Classroom Activity Chapter 14 Number 3 8 February 2017

Printed Name: ______________________________ Chm 205 Student number: ______

1. An equilibrium mixture of CO, O_2 and CO_2 at a certain temperature contains 0.0010 M CO_2 and 0.0015 M O_2. At this temperature, K_c equals 140 for the reaction:

M	$2\,CO(g)$	+	$O_2(g)$	$\rightleftharpoons$	$2\,CO_2(g)$
E			0.0015		0.0010

What is the equilibrium concentration of CO?

$$K_c = \frac{[CO_2]^2}{[CO]^2[O_2]}$$

$$140 = \frac{0.0010^2}{(CO)^2\, 0.0015}$$

$[CO] = 0.0022\,M$

1b. To the reaction mixture at equilibrium, suppose some additional $CO_2(g)$ was added so that $[CO_2] = 0.0020$ M. Solve for Q_c in order to predict the direction in which the shift will occur as equilibrium in established. Without actually solving for x, set up the MICE table to reflect how the predicted changes.

M	$2\,CO(g)$	+	$O_2(g)$	$\rightleftharpoons$	$2\,CO_2(g)$
I	0.0022		0.0015		0.0020
C					
E					

$$Q_c = \frac{.0020^2}{.0022^2 * 0.0015} = 550$$

$Q_c > K_c$ ∴ shift L

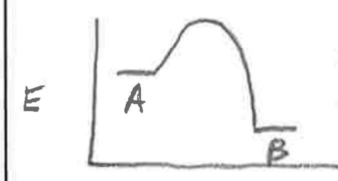
1c. To the reaction mixture at equilibrium as described in Question 1a, suppose the volume of the vessel could be adjusted so that in was now half the size it was in Question 1a. What are the new initial concentrations? Solve for Q_c in order to predict the direction in which the shift will occur as equilibrium in established. Without actually solving for x, set up the MICE table to reflect how the predicted changes.

M	$2\,CO(g)$	+	$O_2(g)$	$\rightleftharpoons$	$2\,CO_2(g)$
I	0.0044		0.0030		0.0020
C					
E					

$$Q_c = \frac{.0020^2}{0.0044^2 * 0.0030} = 69$$

$Q_c < K_c$ ∴ shift R

1d. The value of K_c decreases as the temperature is increased. Is the reaction exothermic or endothermic? Sketch a reaction profile (using one hump because we do not know the mechanism) to describe the reaction.



"Increasing the temperature always favors the endothermic direction"

As temperature is increased, [A]↑ and [B]↓ so K_c ↓

2. What effect will a change in temperature have on the value of K_p?

A. It will have no effect on the value of K_p.

B. The value of K_p always decreases with an increase in temperature.

C. The value of K_p always increases with an increase in temperature.

D. The value of K_p will decrease or increase with an increase in temperature, depending on whether the reaction is exothermic or endothermic.

3. The dissolution of calcium hydroxide is exothermic. What happens when the solution of $Ca(OH)_2$ is heated?

$$Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2\,OH^-(aq)$$

A. The amount of $Ca(OH)_2(s)$ decreases.

B. The amount of $Ca(OH)_2(s)$ increases.

C. The amount of $Ca(OH)_2(s)$ remains unchanged.

D. The $Ca(OH)_2(s)$ completely dissolves.

4. A reaction reaches dynamic equilibrium at a given temperature when

A. the amount of products exceeds the amount of reactants.

B. k_{fwd} equals k_{rev}.

C. opposing reactions cease and the system is static.

D. the relative amounts of reactants and products are constant and $rate_{fwd} = rate_{rev}$.

Classroom Activity Chapter 14 Number 2 6 February 2017

1. Consider the reaction given below. At 425 °C, the equilibrium constant, $K_c = 54.4$. Starting with, $[HI]_I = 0.300$ M HI, what are the equilibrium concentrations of all three gases, $[HI]_E$, $[H_2]_E$, and $[I_2]_E$?

