

## 1 Reaction Kinetics & Rate Laws

1. According to collision theory, in order for a chemical reaction to proceed, the reactants must collide with proper \_\_\_\_\_ and with sufficient \_\_\_\_\_.
2. Draw a graph of energy vs reaction progress as was shown in class.

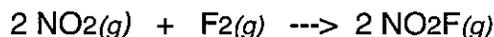
3. Factors which can change the speed of a reaction are:

- 1) \_\_\_\_\_
- 2) \_\_\_\_\_
- 3) \_\_\_\_\_
- 4) \_\_\_\_\_

### Rate Laws

After timing the clock reaction at various concentrations, you have probably come to appreciate that the rate of a reaction depends (in part) on the \_\_\_\_\_ of the \_\_\_\_\_.

A **rate law** is an equation that relates the rate of a reaction to the concentration of reactants.

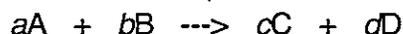


The rate of the above reaction is observed to be proportional to the concentration of nitrogen dioxide. When  $[\text{NO}_2]$  is doubled, the rate doubles. The rate is also proportional to  $[\text{F}_2]$ ; doubling the fluorine doubles the rate. The following equation is the rate law:

$$\text{rate} = k[\text{NO}_2][\text{F}_2]$$

$k$  is called the **rate constant**. It is a *proportionality constant in the relationship between rate and concentrations*. In general, when  $k$  is small, the reaction is \_\_\_\_\_. When  $k$  is large, the reaction is \_\_\_\_\_.

For a general reaction, consider A and B are reactants to produce C and D, according to this reaction:



The rate law for this reaction would be as follows:

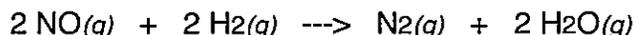
$$\text{Rate} = k[\text{A}]^m[\text{B}]^n$$

The exponents  $m$  and  $n$  must be determined experimentally, and cannot be obtained simply by looking at the balanced equation. What are the values of the exponents for  $[\text{NO}_2]$  and  $[\text{F}_2]$  in the example above? \_\_\_\_\_

**Reaction Order** is the power to which the concentration of a reactant must be raised to give the observed relationship between concentration and rate. For a given reactant, the reaction order equals the exponent of its concentration.

For the reactions of  $\text{NO}_2$  and  $\text{F}_2$ , the reaction is \_\_\_\_\_ with respect to  $\text{NO}_2$  and is \_\_\_\_\_ with respect to  $\text{F}_2$ . The **overall order** of a reaction equal the sum of the orders of the reactants. In this example, the reaction is \_\_\_\_\_ overall.

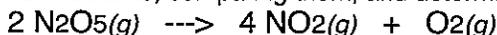
problem:



The experimentally determined rate law is:  $\text{Rate} = k[\text{NO}]^2[\text{H}_2]$

This reaction is \_\_\_\_\_ order in  $\text{NO}$ , \_\_\_\_\_ order in  $\text{H}_2$ , and \_\_\_\_\_ order overall.  
(note: it is possible for reaction orders to be fractions or negative #'s. For more info, ask someone else...)

**Determining the Rate Law:** a simple way to determine the rate law for a reaction is the *initial-rate* method. It involves measuring only the initial concentrations of reactants, comparing them, and determining the reaction orders. Example:



	initial $[\text{N}_2\text{O}_5]$ (mol/L)	initial rate (mol/L · s)
experiment 1	0.010	$4.8 \times 10^{-6}$
experiment 2	0.020	$9.6 \times 10^{-6}$

Notice a pattern? If  $[\text{N}_2\text{O}_5]$  doubled, then the rate is \_\_\_\_\_. Remember:  $\text{Rate} = k[\text{N}_2\text{O}_5]^m$ . If  $m=1$ , then the rate would be directly proportional to  $[\text{N}_2\text{O}_5]$ . Does this match the experimental data? \_\_\_\_\_. Since we doubled the concentration, the new rate is  $2^m$  times the original. If  $m=2$ , then the rate would be  $2^2$ , or 4 times as large. Does this match the data? \_\_\_\_\_. Ok then, the reaction is \_\_\_\_\_ order with respect to  $[\text{N}_2\text{O}_5]$ , and \_\_\_\_\_ order overall.

