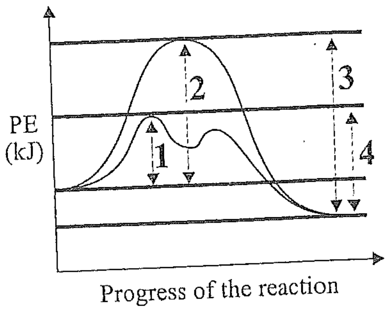
UNIT TEST 2 – DYNAMIC EQUILIBRIUM

1. Consider the following PE diagram:



Identify the activation energy for the forward uncatalysed reaction.

A. 1

B. 2

C. 3

D. 4

CHALLENGER QUESTION 58.5

2. In which of the following will entropy and enthalpy factors favour the establishment of an equilibrium?

A. $CaCO_{3(s)} + 178\,kJ \xrightarrow{?} CaO_{(s)} + CO_{2(g)}$

B. $Mg_{(s)} + 2HCl_{(aq)} \xrightarrow{?} MgCl_{2(aq)} + H_{2(g)} + 425\,kJ$

C. $2C_{(s)} + 2H_{2(g)} \xrightarrow{?} C_2H_{4(g)}$ $\Delta H = +52.3\,kJ$

D. $2C_2H_{6(g)} + 7O_{2(g)} \xrightarrow{?} 4CO_{2(g)} + 6H_2O_{(g)}$ $\Delta H = -1\,560\,kJ$

3. Consider the following equilibrium:
$CO_{(g)} + 2H_{2(g)} \rightleftarrows CH_3OH_{(g)}$ $\Delta H = -91\,kJ$

Which of the factors below would increase the concentration of CH_3OH at equilibrium?

A. an addition of CO

B. an increase in the volume

C. a decrease in the pressure

D. an increase in the temperature

CHALLENGER QUESTION 44.6

4. Consider the following equilibrium:
$PCl_{3(g)} + Cl_{2(g)} \rightleftarrows PCl_{5(g)}$

If the volume of the system is decreased, how will the reaction rates in the new equilibrium compare with the rates in the original equilibrium?

	Forward Rate	Reverse Rate
A.	increases	increases
B.	increases	decreases
C.	decreases	decreases
D.	decreases	increases

5. Consider the following equilibrium:

$$H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)} \quad \Delta H = -71.9\,\text{kJ}$$

colourless purple colourless

Which of the following would allow you to conclude that the system has reached equilibrium?

A. The pressure remains constant.

B. The reaction rates become zero.

C. The colour intensity remains constant.

D. The system shifts completely to the right.

6. Consider the following equilibrium:

$$Fe_2O_{3(s)} + 3CO_{(g)} \rightleftarrows 2Fe_{(s)} + 3CO_{2(g)}$$

Identify the equilibrium constant expression.

A. $K_{eq} = \dfrac{[CO_2]^3}{[CO]^3}$

B. $K_{eq} = \dfrac{[CO_2]}{[CO]}$

C. $K_{eq} = \dfrac{[CO_2]^3[Fe]^2}{[Fe_2O_3][CO]^3}$

D. $K_{eq} = \dfrac{[Fe_2O_3][CO]^3}{[CO_2]^3[Fe]^2}$

7. Consider the following equilibrium system:

$$2NO_{(g)} + Cl_{2(g)} \rightleftarrows 2NOCl_{(g)} \quad \Delta H = -77\,\text{kJ}$$

In which direction will the equilibrium shift and what happens to the value of K_{eq} when the temperature of the system is increased?

	Shift	K_{eq}
A.	right	increases
B.	right	decreases
C.	left	increases
D.	left	decreases

8. Consider the following equilibrium:

$$CO_{(g)} + 2H_{2(g)} \rightleftarrows CH_3OH_{(g)}$$

At equilibrium it was found that $[CO] = 0.105$ mol/L, $[H_2] = 0.250$ mol/L and $[CH_3OH] = 0.00261$ mol/L. Which of the following is the equilibrium constant value?