M	$2\ HI(g)$	→	$H_2(g)$	+	$I_2(g)$
I	0,300		0		0
C	$-2x$		x		x
E	$0,300-2x$ $=0,090$		$x=0,140$		$x=0,140$

$$K_c = \frac{[H_2][I_2]}{[HI]^2} = \frac{x^2}{(0.30-2x)^2} = 54.4$$

Take square root:

$$\frac{x}{0.300-2x} = 7.376 \Rightarrow x = 0.14$$

1b. Suppose that $[HI]_I = 0.0700$ M, $[H_2]_I = [I_2]_I = 0.400$ M. Which direction must the equilibrium shift in order to attain equilibrium? What are the equilibrium concentrations of all three gases, $[HI]_E$, $[H_2]_E$, and $[I_2]_E$?

M	$2\ HI(g)$	→	$H_2(g)$	+	$I_2(g)$	
I	0,0700		0,400		0,400	Q_c!
C	$+2x$		$-x$		$-x$	
E	$0,070-2x$		$0,400+x$		$0,400+x$	

$$Q_c = \frac{0,400^2}{0,0700^2} = 32,7 \ \therefore \text{ shift R}$$

$$K_c = \frac{(0,400+x)^2}{(0,070-2x)^2} = 54,4 \Rightarrow x = 0,00738$$

$[HI] = 0,0552$ $[H_2] = [I_2] = 0,407$

2. Consider the reaction at 250 °C:

$CO(g)$ + $2\ H_2(g)$ → $CH_3OH(g)$ $K_c = 42.8$

2a. What is the relationship between K_c and K_p?

$$K_p = K_c(RT)^{\Delta n_{gas}} \quad \Delta n_g = -2$$

2b. Solve for K_p. Given: R = 0.0821 L atm mol^{-1} K^{-1}]

$$K_p = 42,8\ (0.0821 \times 523)^{-2} = 0.0232$$

3. What best describes what would happen to the system given in Question 1, at equilibrium, if some HI were added?

A. The system would no longer be at equilibrium and would shift to the left in order to reestablish equilibrium.

B. After equiibrium is reestablished, the concentration of all three gases would be greater than they were prior to the addition of more HI(g).

C. The system would remain at equilibrium, however the amount of $H_2(g)$ and $I_2(g)$ would increase and the amount of HI(g) would decrease.

D. The concentration of HI(g) would end up lower than before the addition of HI(g).

4. What best describes what would happen to the system given in Question 2, at equilibrium, the volume of the vessel were decreased?

A. The system would no longer be at equilibrium and would shift to the left in order to reestablish equilibrium.

B. After equiibrium is reestablished, the value of K_c would be larger.

C. After equiibrium is reestablished, the value of K_c would be smaller.

D. The system would no longer be at equilibrium and $Q_c < K_c$.

5. The equilibrium constant is equal to 5.00 at 1300 K for the reaction:

$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g).$$

If initial concentrations are $[SO_2] = 1.20$ M, $[O_2] = 0.45$ M, and $[SO_3] = 1.80$ M, the system is

A. at equilibrium.

B. not at equilibrium and will remain in an unequilibrated state.

C. not at equilibrium and will shift to the left to achieve an equilibrium state.

D. not at equilibrium and will shift to the right to achieve an equilibrium state.

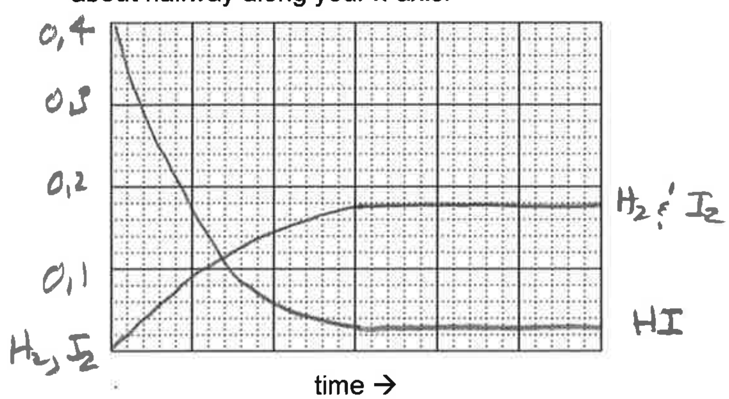
Classroom Activity Chapter 14 Number 1 3 February 2017

1. Consider the reaction given below. At 425 °C, the initial concentration of HI, $[HI]_I$ is 0.4000 M and the other concentrations are 0.0000 M. After equilibrium is reached, $[HI]_E$ = 0.0254 M. What are the equilibrium concentrations of $H_2(g) + I_2(g)$ and what is the value for K_c?