To determine the value of the rate constant  $k$ , simply substitute values of the **rate** and  $[\text{N}_2\text{O}_5]$  from any one of the experiments on the data table into the rate law.

$$\text{Rate} = k[\text{N}_2\text{O}_5]^m$$

$$9.6 \times 10^{-6} = k (0.020)^1$$

$$k = 4.8 \times 10^{-4}$$

example:



$$\text{rate} = k[\text{A}]^m$$

	initial $[\text{A}]$ (mol/L)	initial rate (mol/L · s)
experiment 1	0.05	$3 \times 10^{-4}$
experiment 2	0.10	$12 \times 10^{-4}$
experiment 3	0.20	$48 \times 10^{-4}$

Find the reaction order of the reaction with respect to A, the overall order of the reaction, and  $k$ .

**solution:**

doubling the concentration causes the rate to increase by \_\_\_\_\_. 2 raised to what power equals this increase? \_\_\_\_\_.

The reaction is \_\_\_\_\_ order with respect to A. The reaction is \_\_\_\_\_ order overall.  $k =$  \_\_\_\_\_.

Problem 1:



	initial $[\text{NO}_2^-]$ (mol/L)	initial $[\text{NH}_4^+]$ (mol/L)	initial rate (mol/L · s)
experiment 1	0.010	0.200	$5.4 \times 10^{-7}$
experiment 2	0.020	0.200	$10.8 \times 10^{-7}$
experiment 3	0.040	0.200	$21.5 \times 10^{-7}$
experiment 4	0.200	0.0202	$10.8 \times 10^{-7}$
experiment 5	0.200	0.0404	$21.6 \times 10^{-7}$
experiment 6	0.200	0.0808	$43.3 \times 10^{-7}$

Find the reaction order with respect to  $\text{NO}_2^-$  &  $\text{NH}_4^+$ , the overall order, and  $k$ .

# REACTION RATES – DETERMINING THE ORDER OF A REACTION

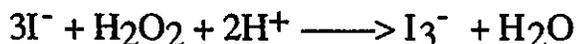
The rate of a reaction is governed by the following relationship:

$$\text{RATE} = k [A]^{N_a}[B]^{N_b}[C]^{N_c}\dots$$

The quantities in brackets are concentrations and are raised to an appropriate power. Together with the constant (k), they give the overall rate of the reaction.

The numerical values of  $N_a$ ,  $N_b$ , and  $N_c$  must be determined by experimentation. These numbers determine the order of the reaction. Added together, they give the overall order of the reaction. The purpose of this experiment is to determine the order of an iodine clock reaction with respect to  $H_2O_2$ .

The reaction to be studied in this experiment is the acid buffered oxidation of iodide to triiodide by hydrogen peroxide:



You will study the rate vs concentration of  $H_2O_2$ .

## MATERIALS

distilled water  
1 x 12-well plates (2)  
1 mL microtip pipet (for water)  
1 mL microtip pipets filled with the following reagents (2 each):  
    solution A                      solution B                      solution C  
cassette box (1)  
clock with sweep second hand or stopwatch

## PROCEDURE

**Caution: Put on your goggles and apron now!!**

### PART A: PREPARING THE STANDARD DILUTIONS

1. Arrange two 1 x 12-well plates such that the numbers are read from left to right.
2. Place drops of solutions into the two plates according to the tables on the next page.

### PLATE 1: SOLUTION A

Cell#	1	2	3	4	5	6	7	8	9	10	11	12
Drops Sol'n A	4	4	4	3	3	3	2	2	2	1	1	1
Drops Water	0	0	0	1	1	1	2	2	2	3	3	3

### PLATE 2: SOLUTIONS B AND C

Cell#	1	2	3	4	5	6	7	8	9	10	11	12
Drops Sol'n B	1	1	1	1	1	1	1	1	1	1	1	1
Drops Sol'n C	4	4	4	4	4	4	4	4	4	4	4	4

#### PART B: THE SHAKEDOWN TECHNIQUE

1. Gently invert Plate 2 onto Plate 1 so that the cells are lined up.
2. Holding the plates firmly together, shake them once vigorously in a downward motion. This is done by dropping your hands as fast as you can and stopping abruptly. There is *no upward motion* so this is a *shakedown* method. **Your partner should start the timing immediately.**
3. Record the time that each cell takes to turn blue. Repeat steps #1 through #3 twice.
4. Dump the contents from the plates into the sink and rinse with distilled water.

#### ANALYSIS

The rate of reaction can be represented by the following equation:  
**Rate =  $k[\text{H}_2\text{O}_2]^{\text{N}1}[\text{I}^-]^{\text{N}2}[\text{H}^+]^{\text{N}3}$ .** The concentrations of  $\text{I}^-$  and  $\text{H}^+$  are constant because the amounts used in plate 2 are constant. Thus, we may write the rate: **Rate =  $k'[\text{H}_2\text{O}_2]^{\text{N}1}$**  where  $(\text{I}^-)^{\text{N}2}(\text{H}^+)^{\text{N}3}$  is absorbed into the pseudo-rate constant,  $k'$ .

