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 2HI : $1H_2$: $1I_2$

3. at 1000°C, reaction shifts left (**Stress!**) and the $[NH_3]$ 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. $K \Uparrow \; T \Uparrow \quad K \propto T$

6. $K_{eq} = 0.20$

7. x = 0.14; $[H_2] = 0.76$ M

8. x = 1.8; $[CO_2] = 1.8$ M $[CO] = 2.3$ M

1. Consider the following exothermic reaction:

$$C_3H_{8(g)} + 5O_{2(g)} \overset{?}{\rightarrow} 3CO_{2(g)} + 4H_2O_{(g)}$$

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:

$$H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)} \qquad K_{eq} = 64$$

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

Solution:

For Example:

	H_2	+	I_2	$\rightleftarrows$	2HI
[I]	x		x		x

$$\text{Trial } K_{eq} = \frac{[HI]^2}{[H_2][I_2]} = \frac{(x)^2}{(x)(x)} = 1$$ ← **2 marks**

Since 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:

$$N_{2(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)} \qquad \Delta H = -92\text{ kJ}$$

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 NH_3. **← 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:

$$SO_{3(g)} + NO_{(g)} \rightleftarrows NO_{2(g)} + SO_{2(g)}$$

In an experiment, 0.100 moles of SO_3 and 0.100 moles of NO are placed in a 1.00 L container. When equilibrium is achieved, $[NO_2] = 0.0414$ mol/L. Calculate the K_{eq} value. **(3 marks)**

Solution:

For Example:

	SO_3	+	NO	$\rightleftarrows$	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

← $1\frac{1}{2}$ marks

$$K_{eq} = \frac{[NO_2][SO_2]}{[SO_3][NO]}$$ **← $\frac{1}{2}$ mark**

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

$$= 0.50$$ **← 1 mark**

5. Consider the following equilibrium:

$$CH_{4(g)} + H_2O_{(g)} \rightleftarrows CO_{(g)} + 3H_{2(g)}$$

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. **(3 marks)**

Solution:

For Example:

This equilibrium is endothermic. ← **1 mark**

Since K_{eq} increases as a result of a temperature increase, equilibrium has shifted to the right. } ← **2 marks**

6. Consider the following equilibrium:

$$CO_{(g)} + Cl_{2(g)} \rightleftarrows COCl_{2(g)}$$

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:

$$K_{eq} = \frac{[COCl_2]}{[CO][Cl_2]} \quad \leftarrow \tfrac{1}{2}\ \textbf{mark}$$

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

$$= \frac{(0.100)}{(1.00)(0.500)} \quad \leftarrow 1\tfrac{1}{2}\ \textbf{mark}$$

$$= 0.20 \quad \leftarrow 1\ \textbf{mark}$$

(Deduct $\frac{1}{2}$ **mark** for incorrect significant figures.)

7. Consider the following equilibrium:

$$CO(g) + H_2O(g) \rightleftarrows CO_2(g) + H_2(g) \qquad K_{eq} = 4.6$$

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 $[H_2]$. **(3 marks)**

Solution:

For Example:

	$CO(g)$	+	$H_2O(g)$	$\rightleftarrows$	$CO_2(g)$	+	$H_2(g)$
[I]	0.50		0.50		0.62		0.62
[C]	$-x$		$-x$		$+x$		$+x$
[E]	$0.50-x$		$0.50-x$		$0.62+x$		$0.62+x$

$$K_{eq} = \frac{[CO_2][H_2]}{[CO][H_2O]}$$

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

$$x = 0.14$$

$$[H_2] = (0.62+x) = 0.76\,M$$

← **3 marks**

8. Consider the equilibrium: $CO_2(g) + H_2(g) \rightleftarrows CO(g) + H_2O(g) \qquad K_{eq} = 1.60$

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:

$$[CO] = [H_2O] = \frac{8.2\text{ mol}}{2.0\text{ L}} = 4.1\,M$$

	$CO_2(g)$	+	$H_2(g)$	$\rightleftarrows$	$CO(g)$	+	$H_2O(g)$	
[I]	0		0		4.1		4.1	
[C]	$+x$		$+x$		$-x$		$-x$	
[E]	x		x		$4.1-x$		$4.1-x$	← **1 mark**

$$K_{eq} = \frac{[CO][H_2O]}{[CO_2][H_2]}$$

$$\sqrt{1.60} = \sqrt{\frac{(4.1-x)^2}{x^2}} \qquad \leftarrow \textbf{1 mark}$$

$$x = 1.8 \qquad \leftarrow \textbf{1 mark}$$

$$[CO_2] = 1.8\,M,\ [CO] = 4.1 - x = 2.3\,M \qquad \leftarrow \textbf{1 mark}$$