Equilibrium (#11-1)
Models of equilibrium

$$A(s) + 2B(g) + C(g) \leftrightarrow + D(l) + F(g)$$

1. Write a Kc expression for this reaction above.

   ↑ only l + g have conc.

   $K_c = \frac{[D][F]}{[B]^2[C]}$

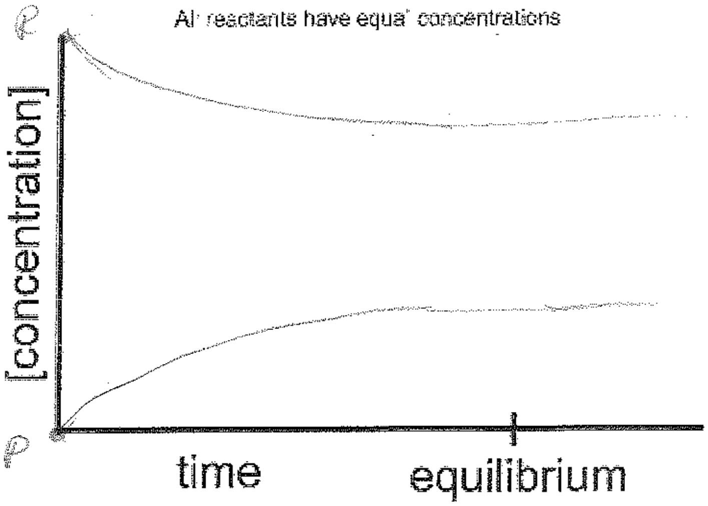
2. Write a Kp expression for the reaction above.

   ↑ only g have P

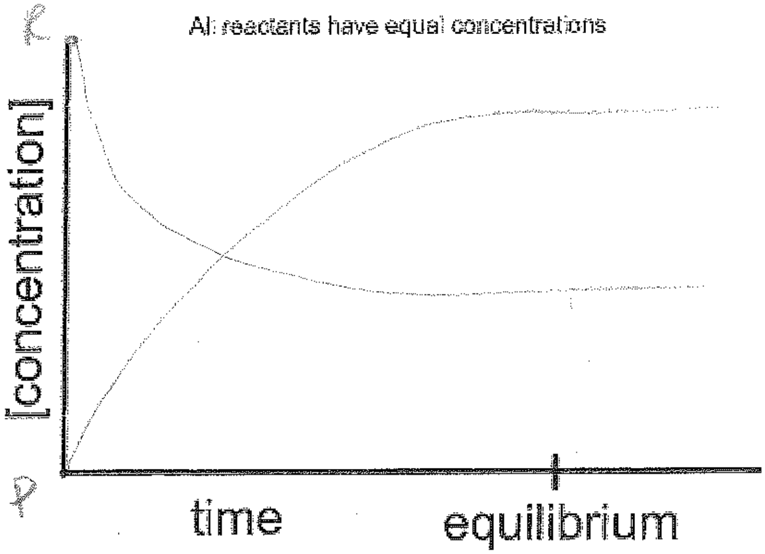
   $K_p = \frac{[P_F]}{[P_B]^2[P_C]}$

3. Fill out the chart below for the reaction above.

Reactant favored



Product favored



4. A(g) ↔ C(g) + 2D(g)

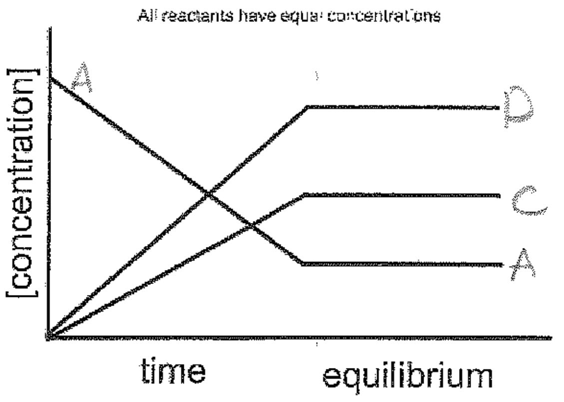
A sample of "A" is placed in a vessel and the reaction occurs until equilibrium.

a. Label line for each chemical it represents.

b. Label when equilibrium is established.

c. Write the Kc and Kp expressions for this reaction.

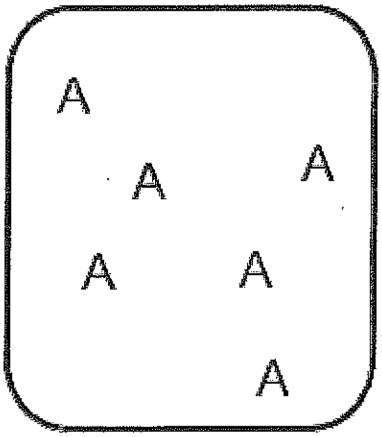
$K_c = \frac{[C][D]^2}{[A]}$ $K_p = \frac{[P_C][P_D]^2}{[P_A]}$



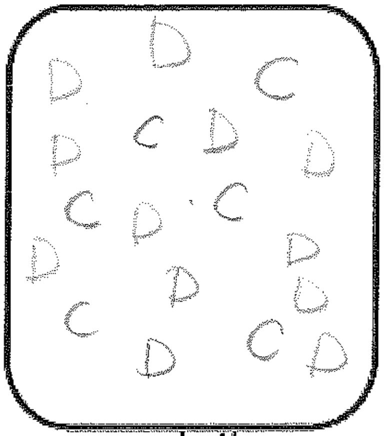
d. Complete the following particulate drawings.

A → C + 2D

(6A)



Initial



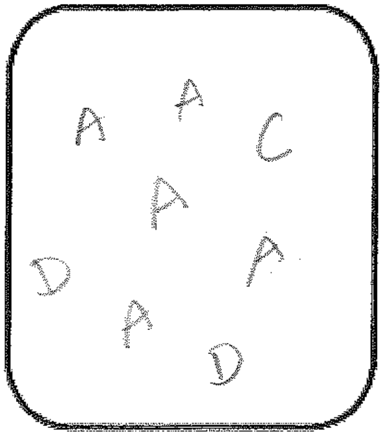
completion

A → C + 2D

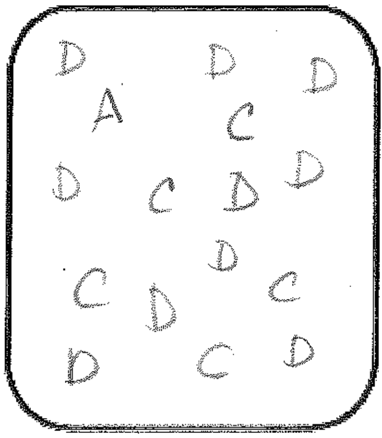
I 6		
S -6	+6	+12
E 0	6	12

A → C + 2D

I 6		
S -1	+1	+2
E 5	1	2



reactant favored



product favored

A → C + 2D

I 6		
S -5	+5	+10
E 1	5	10

(#11-1)

Introduction to Completion vs. Equilibrium

Completion:

.5 grams of methane burns in excess O_2 to produce CO_2 and H_2O. $0.5g \left| \frac{1mol}{16g} \right. = 0.031mol$

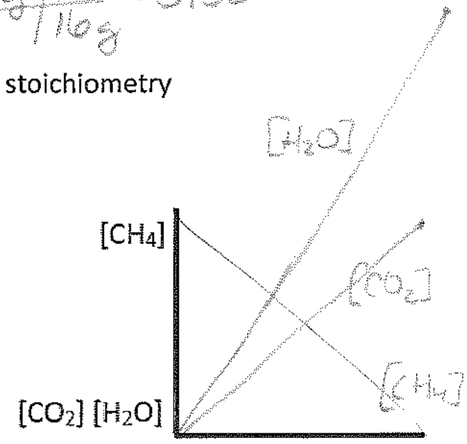
1. Write a balanced combustion reaction with an ISE table showing stoichiometry

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

	CH_4	$+2O_2$	CO_2	$+2H_2O$
I	0.31	excess	0	0
S	-0.31	-0.62	+0.31	+0.62
E	0		0.31	0.62

2. Draw a graphical representation tracking concentrations:

(not tracking O_2 because their is no starting point)



*completion

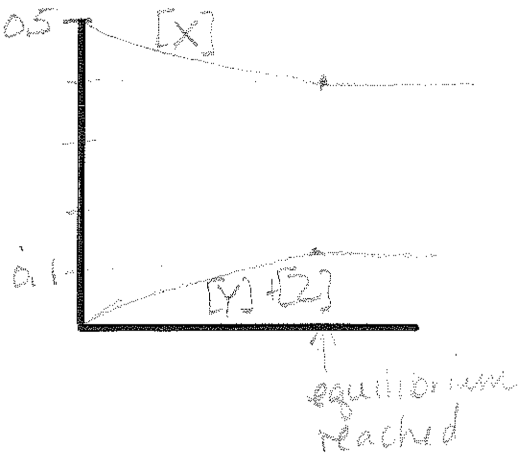
3. A Chemical process of .5M of X only goes 20% to product producing Y and Z. Write a balanced reaction with an ISC table showing ISE stoichiometry.