1. Plot a graph of your results. Use the x-axis for time in seconds and the y-axis for drops of  $\text{H}_2\text{O}_2$  solution. Draw the best fitting curve to this plot.
2. Plot the inverse of the drop count ( $\frac{1}{\text{drops}}$ ) vs time and fit the best curve to these data.

3. Plot the inverse of the drop count squared ( $\frac{1}{\text{drops}^2}$ ) vs time and fit the best curve to these data.

The order of the reaction with respect to  $\text{H}_2\text{O}_2$  can be determined from the graph that best approximates a **straight** line.

GRAPH

Drops vs Time [Zero Slope]

Inverse drop count vs Time

Inverse drop count squared vs Time

ORDER

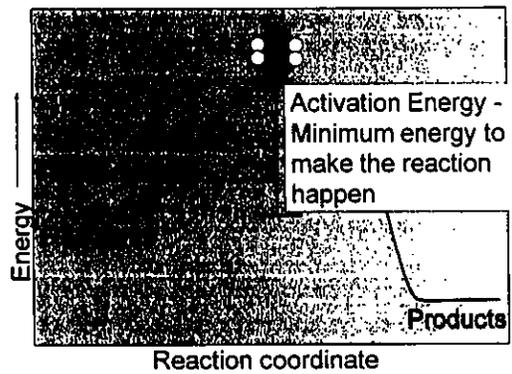
Zero Order

First Order

Second Order

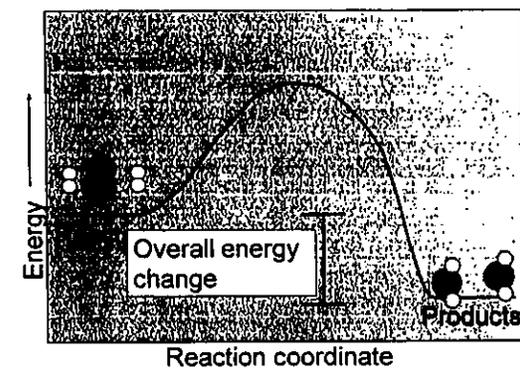
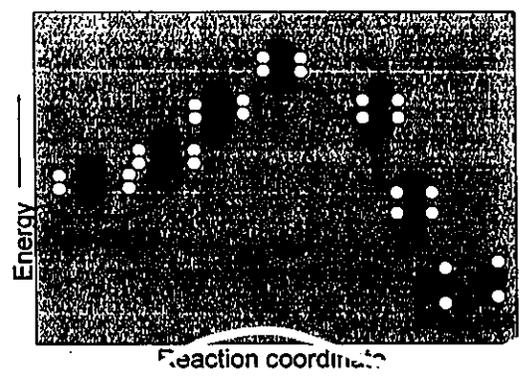
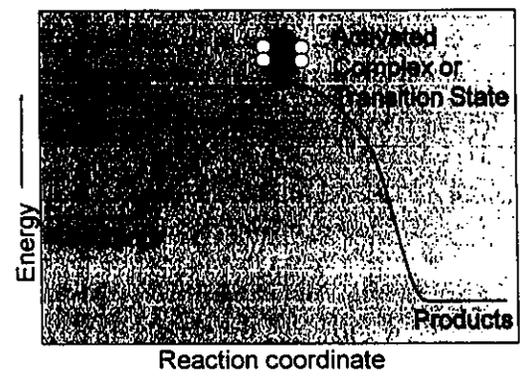
4. According to your graphs, what is the order of the reaction with respect to  $\text{H}_2\text{O}_2$ ?
5. Briefly describe how you would use variations of this lab to determine the order for the  $\text{I}^-$  and  $\text{H}^+$ .
6. Having determined the order for the three reactants, how would you determine the value of  $k$ ?