A. 9.94×10^{-2}

B. 0.398

C. 2.51

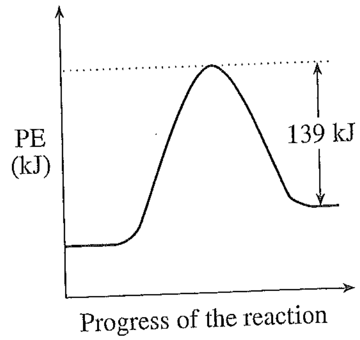
D. 10.0

9. The following forward reaction has an $E_a = 167$ kJ:

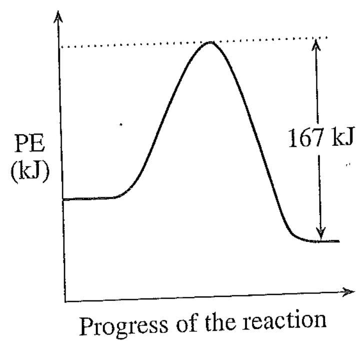
$$28\,\text{kJ} + H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$$

Which of the PE diagrams below represents this reaction?

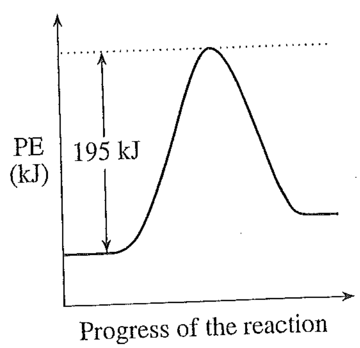
A.



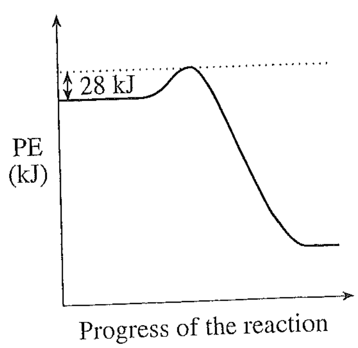
B.



C.



D.



10. A small amount of HCl is added to the following equilibrium system:

$$2CrO_{4\,(aq)}^{2-} + 2H^{+}_{(aq)} \rightleftarrows Cr_2O_{7\,(aq)}^{2-} + H_2O_{(l)}$$

How do the $\left[CrO_4^{2-}\right]$ and the reverse reaction rate change as equilibrium is re-established?

	$\left[CrO_4^{2-}\right]$	**Reverse Rate**
A.	increases	increases
B.	increases	decreases
C.	decreases	decreases
D.	decreases	increases

Use the following equilibrium to answer the next two questions.

$$N_2O_{4(g)} \rightleftarrows 2NO_{2(g)}$$

colourless brown

11. If N_2O_4 is placed in a flask at a constant temperature, which of the following is true as the system approaches equilibrium?

A. The colour gets darker as $[NO_2]$ increases.

B. The colour gets lighter as $[NO_2]$ decreases.

C. The colour gets darker as $[N_2O_4]$ increases.

D. The colour gets lighter as $[N_2O_4]$ decreases.

12. The system above reaches equilibrium. Considering enthalpy and entropy factors, which of the following is true with respect to the forward reaction?

   A. The entropy is increasing and the reaction is exothermic.

   B. The entropy is decreasing and the reaction is exothermic.

   C. The entropy is increasing and the reaction is endothermic.

   D. The entropy is decreasing and the reaction is endothermic.

13. In which of the following reactions do the tendencies for minimum enthalpy and maximum entropy both favour reactants?

   A. $3O_{2(g)} \overset{?}{\rightleftarrows} 2O_{3(g)} \quad \Delta H = +285\,kJ$

   B. $N_{2(g)} + 3H_{2(g)} \overset{?}{\rightleftarrows} 2NH_{3(g)} \quad \Delta H = -92\,kJ$

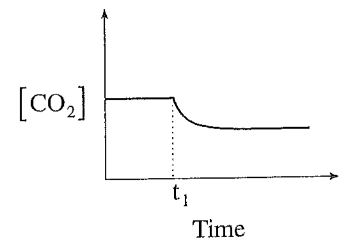
   C. $2BrCl_{(g)} \overset{?}{\rightleftarrows} Br_{2(g)} + Cl_{2(g)} \quad \Delta H = -29.3\,kJ$

   D. $CaCO_{3(s)} \overset{?}{\rightleftarrows} CaO_{(s)} + CO_{2(g)} \quad \Delta H = +175\,kJ$

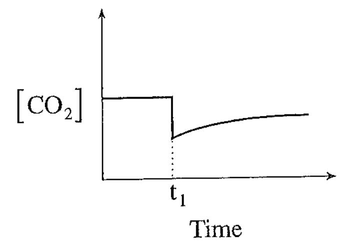
14. Consider the following equilibrium:
$CaCO_{3(s)} \rightleftarrows CaO_{(s)} + CO_{2(g)} \quad \Delta H = +175\,kJ$

Which of the following diagrams best represents the change in the concentration of CO_2 as temperature is decreased at time t_1?