20% of 0.5M = 0.1M

$X \leftrightarrow Y + Z$

	X	Y	Z
I	0.5		
S	-0.1	+0.1	+0.1
E	0.4	0.1	0.1

4. Draw a graphical representation of this process?
   a. Write out the equilibrium expression Kc.
   b. Determine the value of K.



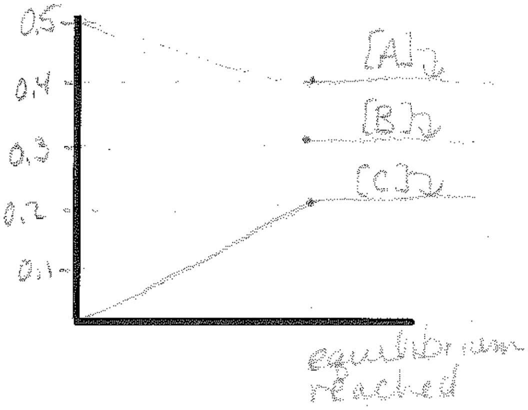
$$K_c = \frac{[Y][Z]}{[X]} = \frac{[0.1][0.1]}{[0.4]} = 0.025$$

A chemical process of $A_{(aq)} + 2B_{(aq)} \leftrightarrow 2C_{(aq)}$. A and B both start at a concentration of .5M and when the process has come to equilibrium 20% of A is converted.

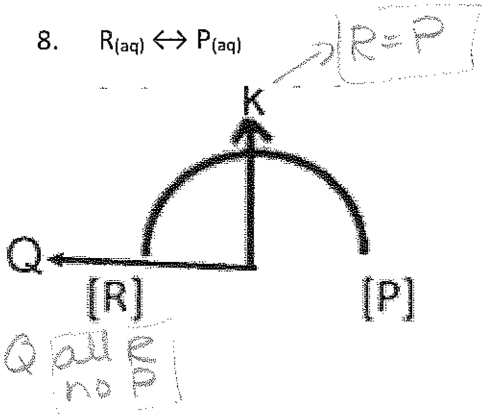
5. Write the equilibrium expression for this reaction.
6. Determine Kc for this process.
7. Draw a graphical representation of concentration.

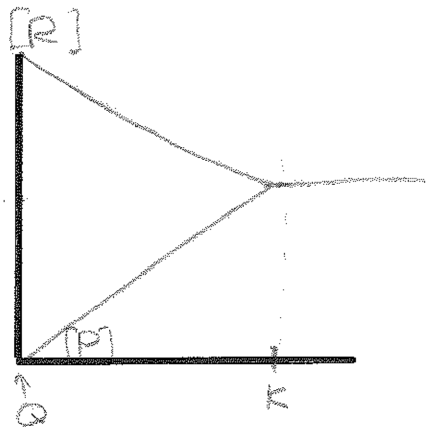
$A + 2B \leftrightarrow 2C$

	A	2B	2C
I	0.5	0.5	
S	-0.1	-0.2	+0.2
E	0.4	0.3	0.2



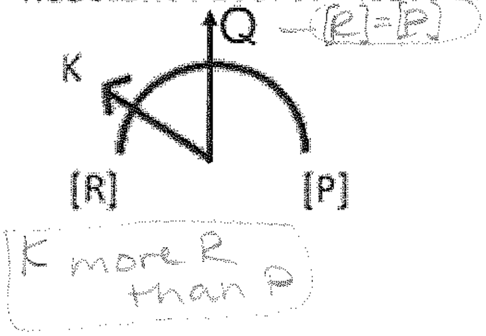
8. $R_{(aq)} \leftrightarrow P_{(aq)}$

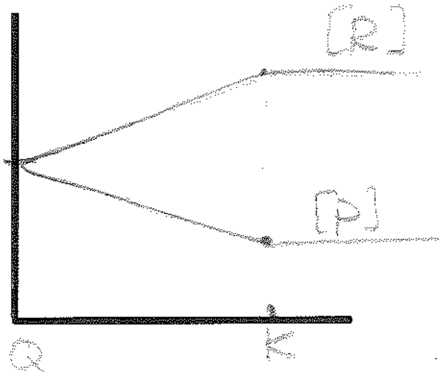




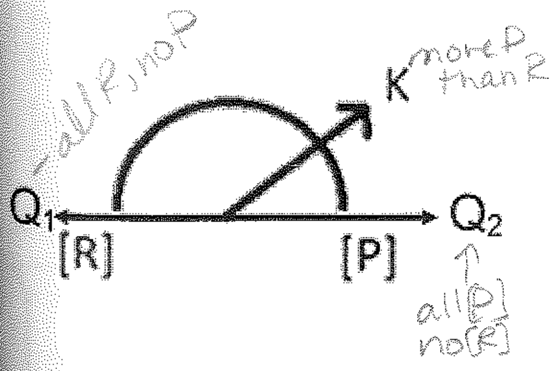
9. $R_{(aq)} \leftrightarrow P_{(aq)}$

Reactant Favored Reaction

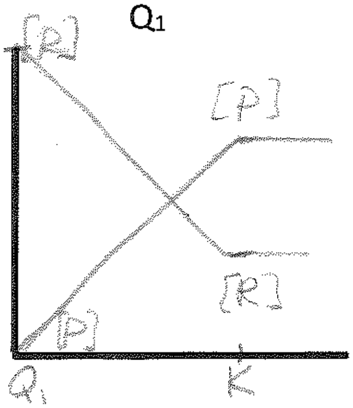




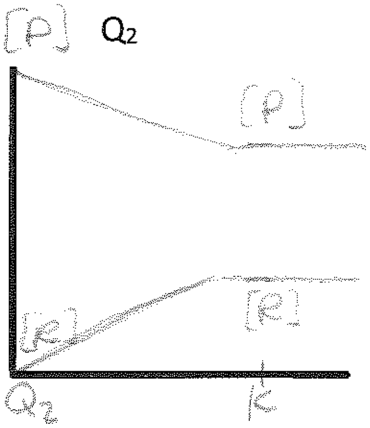
10. $R_{(aq)} \leftrightarrow P_{(aq)}$



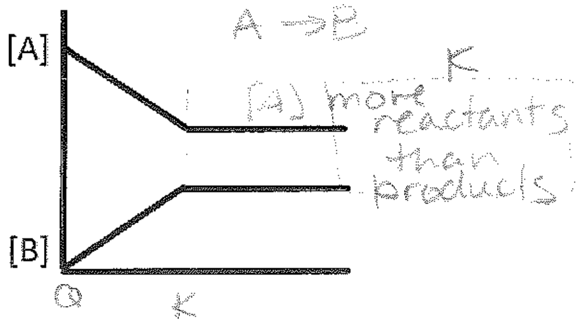
Q_1

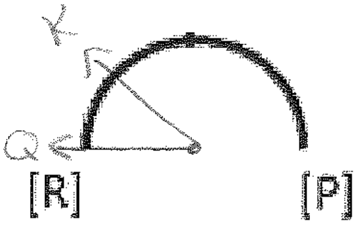


Q_2



11. Using the graph draw an arrow on the gauge indicating placement of equilibrium ratios. Draw an arrow on the graph to the right for both reaction Quotient and K. [A ↔ B]





(#11-1)

Can I determine the extent (favorability) of a chemical reaction?

Write the equilibrium expression for the following chemical reactions.

1. For each of the following, provide a Kc and a Kp expression.

   a. $A(aq) + B(s) \leftrightarrow C(g)$ $K_c = \frac{[C]}{[A]}$ $K_p = (P_C)$

   b. $2A(s) + B(g) \leftrightarrow 2C(g)$ $K_c = \frac{[C]^2}{[B]}$ $K_p = \frac{(P_C)^2}{(P_B)}$

   c. $A(l) + B(aq) \leftrightarrow 3C(aq)$ $K_c = \frac{[C]^3}{[B]}$ no Kp → no gas

   d. $Ag_2SO_4(s) \leftrightarrow 2Ag^+_{(aq)} + SO_4^{-2}{}_{(aq)}$ (Balance me please) Solubility → Ksp

      $K_c = K_{sp} = [Ag^+]^2[SO_4^{-2}]$

   e. $PbSO_4(s) \leftrightarrow Pb^{+2} + SO_4^{-2}$ (provide balanced dissolving equation) no Kp no gas

      $K_c = K_{sp} = [Pb^{+2}][SO_4^{-2}]$

Reaction Extent

2. A reaction that goes to greatly to extent produces a lot of ([✓ products]/reactants) means "towards products"

3. A reaction that is product favored will have a ([✓ large K]/small K)

   $K = \frac{[P]}{[R]}$

4. Question 1d above has a K value = 1.0E-5 and 1e has a K value = 6.3E-7.

   0.00001 0.00000063

   a. Which reaction will produce more product at equilibrium?