M	$2\,HI(g)$	→	$H_2(g)$	+	$I_2(g)$
I	0,400		0		0
C	−2X		+X		+X
E	0,400-2X		X = 0,1873		X = 0,1873

$$\boxed{\begin{gathered} 0{,}400 - 2X = 0{,}0254 \\ X = 0{,}1873 \\ K_c = \frac{[H_2][I_2]}{[HI]^2} = \frac{0{,}1873^2}{0{,}0254^2} = 54{,}4 \end{gathered}}$$

1b. Plot these concentrations on the y-axis as a function of time on this graph. The kinetics region is not zero order. Assume equilibrium is reached about halfway along your x-axis.



1c. What would be the equilibrium constant if the initial concentrations of HI, $[HI]_I$ is 0.600 M and the other concentrations are 0.00 M?

It'll still be 54,5

1d. What is the value of K_c for the reverse reaction?

$$H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)$$

$$\boxed{K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{1}{54{,}5} = 0{,}0184}$$

2. Consider the reaction given below. At 250 °C, the initial concentrations are 0.246 M CO, 0.491 M H_2, and 0.00 M CH_3OH. At equilibrium, [CO] = 0.096 M. What are the equilibrium concentrations of all three gases and what is the value for K_c?

M	$CO(g)$	+	$2\,H_2(g)$	→	$CH_3OH(g)$
I	0,246		0,491		0
C	−X		−2X		+X
E	0,246-X		−2X = 0,0191		X = 0,15

$$\boxed{\begin{gathered} 0{,}246 - X = 0{,}096 \rightarrow X = 0{,}15 \\ K_c = \frac{[CH_3OH]}{[CO][H_2]^2} = \frac{0{,}15}{0{,}096 * 0{,}191^2} = 42{,}8 \end{gathered}}$$

Questions in final exam format:

3. Which one of the following statements does **not** describe the equilibrium state?

A. Equilibrium is dynamic and there is no net conversion to reactants and products.

B. The concentration of the reactants is equal to the concentration of the products.

C. The concentration of the reactants and products reach a constant level.

D. The rate of the forward reaction is equal to the rate of the reverse reaction.

4. If $K_c = 7.04 \times 10^{-2}$ for the reaction:

$$2\,HBr(g) \rightleftharpoons H_2(g) + Br_2(g)$$

What is the value of K_c for this reaction?

$$\tfrac{1}{2}\,H_2(g) + \tfrac{1}{2}\,Br_2(g) \rightleftharpoons HBr(g)$$

A. 3.52×10^{-2} B. 0.265

C. 3.77 D. 28.4

5. Which of the following is the correct equilibrium constant expression for this reaction? (Hint: Balance the reaction first)

$$4\,Fe(s)\ 3\,O_2(g) \rightleftharpoons 2\,Fe_2O_3(s)$$

A. $K_c = [Fe_2O_3]/([Fe][O_2])$ B. $K_c = 1/[O_2]$

C. $K_c = [Fe_2O_3]/([Fe] + [O_2])$ D. $K_c = [O_2]^{-3}$

Classroom Activity Chapter 14 Number 4 — 10 February 2017

1. On the previous worksheet, we explored the LeChalelier Principle relating to adding a reagent to a system at equilibrium. Without doing any Qc calculations, predict which direction the following equilibrium would shift if...