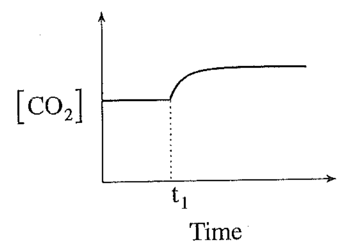
A.



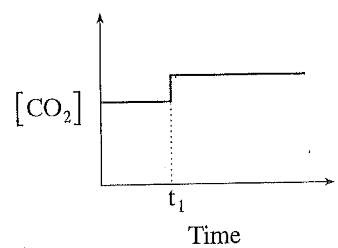
B.



C.



D.



15. The Haber Process is used to produce ammonia commercially according to the following equilibrium:
$N_{2(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)} + \text{energy}$

Which of the following conditions will produce the highest yield of ammonia?

A. increase temperature and increase pressure

B. increase temperature and decrease pressure

C. decrease temperature and increase pressure

D. decrease temperature and decrease pressure

16. Consider the following equilibrium:
$Cr_2O_{7\,(aq)}^{2-} + 2OH^-_{(aq)} \rightleftarrows 2CrO_{4\,(aq)}^{2-} + H_2O_{(l)}$
$K_{eq} = 4.14$

The concentration of ions at equilibrium was measured at a specific temperature and found to be $[Cr_2O_7^{2-}] = 0.100$ M and $[OH^-] = 0.020$ M. What is the equilibrium $[CrO_4^{2-}]$?

A. 1.7×10^{-4} M

B. 3.1×10^{-3} M

C. 1.3×10^{-2} M

D. 2.0×10^{-1} M

17. In which of the following equilibria does the concentration of reactants equal the concentration of products?

A. $N_2O_{4(g)} \rightleftarrows 2NO_{2(g)}$ $K_{eq} = 0.71$

B. $H^+_{(aq)} + OH^-_{(aq)} \rightleftarrows H_2O_{(l)}$ $K_{eq} = 1.0 \times 10^{14}$

C. $CO_{2(g)} + H_{2(g)} \rightleftarrows CO_{(g)} + H_2O_{(g)}$ $K_{eq} = 0.279$

D. $SnO_{2(s)} + 2H_{2(g)} \rightleftarrows Sn_{(s)} + 2H_2O_{(g)}$ $K_{eq} = 1.00$

18. Consider the following equilibrium:
$C_{(s)} + H_2O_{(g)} \rightleftarrows CO_{(g)} + H_{2(g)}$

At equilibrium, 4.0×10^{-2} mol H_2, 4.0×10^{-2} mol CO, 1.0×10^{-2} mol H_2O and 1.0×10^{-2} mol C were present in a 1.0 L container. What is the value of K_{eq}?

A. 0.063 B. 0.16

C. 6.3 D. 16

19. Consider the following equilibrium:
$2COF_{2(g)} \rightleftarrows CO_{2(g)} + CF_{4(g)}$

Initially, 0.12 M CO_2 and 0.20 M CF_4 are placed in a container. At equilibrium, it is found that the $[COF_2]$ is 0.040M. What is the value of K_{eq}?