      1d with larger K will have more product

   b. Is the answer from letter a considered product favored? no

      K is less than 1 and is reactant favored

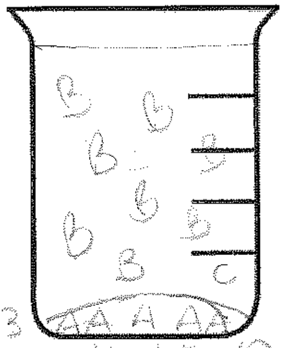
5. Reaction $A(s) + B(aq) \leftrightarrow 2C(aq)$

   a. Top beaker has a K = 1.0 E-10 reactant favored

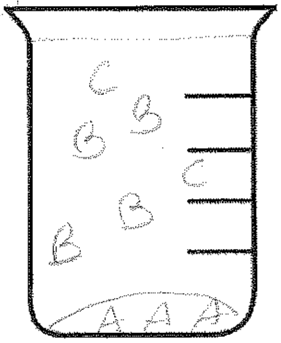
   b. Bottom beaker has a k = 1.0 Products : Reactants

Draw the beaker at equilibrium

$K = \frac{[C]^2}{[B]} = \frac{2^2}{4} = 1$



K = 1 E-10



K = 1

Quiz 1 topic Reminder

I understand the value of K and can generate an equilibrium expression

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(#11-1) I understand the value of K and can generate an equilibrium expression

(#11-1a)

- I can write an equilibrium constant expression Kc or Keq
- I can write an equilibrium constant expression Kp (partial pressures.)
- I can write an equilibrium constant expression for solubility Ksp.

(#11-1b)

- I can explain how the value of the equilibrium constant (k) relates to the extent of the reaction.

(#11-1c)

- I can calculate a new K value from an old K value if the reaction is altered (Altered means reversed or coefficients are a multiple of the original.)

(#11-1d)

- I can specifically model an equilibrium system using particulate diagrams.
- I can determine if a reaction is at equilibrium.

$$A(aq) + B(s) \leftrightarrow 2C(g)$$

$2C(g) \rightarrow A(aq)$

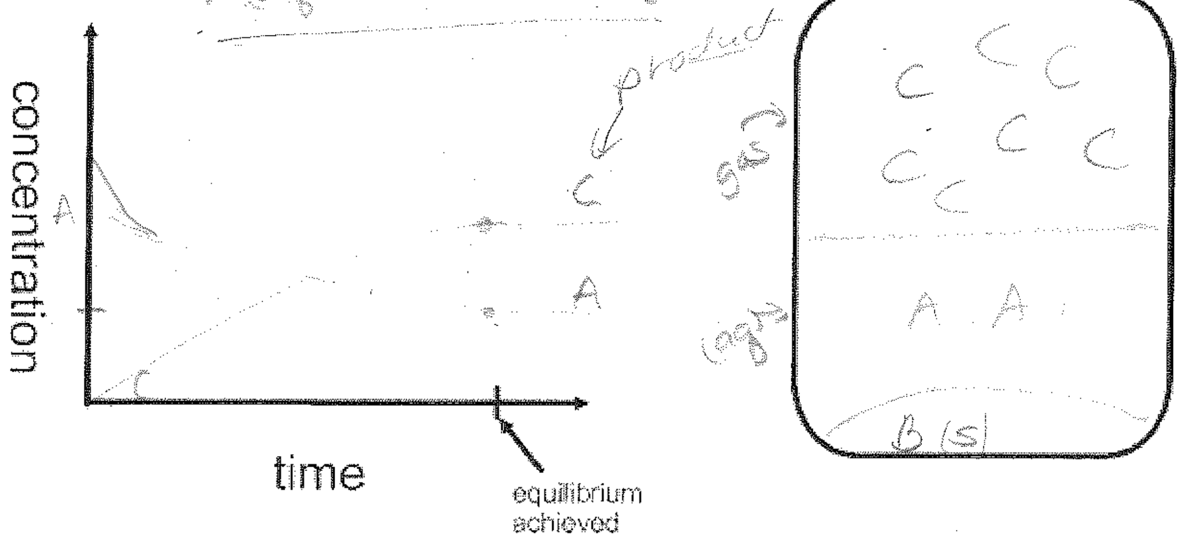
$Kp = \frac{1}{(P_c)^2}$

1. Write the equilibrium expression
   a. Kc = $\frac{[C]^2}{A}$
   b. Kp = $(P_c)^2$

2. In a 2L ridged vessel, 0.2M A and solid B are added. Since the reaction is slightly product favored, draw a graph below and ridged container representing the reaction at equilibrium.

3.

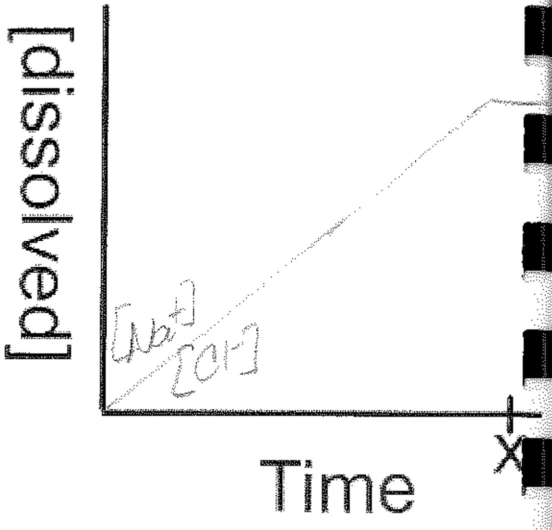
$A(aq) + B(s) \rightleftharpoons 2C(g)$



(#11-2 #11-1)

Modeling and determining equilibrium of solubility

You take a sample of salt (table) and you start to dissolve it in water. You continue to add and stir the solution. After a period of time it appears solid is forming on the bottom. (saturated)



a. Write the solubility equation.

$NaCl_{(s)} \rightleftharpoons Na^{+1} + Cl^{-1}$

b. This chemical reaction goes to (completion/[✓ equilibrium]).

c. To the right draw a picture of table salt dissolving. The reaction gets saturated at time X. Graph →

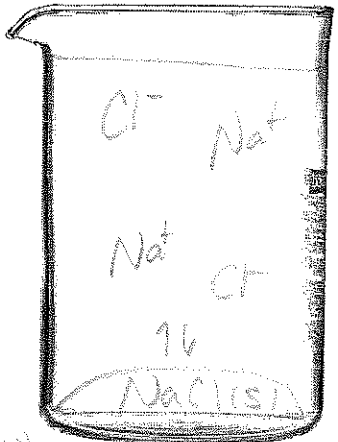
d. A student hypothesizes that stirring increases solubility. Confirm or counter this statement. (Counter)

Stirring increases the rate in which equilibrium is achieved.

e. How might you increase the solubility of table salt in water.

increase the temperature

f. Draw the picture of table salt in the beaker to the right at time X.



A student dissolved Ag_2SO_4 in 100mL of water. The student added 5g of silver sulfate to the solution and measured the concentration of SO_4^{-2} to be 0.0135M. Answer the following questions.

$5g \cdot \frac{1 mol}{311.95} = 0.0160$

Ag 107.9 × 2 = 215.8
S 32.1 = 32.1
O 16 × 4 = 64
} 311.9 g/mol

a. Draw the beaker.

b. Write out the dissolving equation and the equilibrium expression.

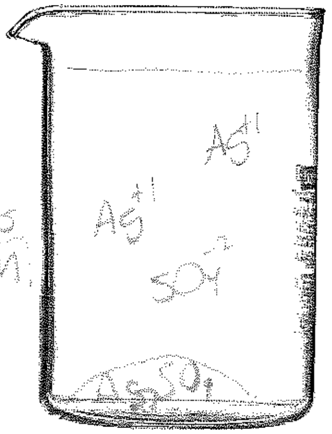
$Ag_2SO_{4(s)} \rightleftharpoons 2Ag^{+1} + SO_4^{-2}$ $K_{sp} = [Ag^{+1}]^2[SO_4^{-2}]$

c. Based on your picture, if the $SO4^{-2}$ = 0.0135, what is the [Ag^+] ?