$3\,H_2(g) + N_2(g) \rightleftharpoons 2\,NH_3(g)$

For each of these, write "Shift Right", "Shift Left" or "No shift"

1a. Some nitrogen gas was added. R

1b. Some hydrogen gas was removed. L

1c. Some ammonia gas was removed. R

1d. Some ammonia gas was added. L

1e. Some argon gas was added. No

2. Another LeChatelier Principle predicts the results of a volume change. Imagine the reaction and equilibrium concentrations:

M	$A(g)$	$\rightleftharpoons$	$2\,B(g)$
E	5.0		2.0

2a. What is the value of K_c?

$$\boxed{K_c = \frac{[B]^2}{[A]} = \frac{2^2}{5} = 0.80}$$

2b. Now suppose that the volume of the contained was decreased to just half its original value. What are the new initial concentrations of A and B? Calculate Q_c. In which direction must a shift occur?

M	$A(g)$	$\rightleftharpoons$	$2\,B(g)$
E	5.0		2.0
I	10.0		4.0
C	+x		−2x
E	10.0+x		4.0−2x

$Q_c = \frac{[B]^2}{[A]} = \frac{4^2}{10} = 1.6$

∴ shift L

3. Repeat the problem for this equilibrium. Again, the volume is decreased by half.

M	$C(g)$	$\rightleftharpoons$	$D(g)$
E	4.0		3.0
I	8.0		6.0
C			
E			

$K_c = \frac{[D]}{[C]} = \frac{3}{4} = 0.75$

$Q_c = \frac{[D]}{[C]} = \frac{6}{8} = 0.75$

no shift

4. Which direction will each reaction shift if the volume is doubled. "Shift R" "Shift L" or "No shift"

4a. $2\,CO(g) + O_2(g) \rightleftharpoons 2\,CO_2(g)$ L

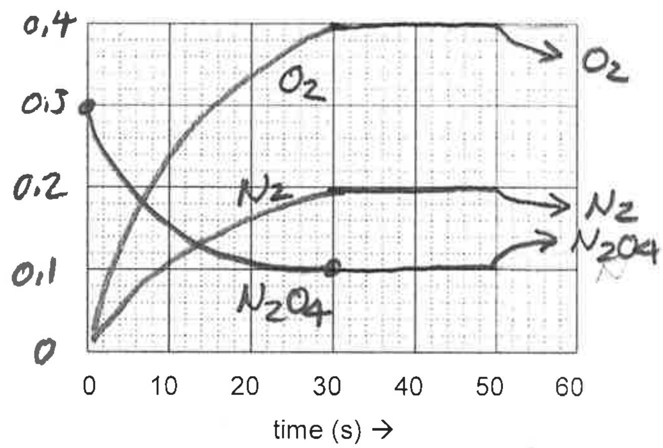
4b. $3\,H_2(g) + N_2(g) \rightleftharpoons 2\,NH_3(g)$ L

4c. $3\,H_2(g) + WO_3(s) \rightleftharpoons W(s) + 3\,H_2O(g)$ No

5. The third LeChalelier Principle involves changing the temperature. The "rule" is that "Increasing the temperature always favors the endothermic direction." Changing the temperature is the only way to change the value of K_c. For endothermic reactions, increasing the temperature will cause a shift is towards the right and K_c increases. For exothermic reactions, the opposite is true.

$N_2O_4(g) \rightleftharpoons N_2(g) + 2\,O_2(g) \quad \Delta H = -22.2$ kJ

5a. Suppose $[N_2O_4]_I = 0.30$ M, $[N_2]_I = [O_2]_I = 0$ M. At some temperature suppose $[N_2O_4]_E = 0.10$ M Further suppose that it takes about 30 s to reach equilibrium. Plot $[N_2O_4]$, $[N_2]$ and $[O_2]$ thru 50 s.



5b. What is the numerical value of K_c?

$$\boxed{K_c = \frac{[N_2][O_2]^2}{[N_2O_4]} = \frac{(0.2)(0.4)^2}{0.10}} = 0.32$$

5c. At t = 50 s, suppose the temperature was raised. With little arrows, show which direction each equilibrium concentration will shift. Also will it take more than 30 s or less than 30 s to reach equilibrium again? Or about 30 s?

It will shift L

With a higher temperature, it should take a little less than 30 s to re-establish equilibrum.