A. 0.089 B. 0.45

C. 8.0 D. 11

20. Which of the solutes below can form an ionic solution with the highest conductivity?

A. PbS B. CH_3Cl

C. $NaNO_3$ D. CH_3COOH

Written Response

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

$N2_{(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)} \quad \Delta H = -92\ kJ$

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

a) Explain why 1 000°C is not used. **(1 mark)**

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

2. 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)**

3. Consider the following exothermic reaction:

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

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

4. Given the reacting system:

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

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)**

5. Consider the following equilibrium:

$N_{2(g)} + O_{2(g)} \rightleftarrows 2NO_{(g)} \quad K_{eq} = 0.0095$

Initially, 0.15 mol N_2 and 0.15 mol O_2 were placed in a 1.0 L container. Calculate the concentration of all species at equilibrium. **(4 marks)**

ANSWERS AND SOLUTIONS
UNIT TEST 2 – DYNAMIC EQUILIBRIUM

1. B	5. C	9. A	13. A	17. D
2. A	6. A	10. D	14. A	18. B
3. A	7. D	11. A	15. C	19. D
4. A	8. B	12. C	16. C	20. C
				WR1.–5. See Solution

1. **B**

   2 is correct. *1* shows the activation energy of the *catalyzed* **forward** reaction. *3* is the activation energy of *uncatalyzed* **reverse** reaction, and *4* is the activation energy of the *catalyzed* **reverse** reaction.

2. A

   For an equilibrium to be established, entropy and enthalpy factors must favour opposite sides of the reaction equation. In reaction A enthalpy favours reactants while entropy favours products (1 particle becomes 2).

3. A

   According to Le Châtelier's Principle, addition of CO would raise its concentration and cause the equilibrium to shift right, increasing the concentration of CH_3OH.

4. A

   Decreasing the volume of the system will increase the pressure. For a homogeneous gaseous system this is equivalent to increasing the concentration of each substance in the reaction. Forward and reverse reaction rates will both be increased.

5. C

   At equilibrium all concentrations remain constant. *Therefore* the colour intensity of the purple $I_{2(g)}$ will remain constant.

6. A

   For the *reaction* a A + b B $\rightleftarrows$ c C + d D,

   $K_{eq} = \dfrac{[C]^c \bullet [D]^d}{[A]^a \bullet [B]^b}$. *Solids* are not included in equilibrium constants therefore the correct answer is A.

7. D

   According to Le Châtelier's Principle, an increase in *temperature* will shift the equilibrium left, since it is an exothermic reaction. Since $K_{eq} = \dfrac{[\mathrm{NOCl}]^2}{[\mathrm{NO}]^2 \bullet [\mathrm{Cl_2}]}$ and concentration of NOCl decreases while concentrations of NO and Cl_2 both increase, K_{eq} must decrease.

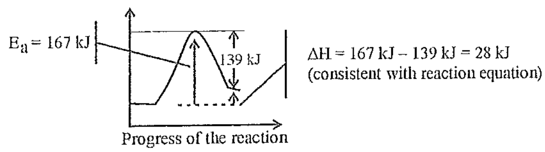
8. **B**

$$K_{eq} = \frac{[\mathrm{CH_3OH}]}{[\mathrm{CO}]\bullet[\mathrm{H_2}]^2}$$

$$= \frac{(0.00261\,\mathrm{mol/L})}{(0.105\,\mathrm{mol/L})\bullet(0.250\,\mathrm{mol/L})^2} = 0.398$$

9. A

This reaction must be **endothermic**, leaving only choices A and C. Given the fact that E_a is 167 kJ for the forward reaction the following the answer can be seen on the diagram with the attached notes:



10. D

Adding HCl will increase $[H^+]$, and according to Le Châtelier's Principle, the equilibrium should shift right to consume some of the added H^+. This should cause the $\left[CrO_4^{\ 2-}\right]$ to decrease and the rate of the reverse reaction to increase as concentrations of products increase.

11. A

If pure, *colourless* $N_2O_{4(g)}$ is in a closed system, it will become an equilibrium mixture of $NO_{2(g)}$ and $N_2O_{4(g)}$. Since $N_2O_{4(g)}$ is colourless and $NO_{2(g)}$ is brown, the system will get darker as $NO_{2(g)}$ is formed.

12. C

Bonds must break $N_2O_{4(g)}$ goes to $NO_{2(g)}$. *Breaking* bonds is endothermic. At the same time one molecule of $N_2O_{4(g)}$ becomes two molecules of $NO_{2(g)}$. This is a more disordered, higher entropy state.

13. A

In the reaction $3\,O_{2(g)} \rightleftharpoons 2\,O_{3(g)}$ molecules are in equilibrium with two. The reactants are more disordered, therefore entropy favours reactants. At the same time the reaction is endothermic *meaning* that the minimum energy side is the reactant side of the equation. Equilibria will be positioned to maximize entropy and minimize energy.