$[Ag^{+1}]$ will be 2x conc. of SO_4^{-2}

$[Ag^{+1}] = 2 \cdot 0.0135$
$\boxed{= 0.027M}$

d. Create an ISE table, Fill it out and determine the K value.



	$Ag_2SO_{4(s)} \rightleftharpoons$	$2Ag^{+1}$	$+ SO_4^{-2}$
I	0.016		
S	-0.00135	+0.0027	+0.00135
E	0.01465	0.0027	0.00135

we don't need this for Ksp →

$K_{sp} = [0.027M]^2 \cdot [0.0135M]$
$K_{sp} = 9.84E-6$

SO_4^{-2}
$0.0135M = \frac{X}{0.1L}$
$X = 0.00135 \text{ mol } SO_4^{-2}$

Ag^{+1}
$0.027M = \frac{X}{0.1L}$
$X = 0.0027 \text{ mol } Ag^{+1}$

Two salts are dissolving, AgBr (Ksp = 5E-13) and AgCl (Ksp = 1.6E-10), Answer the following questions.

smaller   0.00000000016

a. Write out the dissolving equation for each.

$AgBr_{(s)} \rightleftharpoons Ag^{+1}_{(aq)} + Br^{-1}_{(aq)}$ $AgCl_{(s)} \rightleftharpoons Ag^{+1}_{(aq)} + Cl^{-1}_{(aq)}$

b. Write out the solubility expression.

$K_{sp} = [Ag^{+1}][Br^{-1}]$ $K_{sp} = [Ag^{+1}][Cl^{-1}]$

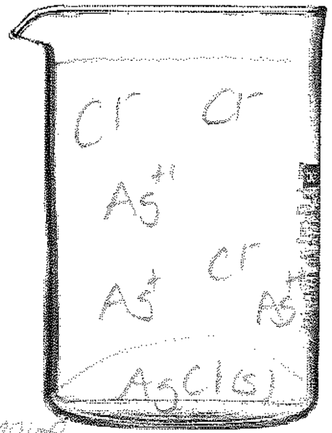
c. A large K value means what relative to solubility?

a large K mean the salt dissolve to a great extent.

d. Which of the two salts is more soluble?

AgCl has a larger K, so AgCl dissolves more

e. Draw a saturated solution of the *more* soluble salt.



A student comes across a solution that is saturated solution of lead(II) chloride. The lead ion has a concentration of 1.5E-5M. Answer the following questions.

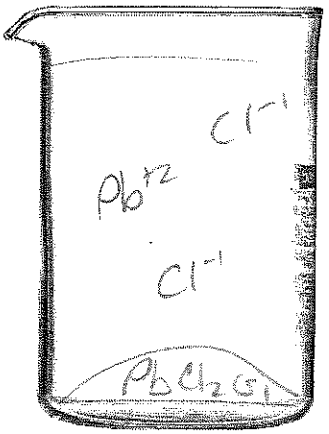
a. Write out the solubility reaction.

$PbCl_{2(s)} \rightleftharpoons Pb^{+2}_{(aq)} + 2Cl^{-1}_{(aq)}$

b. Write out the solubility equilibrium expression.

$K_{sp} = [Pb^{+2}][Cl^{-1}]^2$

c. Draw a picture of this reaction mixture.



d. What is the concentration of the Chloride ion?

3.0E-5 M

e. What is the equilibrium constant for this reaction?

$PbCl_2 \rightleftharpoons Pb^{+2} + 2Cl^{-1}$

S   1.5E-5   3.0E-5

$K_{sp} = (1.5E5)(3.0E{-5})^2 = 1.35E^{-14}$

(#11-2)

Equilibrium: Concentration Determination

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1. (brown574) For the Haber process, $N_{2(g)} + 3H_2 \leftrightarrow 2NH_{3(g)}$ $K_p = 1.45E\text{-}5$ @ 500°C. In an equilibrium mixture of the three gases, the partial pressure of H_2 is 0.928atm and that of N_2 is 0.432 atm.
   a. Write the equilibrium expression for the reaction. $K_p = \frac{(P_{NH_3})^2}{(P_{N_2})(P_{H_2})^3}$

   b. What is the partial pressure of NH_3 in this equilibrium mixture?

$$1.45E\text{-}5 = \frac{(P_{NH_3})^2}{(0.432)(0.928)^3} \qquad (P_{NH_3})^2 = 5.0E\text{-}6$$

$$\boxed{P_{NH_3} = 0.0022 \text{ atm}}$$

2. Nitrogen gas is mixed with oxygen gas to ~~from~~ form nitrogen monoxide.

$$N_2 + O_2 \leftrightarrow 2NO \quad Kc\ 4.1\ E\text{-}4$$

In a 2L rigid tank, 0.5mol of $N_2$is mixed with 0.86 mol of O_2 gas at 2000.K. The two gases react forming an equilibrium.

a. Write out the equilibrium expression. $K_c = \frac{[NO]^2}{[N_2][O_2]}$

	N_2 +	O_2 →	$2NO$	
I	0.5	0.86		completion
S	-0.5	-0.5	+1.0	
E	0	0.36	1.0	

b. If this reaction were to go to completion, what is the value of "x"?

(x is the shift -- the Stoich Change

It goes to completion $x = 0.5$ mol

	N_2 +	O_2 →	NO	
I	0.5	0.86		Equilibrium
S	-x	-x	+2x	
E	0.5-x	0.86-x	2x	

c. Based upon the size of K the value of "x" is (big, small, very small)?

very small x

K is small – reactant favored

d. Determine the final concentrations of each species at equilibrium.

$[N_2] = \frac{0.5}{2L} = 0.25M$ N_2

$[O_2] = \frac{0.86}{2L} = 0.43M$

$$4.1E\text{-}4\ \frac{(2x)^2}{(0.25-x)(0.43-x)}$$

negligible

$$4.1E\text{-}4 = \frac{4x^2}{(0.25)(0.43)} \qquad x^2 = 1.1E\text{-}5 \qquad x = 0.0033$$

$[N_2] = 0.25M$

$[O_2] = 0.43M$

$[NO] = 2x = 0.0066M$

3. The reversible reaction $CH_{4(g)} + H_2O_{(g)} \leftrightarrow CO_{(g)} + 3H_{2(g)}$ has been used as a commercial source of hydrogen. At 1500°C, an equilibrium mixture of these gases was found to have the following concentrations: [CO] = 0.300M, $[H_2]$ = 0.800M, $[CH_4]$ = 0.400M. At 1500°C Kc = 5.67 for this reaction.
   a. Write the equilibrium expression. $K_c = \frac{[CO][H_2]^3}{[CH_4][H_2O]}$

   b. If this reaction went to completion, how big would "x" be? "x" is the variable in the ISE table.

Depends on what you start with

$[H_2O] = \frac{(0.3)(0.8)^3}{(0.4)(5.67)}$

   c. Does this reaction go to equilibrium?

at eq in problem so yes

$[H_2O] = 0.068$

   d. Approximate the size of "x", can this value be ignored in the ISE table (use the short cut rule).

   e. Solve for "x", using the quadratic if you can. Otherwise set it up for the quadratic.
   What was the equilibrium concentration of $H_2O_{(g)}$ in this mixture?

$CH_4 + H_2O \rightleftarrows CO + 3H_2$

I				
S	x		x	3x
E	0.400	?	0.300	0.800
		0.068		

$$\frac{(0.3)(0.8)^3}{(0.4)[H_2O]} = 5.67$$

$$[H_2O] = \frac{(0.3)(0.8)^3}{(0.4)}$$

(#11-2)

Equilibrium: K Determination

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1. (brown570) A mixture of Hydrogen and nitrogen in a reaction vessel is allowed to attain equilibrium at 472°C. The equilibrium mixture of gases was analyzed and found to contain 0.1207M H_2, 0.0402M N_2, and 0.00272 MNH_3. From this data calculate the equilibrium constant Kc for $N_{2(g)} + 3H_{2(g)} \leftrightarrow 2NH_{3(g)}$

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(0.00272)^2}{(0.0402)(0.1207)^3} = 0.105$$

2. (brown 571)A mixture of 5.000x 10^{-3} mol of H_2 and 1.000 x 10^{-2} mol of I_2 is placed in 5.000L container at 448°C and allowed to come to equilibrium. Analysis of the equilibrium mixture shows that the concentration of HI is 1.87E-3M. Calculate K_c at 448°C for the reaction

$$H_{2(g)} + I_{2(g)} \leftrightarrow 2HI_{(g)}$$

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

I 5E-3 1E-3

S -x -x +2x

E 5E-3-x 1E-3-x 1.87E-3

2x = 1.87E-3

x = 0.000935

$[H_2]$ = 5E-3 - 0.000935 = 0.0041

$[I_2]$ = 1E-2 - 0.000935 = 0.0091

$$K_c = \frac{[1.87E-3]^2}{[9.1E-3][9.1E-3]} = \boxed{0.094}$$

3. A reaction mixture of three gases, A, B, and C are all 1.0 M at 200K. The reaction below runs for a period of time and forms an equilibrium balance where a little solid formed on the bottom. The concentration of A at equilibrium is .5M.

