = 1.8 \text{ M}, [\text{CO}] = 4.1 - x = 2.3 \text{ M}$$

← 1 mark

← 1 mark

← 1 mark