14. A

A decrease in temperature should shift the *equilibrium* left, causing $\left[CO_2\right]$ to decrease.

15. C

To get the highest yield of ammonia, the *equilibrium* needs to shift right. Using Le Châtelier's Principle it can be determined that decreasing the temperature will shift the equilibrium right as will increasing the pressure since there are fewer moles of gaseous products than gaseous reactants.

16. C

$$K_{eq} = 4.14 = \frac{\left[CrO_4^{2-}\right]^2}{\left[Cr_2O_7^{2-}\right]\left[OH^-\right]^2}$$

$$4.14 = \frac{\left[CrO_4^{2-}\right]^2}{(0.100\,M)(0.020\,M)^2}$$

$$\left[CrO_4^{2-}\right] = 1.3 \times 10^{-2}\ M$$

17. D

In *choice* D, $K_{eq} = 1.00 = \dfrac{[H_2O]^2}{\left[H_2\right]^2}$.

Therefore the reactant and product concentrations are equal. The solids cannot vary in concentration and they are not considered in this comparison.

18. B

$$K_{eq} = \frac{[H_2][CO]}{[H_2O]} = \frac{(0.040\,M)(0.040\,M)}{(0.010\,M)} = 0.16$$

19. D

Use an ICE BOX where I = initial, C = change and E = equilibrium.

	$2\,COF_{2(g)}$	$\rightleftharpoons CO_{2(g)}$	$+ CF_{4(g)}$
I	0	0.12 M	0.20 M
C	+0.040 M	$-0.040\text{ M} \times \frac{1}{2}$	$-0.040\text{ M} \times \frac{1}{2}$
E	0.040 M	0.10 M	0.18 M

$$K_{eq} = \frac{[CO_2][CF_4]}{[COF_2]^2} = \frac{(0.10\,M)(0.18\,M)}{(0.040\,M)^2} = 11$$

20. C

The solution with the highest ion concentration will be the best conductor. PbS, choice A, has a low solubility and therefore a low ion *concentration*. CH_3Cl, choice B, is molecular and will not dissociate to form ions. CH_3COOH, choice D, is a weak acid and its degree of ionization is very low. $NaNO_3$, choice C, is an ionic compound of high solubility and therefore will have a large ion concentration as it is presumed to dissociate completely.

Written Response

1. *Consider the following reaction for the Haber Process for ammonia production:*

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

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

a) *Explain why 1 000°C is not used.* **(1 mark)**

Solution:

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:

The rate of the reaction would be too low, thereby reducing the yield of NH_3. ←**1 mark**

2. *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:

	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

$\left(1\frac{1}{2}\text{ marks}\right)$

$$K_{eq} = \frac{[NO_2][SO_2]}{[SO_3][NO]} = \frac{(0.0414)(0.0414)}{(0.059)(0.059)} \quad \leftarrow \frac{1}{2}\text{ mark}$$

$$= 0.50 \quad \leftarrow 1\text{ mark}$$

3. *Consider the following exothermic reaction:*

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

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

Solution:

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**

4. *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:

Example:

	H_2	+	I_2	⇄	2HI
[I]	x		x		x

← **2 marks**

$$\text{Trial } K_{eq} = \frac{[HI]^2}{[H_2][I_2]} = \frac{(x)^2}{(x)(x)} = 1 \quad \leftarrow 1\text{ mark}$$

Since Trial $K_{eq} < K_{eq}$, equilibrium is established by proceeding to the right.

5. *Consider the following equilibrium:* **(4 marks)**

For Example:

	N_2	+	O_2	⇄	2NO
[I]	0.15		0.15		0
[C]	$-x$		$-x$		$+2x$
[E]	$0.15-x$		$0.15-x$		$2x$

$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$$

$$0.0095 = \frac{(2x)^2}{(0.15-x)^2}$$

$$\sqrt{0.0095} = \sqrt{\frac{(2x)^2}{(0.15-x)^2}}$$

$$x = 6.97\times10^{-3}$$

$$[N_2] = [O_2] = 0.15 - x = 0.14\text{M}$$

$$[NO] = 2(x) = 0.014\text{M}$$