(b) $2A_{(g)} + B_{(g)} \leftrightarrow C_{(g)} + D_{(s)}$

I 1 1 1

S -2x -x +x

E 1-2x 1-x x

↑ =0.5M 0.75 1.25

1-2x = 0.5

2x = 0.5

x = 0.25

   a. Write the equilibrium expression.
   b. Complete an ISE table below the equation.
   c. Determine K.

(a) $K_c = \frac{[C]}{[A]^2[B]}$

(c) $K = \frac{[1.25]}{[0.5]^2(0.75)} = \boxed{6.67}$

4. The following reaction goes to equilibrium at 500K. $A_{(g)} \leftrightarrow 2B_{(g)} + C_{(s)}$
Original pressure of A is .55atm and reduces to .15atm at equilibrium.
   a. Write the Kp equilibrium expression.

$$K_p = \frac{(P_B)^2}{(A)}$$

I 0.55

S -x +2x

E 0.15 2x

x = 0.4

B = 2x = 0.8

   b. Complete an ICE table below the reaction.

   c. Determine Kp. $K_p = \frac{(0.8)^2}{(0.15)} = 4.3$

   d. (Challenge) If the container is 1L in size, determine the mass of C (carbon) produced at equilibrium?

B : 0.8atm

use PV = nRT → det. mol of B

$n = \frac{PV}{RT} = \frac{(0.8atm)(1L)}{(0.0821)(500K)} = 0.019$ mol B

$0.019 \text{ mol B} \times \frac{1 \text{ mol C}}{2 \text{ mol B}} \times \frac{12g}{1 \text{ mol C}} = 0.12g$ C

Reaction Quotient

How do you predict which direction a reaction will proceed to reach equilibrium?

Why?

When a reaction reaches equilibrium there must be some non-negligible amount of every species in the reaction, otherwise the reaction cannot react in both directions. Knowing this, it is very easy to predict which direction a reaction will go to reach equilibrium when one of the components of the reaction has an initial concentration of zero. Many of the problems you have worked with thus far have some component at zero concentration, but real life does not work that way. Most of the time, the reaction in question has some measureable amount of every species. Deciding which way the reaction will go to reach equilibrium then becomes more challenging.

Model 1 – A Theoretical Equilibrium

Trial 1	A(g)	+ B(g) ⟷	C(g)
Initial	1.000 M	1.000 M	1.000 M
Change		←	
Equilibrium	1.464 M	1.464 M	0.536 M

Trial 2	A(g)	+ B(g) ⟷	C(g)
Initial	2.000 M	0.500 M	0.500 M
Change		←	
Equilibrium	2.150 M	0.650 M	0.350 M

Trial 3	A(g)	+ B(g) ⟷	C(g)
Initial	1.000 M	0.500 M	1.500 M
Change		←	
Equilibrium	1.864 M	1.364 M	0.636 M

Trial 4	A(g)	+ B(g) ⟷	C(g)
Initial	1.600 M	1.000 M	0.400 M
Change		no shift	
Equilibrium	1.600 M	1.000 M	0.400 M

Trial 5	A(g)	+ B(g) ⟷	C(g)
Initial	1.400 M	1.200 M	0.400 M
Change		→	
Equilibrium	1.388 M	1.188 M	0.412 M

Trial 6	A(g)	+ B(g) ⟷	C(g)
Initial	0.750 M	2.000 M	0.250 M
Change		→	
Equilibrium	0.675 M	1.925 M	0.325 M

1. Examine Model 1.
   a. Write the theoretical chemical reaction that is used in the trials of Model 1.

   $A_{(g)} + B_{(g)} \rightleftharpoons C_{(g)}$

   b. If 0.50 M of reactant A reacts, predict the change in concentration of B and C.

   Same stoichiometric change - all coefficients are 1
   0.5M B reacts or created and 0.5M C is created or +

2. What variables were changed in the different trials shown in Model 1?

   Initial amount of A, B, and C

3. In Trial 1 of Model 1 there is an arrow in the "change" section of the table.

   a. Explain what that arrow represents.

   The equilibrium is shifting toward the reactants

   b. What evidence is present in the table to indicate the direction the arrow should be pointing?

   Look at how amounts change from I to E. In trial 1, the amount of reactants increase and the amount of products decrease

4. With your group, determine which direction each of the other trials in Model 1 reacted to reach equilibrium. Indicate that direction with an arrow in the "change" section of the table. see model 1

5. Is it true that there are equal concentrations of reactants and products when all of the reactions in Model 1 reach equilibrium? Justify your answer with evidence from Model 1.

   no reactants do not = products at equilibrium

   for example Trial 1 reactact are 1.464 each + products are 0.536

6. According to Model 1, are the final concentrations of all species in the reaction the same when the reaction reaches equilibrium, regardless of the initial concentration?

   no.

7. Does the reaction in Model 1 always proceed in the forward direction when there are more reactants than products? Justify your answer with evidence from Model 1.

   no → trial 4 - - no shift

   → trial 2 - more A + B thus proceed in reverse direction

8. Write the equilibrium constant expression for the reaction in Model 1.

$$K_{eq} = \frac{[C]}{[A][B]}$$

9. Discuss with your group how you could determine the equilibrium constant, K_{eq}, for the reaction in Model 1. Divide the work among group members. Use data from multiple trials to calculate the equilibrium constant for the reaction and determine the average. Show all work.

Trial 1 $\frac{(.536)}{(1.464)(1.464)} = 0.25$

Trial 2 $\frac{(0.350)}{(2.150)(0.650)} = 0.25$

Trial 3 $\frac{(0.636)}{(0.864)(1.364)} = 0.25$

Trial 4 $\frac{(0.4)}{(1.6)(1.0)} = 0.25$

Trial 5 $\frac{(0.412)}{(1.388)(1.188)} = 0.25$

Trial 6 $\frac{(0.325)}{(0.675)(1.925)} = 0.25$

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Read This!

The key to knowing which direction a reaction will need to proceed in order to reach equilibrium is knowing if you have too much reactant or too much product compared to the equilibrium state. Keep in mind, however, that there are many combinations of reactant and product concentrations that constitute an equilibrium state.

Model 2 – Comparing Q and K_{eq}

	⑫	⑩	⑮	⑪
Trial	**Reaction Quotient, Q**	**Equilibrium Constant, K_{eq}**	**Q versus K_{eq}**	**Direction to Equilibrium**
1	1	0.25	$Q > K_{eq}$	⟵
2	0.5	0.25	$Q > K_{eq}$	⟵
3	3	0.25	$Q > K_{eq}$	⟵
4	0.25	0.25	$Q = K_{eq}$	no shift
5	0.238	0.25	$Q < K_{eq}$	⟶
6	0.167	0.25	$Q < K_{eq}$	⟶

10. Fill in the Equilibrium Constant column in Model 2 using data from Model 1. (from #9)
11. Fill in the Direction to Equilibrium column in Model 2 using data from Model 1.

Read This!

The **reaction quotient** for a reaction is the ratio of products to reactants, similar to the equilibrium constant. The difference is you calculate the ratio with initial conditions.

$$K_{eq} = \frac{[C]_{eq}}{[A]_{eq}[B]_{eq}} \qquad Q = \frac{[C]_{initial}}{[A]_{initial}[B]_{initial}}$$

12. Calculate the reaction quotient for each of the trials in Model 1 and record the data in Model 2 in the appropriate column. Divide the work among group members. Show your work below.

Trial 1 $\frac{(1)}{(1)(1)} = 1$

Trial 2 $\frac{(0.5)}{(2)(0.5)} = 0.5$

Trial 3 $\frac{(1.5)}{(1.0)(0.5)} = 3$

Trial 4 $\frac{(0.4)}{(1.6)(1.0)} = 0.25$

Trial 5 $\frac{0.4}{(1.4)(1.2)} = 0.238$

Trial 6 $\frac{(0.25)}{(0.75)(2.0)} = 0.167$

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13. Consider how the concentration values in the reaction quotient change when the reaction proceeds in the forward direction. →

   a. Does the numerator increase or decrease? →

   (top number is product)

   Product increase

   b. Does the denominator increase or decrease?

   (bottom number-reactants)

   Reactants decrease

   c. Overall, does the Q ratio increase or decrease when the reaction proceeds in the forward direction? Q ratio increases

   Trial 6 $Q = 0.167$ $K_{eq} = 0.25$

14. Consider how the concentration values in the reaction quotient change when the reaction proceeds in the reverse direction. ←

   a. Does the numerator increase or decrease?

   products decrease

   b. Does the denominator increase or decrease?

   reactants increase

   c. Overall does the Q ratio increase or decrease when the reaction proceeds in the forward direction?

   Q ratio decreases

   Trial 1 $Q = 1$, $K_{eq} = 0.25$

15. Fill in the Q versus K_{eq} column in the table in Model 2. Write $Q > K_{eq}$, $Q < K_{eq}$ or $Q = K_{eq}$. see model 2

16. Complete the following statements.

   a. When the reaction quotient is **greater than** the equilibrium constant, the reaction proceeds more in (the forward, [✓ the reverse], neither) direction to reach equilibrium.

   b. When the reaction quotient is **less than** the equilibrium constant, the reaction proceeds more in ([✓ the forward], the reverse, neither) direction to reach equilibrium.

   c. When the reaction quotient is **equal to** the equilibrium constant, the reaction proceeds more in (the forward, the reverse, [✓ neither]) direction to reach equilibrium.

17. Consider the following reaction.

$$2SO_2(g) + O_2(g) \longleftrightarrow 2SO_3(g) \qquad K_{eq} = 3900 \text{ at } 2000 \text{ K}$$

a. Write the equilibrium constant expression for the reaction.

$$Keq = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

b. Write the reaction quotient expression for the reaction.

$$Q = \frac{[SO_3]^2_{init}}{[SO_2]^2_{init}[O_2]_{init}}$$

c. A reaction vessel contains 0.150 M sulfur dioxide, 0.150 M oxygen and 2.000 M sulfur trioxide. Predict the direction the reaction must shift to reach equilibrium. Show a calculation to justify your answer.

$$Q = \frac{(0.50)^2}{(2)^2(0.15)} = 0.417 \qquad Keq = 3900$$

$Q < Keq$ so reaction will continue in the forward direction

18. Consider the following reaction.

$$2HI(g) \longleftrightarrow H_2(g) + I_2(g) \qquad K_{eq} = 0.25 \text{ at } 25\,°C$$

A reaction vessel contains 0.500 M hydrogen, 0.500 M iodine vapor and 0.750 M hydrogen iodide. Predict the direction the reaction must shift to react equilibrium. Show a calculation to justify your answer.

$$Q = \frac{[HI]^2}{[H_2][I]} = \frac{(0.75)^2}{(0.5)(0.5)} = 2.25 \qquad Keq = 0.25$$

$Q > Keq$ so reaction will shift in the reverse direction

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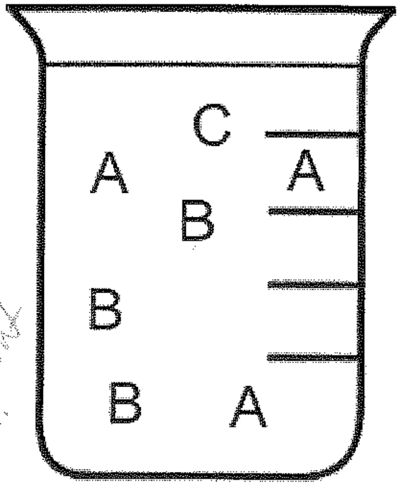
(#11-3)

Where am I and where am I going?

$A(aq) + B(aq) \rightleftharpoons C(ac$



1. The beaker below is ~~an~~ at equilibrium ~~expression~~. Answer the following questions. Reaction is being run at 25C.
   a. Write the equilibrium expression for this reaction.

      $K = \frac{[C]}{[A][B]}$

   b. Is the K value for this substance bigger or smaller than 1?

      $K = \frac{1}{(3)(3)} = \frac{1}{9}$ smaller than 1

   c. Would you consider this reaction (product favored/reactant favored). reactant favored

      but reactant f'd.

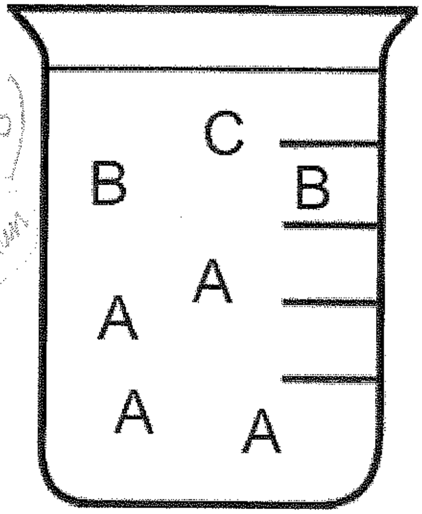
   d. Each letter will represent 0.1M, what is the value of K?

      $K = \frac{(0.1)}{(0.3)(0.3)} = 1.11$

   e. What is the only way to change the equilibrium constant?

      change the temperature

$A(aq) + B(aq) \rightleftharpoons C(a$



2. To the right is another beaker with the _same reaction as above_ at 25C.
   a. What is the K for this reaction?

      $K = 1.11$ (same reaction at same temp so K is same)

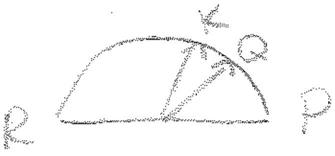
   b. Each letter represents a 0.1M concentration, Determine the reaction quotient or this reaction.

      $Q = \frac{0.1}{(0.4)(0.2)} = 1.25$ not at equilibrium

   c. Is this reaction
      i. At equilibrium? no
      ii. Too much reactant? no
      [✓ iii. Too much product?]

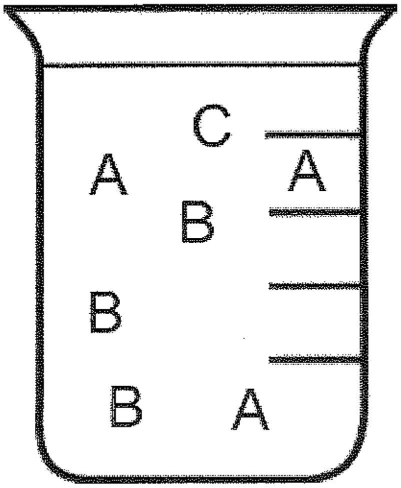
      $Q > K$

   d. Given time all reactions will achieve equilibrium. To achieve equilibrium this reaction will
      i. Stay where it is.
      [✓ ii. Lo~~o~~se product and gain reactant]
      iii. Lo~~o~~se reactant and gain product.



3. To the right is another beaker running the same reaction at the same temperature.

$A(aq) + B(aq) \rightleftharpoons C(a$



   a. What is the K value for this reaction?

   K=1.11 (same reaction, same temp. so same K)

   b. What is the reaction quotient for the reaction to the right? (each letter represents 0.1M)

   $K = \frac{0.1}{(0.3)(0.3)} = 1.11$

   c. Is this reaction
      [✓ i. At equilibrium?]
      ii. Too much reactant?
      iii. Too much product?

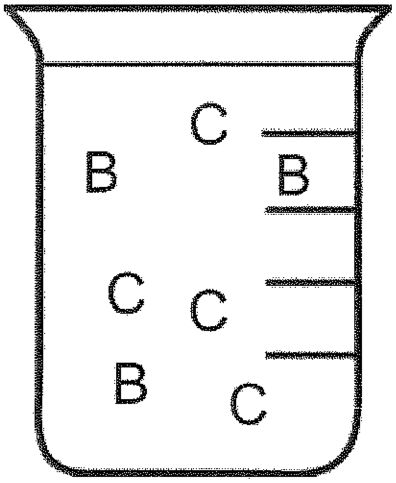
   d. Given time all reactions will achieve equilibrium. To achieve equilibrium this reaction will
      [✓ i. Stay where it is.]
      ii. Lo~~o~~se product and gain reactant
      iii. Lo~~o~~se reactant and gain product.

4. To the right is another beaker running the same reaction at the same temperature.

   a. What is the K value for this reaction?

   K=1.11 (same reaction same temp so same K)

$A(aq) + B(aq) \rightleftharpoons C(aq$



   b. What is the reaction quotient for the reaction to the right? (each letter represents 0.1M)

   $Q = \frac{0.4}{1(0.3)} = 1.3$

   c. Is this reaction
      i. At equilibrium?
      ii. Too much reactant?
      [✓ iii. Too much product?]

   d. Given time all reactions will achieve equilibrium. To achieve equilibrium this reaction will
      i. Stay where it is.
      [✓ ii. Lo~~o~~se product and gain reactant]
      iii. Lo~~o~~se reactant and gain product.

5. What can change if we alter the temperature?

   K will change

Q3 #11-3

I know how a reaction will proceed when approaching equilibrium?

(#11-3) I know how a reaction will proceed when approaching equilibrium?

a. (#11-3a1) I can solve for a reaction quotient.
b. (#11-3a2) I can relate reaction quotient to equilibrium constant to judge how a reaction will proceed to get to equilibrium. (Q vs. K) (*note: will a precipitate form?*)
c. (#11-3b1) I can determine if a stress (change) actually alters the equilibrium position.
d. (#11-3b2) I can determine how the reaction will change if the reaction mixture was altered from equilibrium to get back to equilibrium (Le Chatelier's principle)

Practice question

$$A(aq) + B(s) \rightarrow 2C(aq) \quad \boxed{K = 10}$$

0.5M excess 0.5m

In a beaker a 0.5M [A] is placed with excess B and 0.5M [C] and allowed to react.

a. What is the reaction quotient? $Q = \frac{(0.5)^2}{(0.5)} = 0.5$
b. How will the value of Q change as the reaction approaches K? Q will increase
c. What will happen the concentration of each chemical as the reaction shifts to equilibrium?
   (place an arrow up or down) [A] ↓ [B] ↓ [C] ↑

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons FeSCN^{2+}(aq)$$

(colourless) (red)

This reaction is at equilibrium and the beaker appears mostly clear. (mostly reactants)

a. What can you say about the relative value of K? K is small

b. Fe^{3+} is added to the solution causing the solution to turn slightly red. Why?
Equilibrium (K) must remain the same so reaction shifts forward, creating red ions (product)

c. In which direction will the reaction shift based upon the following stress?
   i. Addition of SCN^- →
   ii. Removal of $FeSCN^{2+}$ →
   iii. Addition of solid iron. no shift
   iv. Which of these will cause a change in K? no change in K unless T is changed

d. Addition of heat causes the solution to turn from clear to red. →
   i. Is this reaction exothermic or [✓ endothermic?]
   ii. What happens to the size of K? K increases (more product with new K)

4. (Brady655) Hydrogen, a potential fuel, is found in great abundance in water. Before the hydrogen can be used as a fuel, however, it must be separated from the oxygen; the water must be split into H_2 and O_2. One possibility is thermal decomposition, but this requires very high temperatures. Even at 1000°C, Kc = 7.3E-18 for this reaction: $2H_2O_{(g)} \leftrightarrow 2H_{2(g)} + O_{2(g)}$

a. If a vessel has a concentration of 0.10M of every species, what is the reaction quotient?

$Q = \frac{(0.1)^2(0.1)}{(0.1)^2} = 0.1$

b. Which way will the reaction proceed as it approaches equilibrium?

$Q > K_{eq}$ shift ← (reverse reaction)

c. If at 1000°C the H_2O concentration in a reaction vessel is set initially at 0.100M what will the $\boxed{H_2}$ concentration be when the reaction reaches equilibrium.

$2H_2 \rightleftharpoons 2H_2 + O_2$

I 0.1

S $-2X$ $+2X$ X

E $0.1-2X$ $2X$ X

$7.3E18 = \frac{(2x)^2(x)}{(0.1-2x)^2}$ neg

$\frac{7.3E{-20}}{4} = \frac{4x^3}{4}$

$\sqrt[3]{x^3} = \sqrt[3]{1.825E{-20}}$

$x = 2.6E{-7}$

$[H_2] = 2x = 2(2.76E{-7}) = 5.2E{-7}$

5. A mixture consisting initially of 3.00 moles NH_3, 2.00 moles of N_2, and 5.00 moles of H_2, in a 5.00 L container was heated to 900 K, and allowed to reach equilibrium.

$$2\ NH_3(g) \leftrightarrows N_2(g) + 3\ H_2(g) \quad K_c = 0.0076 \text{ @ } 900\ K$$

a. Write out the equilibrium expression. $K_c = \frac{[H_2]^3[N_2]}{[NH_3]^2}$

b. What is the value of Q? $Q = \frac{(5)^3(2)}{(3)^2} = 27.8$ $K = 0.0076$

c. Which way will this reaction proceed to approach K?

$Q > K$ shift ←

Reaction will proceed in the reverse direction

6.

(#11-3)

Honors Chemistry

Le Chatelier's Principle

Match the change to the equilibrium system below with the letter of the appropriate response. Each letter can be used once, more than once, or not at all.

$$2SO_2(g) + O_2(g) \Leftrightarrow 2SO_3(g) + \text{heat}$$

a → 1. O_2 is added to the system. → a. The reaction shifts to the right.

b ← 2. SO_3 is added to the system. b. The reaction shifts to the left.

b ← 3. The temperature of the system is increased. c. No Shift.

no shift 4. A catalyst is added to the system.

a → 5. The volume is decreased.

⇒ pressure is increased

If the statement is true, write "true." If it is false, change the underlined word or words to make the statement true. Write your answer on the line provided.

$$NH_4Cl(s) + \text{heat} \Leftrightarrow NH_3(g) + HCl(g)$$

false 5. The above reaction is exothermic. endothermic

true 6. The production of ammonia from ammonium chloride will increase at higher temperature. →

true 7. For the above reaction at equilibrium, an increase in the concentration of HCl causes a decrease in gaseous ammonia concentration. ←

8. The following equilibrium may be established with carbon dioxide and steam.

$$CO(g) + H_2O(g) \leftrightarrow CO_2(g) + H_2(g) + \text{heat}$$

go to p. 79 for questions on this one

(#11-3)
Equilibrium: Le Chatelier's principle

endothermic

1.(brown580) Consider the following equilibrium $N_2O_{4(g)} \leftrightarrow 2NO_{2(g)}$ $\Delta H = 58.0$ kJ/mol In what direction will the equilibrium shift when each of the following changes is made to a system at equilibrium

a. add N_2O_4 $\rightarrow$
b. remove NO_2 $\rightarrow$
c. increase the total pressure by adding N_2 no shift
d. increase the volume $\rightarrow$
e. decrease the temperature. $\leftarrow$

2.(Brady640) The air pollutant sulfur dioxide can be removed form a gas mixture by passing the gases over calcium oxide. The equation is $CaO_{(s)} + SO_{2(g)} \leftrightarrow CaSO_{3(s)}$ If the reaction is currently at equilibrium which means it looks as if the reaction has essentially ended… how will these alterations affect the direction the reaction will shift to return to equilibrium?

a. Addition of $CaO_{(s)}$ - solid - no shift
b. Addition of $SO_{2(g)}$ $\rightarrow$
c. Addition of a catalyst? no shift
d. Addition of additional O_2? no shift

3. (brady645) Consider the equilibrium $PCl_{3(g)} + Cl_{2(g)} \leftrightarrow PCl_{5(g)}$ + energy , for which $\Delta H° = -88$kJ. How will the amount of Cl_2 at equilibrium be affected by a) adding PCl_3 b) adding PCl_5 c) raising the temperature, and d) decreasing the volume of the container? E. How will all of these changes affect Kp.

a. add PCl_3 $\rightarrow$ $[Cl_2]\downarrow$
b. add PCl_5 $\leftarrow$ $[Cl_2]\uparrow$
c. raise temp $\leftarrow$ $[Cl_2]\uparrow$
d. decrease vol $\rightarrow$ $[Cl_2]\downarrow$
e. Kp changes $\rightarrow$ only change in temp (c) affect Kp

4. (brown573) At 448°C the equilibrium constant, K_c, for the reaction $H_{2(g)} + I_{2(g)} \leftrightarrow 2HI_{(g)}$ is 51. Predict how the reaction will proceed to equilibrium at 448°C if we start with 2.0E-2mol of HI, 1.00E-2 mol H_2 and 3.0E-2mol of I_2 in a 2.0L container. In other words, which way will it shift, products or reactants.

$$Q = \frac{[2E^{-2}]^2}{[1E^{-2}][3E^{-2}]} = 1.3 \quad K = 51 \quad Q < K$$

shift toward products

$CO_{(g)} + H_2O_{(g)} \rightleftharpoons CO_{2(g)} + H_{2(g)} + heat$

Predict the direction of equilibrium shift (right, left, or no shift) if the following changes occur:

a.) The addition of more H_2O? $\rightarrow$

b.) The removal of some H_2? $\rightarrow$

*c.) Raising the temperature? $\leftarrow$

d.) Addition of a catalyst? no shift

e.) Increasing the volume? no shift (same amount of mol on each side)

9. What would be the effect of each of the following on the concentration of CO (increase, decrease, or no effect) when the following stresses are placed on the equilibrium involving the synthesis of methanol?

$$CO(g) + 2H_2(g) \leftrightarrow CH_3OH(g)$$

3 2

a.) The removal of CH_3OH? $\rightarrow$ $[CO]\downarrow$

b.) Lowering the concentration of H_2? $\leftarrow$ $[CO]\uparrow$

c.) The addition of a catalyst? no shift (no change [CO])

d.) Decreasing the volume? $\rightarrow$ $[CO]\downarrow$

10. A small percentage of nitrogen gas and oxygen gas in the air combine at the high temperatures found in automobile engines to produce NO(g), an air pollutant.

$$N_2(g) + O_2(g) + heat \leftrightarrow 2NO(g)$$

Higher engine temperatures are used to minimize carbon monoxide production. What effect do higher engine temperatures have on the production of NO? Why?

if increase temperature, the reaction will shift toward the products and create more $NO_{(g)}$

Name
Kc/Kp Multiple choice

d 1. (ebbing14.3)
The equilibrium expression for K_c for the system
$CO_{2(g)} + CaO_{(s)} \rightarrow CaCO_{3(s)}$ is

a. $[CaCO_3]/[CO_2][CaO]$
b. $[CaCO_3]/[CO_2]$
c. $[CO_2]$
[✓ d. $1/[CO_2]$]
e. $[CO_2][CaO]$

e 2. (Ebbing 14.4)
In which of the following does the reaction go the least to completion (see the following K values)
a. 10E5
b. 10E3
c. 10E0
d. 10E-3
[✓ e. 10E-5] smallest K is most reactants

a 3. (Ebbing 14.7)
Carbon disulfide and chlorine react according to the following equation:
$CS_{2(g)} + 3Cl_2 \leftrightarrow S_2Cl_{2(g)} + CCl_{4(g)}$
When 1.00 mol of CS_2 and 3.00 mol of Cl_2 are placed in a 2.00L container and allowed to come to equilibrium, the mixture is found to contain 0.250mol of CCl_4. What is the amount of Cl_2 at equilibrium?
[✓ a. 2.25 mol]
b. 2.75 mol
c. 0.75 mol
d. .25 mol
e. .50 mol

$CS_{2(g)} + 3Cl_{2(g)} \rightleftharpoons S_2Cl_{2(g)} + CCl_{4(g)}$
I 1 3
S -0.25 -0.75 +0.25 +0.25
E 0.75 $\boxed{2.25}$ 0.25 0.25

d 5. (Ebbing14.11)
Which expression correctly describes the equilibrium constant for the following reaction?
$4NH_{3(g)} + 5O_2 \leftrightarrow 4NO_{(g)} + 6H_2O(g)$
a. $Kc = 4[NH_3] + 5[O_2]/6[H_2O] + 4[NO]$
b. $Kc = 6[H_2O] + 4\,[NO]/4[NH_3]+5[O_2]$
c. $Kc = [H_2O][NO]/NH_3][O_2]$
[✓ d. $K_c = [H_2O]^6[NO]^4\,/[NH_3]^4[O_2]^5$]
e. $Kc = [NH_3]^4[O_2]^5/[H_2O]^6[NO]^4$

$$K_c = \frac{[H_2O]^6[NO]^4}{[NH_3]^4[O_2]^5}$$

b 7. (ebbing14.19)
Consider the reaction system
$Br_2(g) + Cl_2(g) \leftrightarrow 2BrCl(g)$

$$K_c = \frac{[BrCl]^2}{[Br_2][Cl_2]} = \frac{(0.015)^2}{(0.006)(0.0095)} = 3.95$$

At a given temperature. When the system is at equilibrium, the molar concentrations of Br_2, Cl_2 and BrCl are 0. 0060M, 0.0095M, and 0.015M, respectively. The value of Kc for this system is
a. .025
[✓ b. 3.9]
c. 27
d. 53
e. 260

d 8. (ebbing14.21)
For the reaction system
$2HI(g) \Leftrightarrow H_2(g) + I_2(g)$

$Q = \frac{[H_2][I_2]}{[HI]^2} = \frac{(1.5E-3)(1.5E-3)}{[1.5E-3]^2} = 1$

reactant favored

$K_c = 0.020$ at 720K. If the initial concentrations of HI, H_2, and I_2 are all 1.50E-3M at 720K, which one of the following statements ~~is~~ are correct?

a. The system is at equilibrium.

[✓ d.] The concentration of HI will increase as the system is approaching equilibrium. ← shift

b. The concentrations of HI and I_2 will increase as the system is approaching equilibrium.

c. The concentrations of H_2 and HI will decrease as the system approaches equilibrium.

e. The concentrations of H_2 and I_2 will increase as the system is approaching equilibrium.

c 9. (ebbing14.22) For the reaction $2H_2S(g) \Leftrightarrow 2H_2(g) + S_2(g)$

I 0.01 1.0 1.5

at a certain temperature Kc equals 4500. What will happen when 0.010 mol of $H_2S(g)$, 1.0 mol of H_2 and 1.5 mol of S_2 are added to a 2.0L container and the system is brought to the temperature at which Kc = 4500.

↗ $Q \neq K$

a. Nothing, the system is at equilibrium.

b. More H_2S will be formed.

[✓ c.] More H_2 will be formed than S_2

d. More S_2 will be formed than H_2.

e. The amount of H_2 formed will be half the amount of S_2.

$Q = \frac{(1)^2(1.5)}{(0.01)^2} = 15000 \quad K = 4500$

$Q < K$ shift →

a 10. (ebbing14.23) A 1.00 mol sample of HI is placed in a 1-L vessel at 460C, and the reaction system is allowed to come to equilibrium. The HI partially decomposes, forming 0.11 mol H_2 and 0.11 mol I_2. What is the equilibrium constant for the reaction?

$H_2(g) + I_2(g) \Leftrightarrow 2HI(g)$

[✓ a. 0.020]

b. 7.1

c. 8.1

d. 50

e. 65

$2HI \rightleftharpoons H_2 + I_2$

I 1.0

S 0.22 +0.11 +0.11

E 0.78 0.11 0.11

$K_c = \frac{(0.11)(0.11)}{(0.78)^2} = 0.0199$

11. (ebbing14.25)
Consider the equilibrium $N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$ at a certain temperature. An equilibrium mixture in an 8.00L vessel contains .800 mol N_2, and 1.20 mol H_2. What is the value of K_c?

* ?

a. 1.85

b. 29.6

c. 37.4

d. 75.8

e. 119

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

I

S

E 0.8 1.2

need more info

$K_c = \frac{(NH_3)^2}{(N_2)(H_2)^3} = \frac{(NH_3)^2}{(0.8)(1.2)^3}$

c 12.

$$4H_{2(g)} + CS_{2(g)} \Leftrightarrow CH_{4(g)} + 2H_2S_{(g)}$$

The system reaches equilibrium according to the equation above. A mixture of 2.5 mol H_2, 1.50 mol CS_2, 1.50 mol CH_4 and 2.00mol H_2S is placed in a 5L reaction vessel. When equilibrium is achieved, the concentration of CH_4 has become 0.25M.

Changes in concentration occur as this system approaches equilibrium. Which expression gives the best comparison of the changes in those concentration shown in the ratio below?

$\Delta[H_2S]/\Delta[CS_2]$

$\frac{\Delta[H_2S]}{\Delta[CS_2]} = \frac{-2}{+1}$

$4H_2 + CS_2 \rightleftharpoons CH_4 + 2H_2S$

I 2.5 1.5 1.5 2

S +4x +x -x -2x

E 3.5 1.75 1.25 1.5

mols

X shift = 0.25

$0.25M = \frac{X\ mol}{5L}$

1.25 mol CH_4

a. +2 /+1
b. +2 /-1
[✓ c.] -2 /+1
d. -1 /+1
e. -1 /+2

b 13.

What is the change in the number of moles of $H_2S(g)$ present as the system moves from its original state to the equilibrium described?

a. -1.25
[✓ b.] -.50
c. -.25
d. +.25
e. +.50

$-2x = -0.50$

(0.25)

d 14.

When equilibrium is achieved, the concentration of CH_4 has become 0.25M.

What is the number of moles of $CS_2(g)$ at equilibrium

a. .25
b. .35
c. .75
d. 1.25
[✓ e.] 1.75

[illegible]

1.5+0.25 = 1.75

b 15.

When equilibrium is achieved, the concentration of CH_4 has become 0.25M.

What is the concentration in moles per liter of H_2 at equilibrium?

a. .5
[✓ b.] .70
c. 1.00
d. 2.5
e. 3.5

$\frac{3.5\ mol\ H_2}{5L} = 0.3$