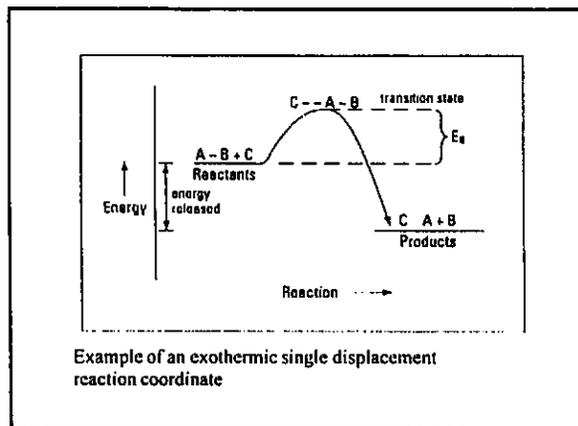
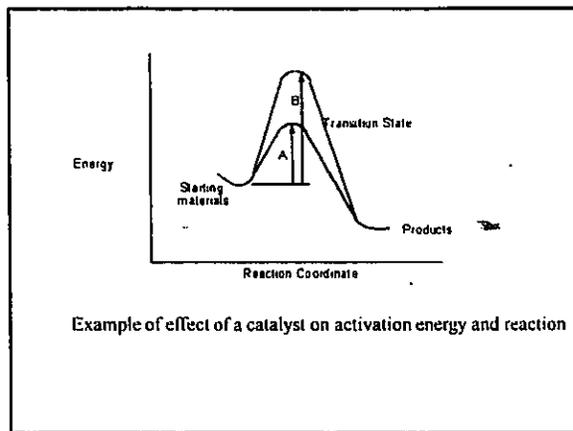
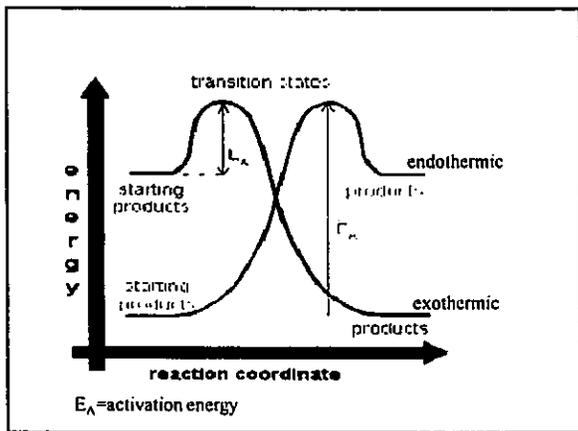
## Reaction Rate Introduction to Kinetics



## Review of Collision Theory

- In order to react molecules and atoms must collide with each other.
- They must also collide with each other hard enough to react.
- Anything that increase collisions and strength of collisions will make the reaction faster.
- Collisions that result in product formation are known as effective collisions



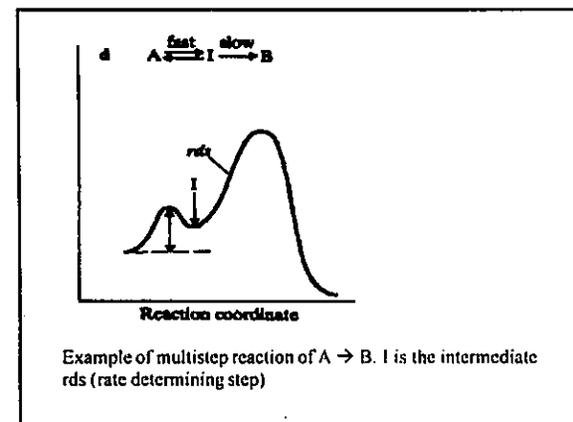


### Multistep Reactions

- Elementary reaction- a reaction that happens in a single step.
- Reaction mechanism is a description of how the reaction really happens.
- It is a series of elementary reactions.
- The product of an elementary reaction is an intermediate.
- An intermediate is a product that immediately gets used in the next reaction.

### Factors that affect Reaction Rate

- **Temperature:** Higher temperatures means faster particles giving more collisions therefore faster Reactions.
- **Concentration:** The more concentrated the closer together the reactants are therefore they collide more often giving faster reactions.
- **Particle size:** Molecules can only collide at the surface. Smaller particles have larger surface area. Therefore smaller particles give faster reactions.
- **Pressure:** increased pressure generally increases collision rate increasing reaction rate.
- **Nature of Reactants:** some reactants are more or less reactive than others. recall the reactivity of elements (metals and nonmetals) from the periodic trends.
- **Catalysts:** substances that speed up a reaction without being consumed by reaction (enzyme). Speeds up reaction by lowering activation energy and giving the reaction a new path along its reaction coordinate. The new path has a lower activation energy. Therefore, the reaction is faster. This factor influences reaction rate more than any other factor.
- **Inhibitor:** a substance that blocks a catalyst.



## Mechanisms and rates

- There is an activation energy for each elementary step.
- Slowest step (rate determining) must have the highest activation energy.
- This rate determining step determines how fast the reaction will occur.

## Formation of Solution Equilibrium

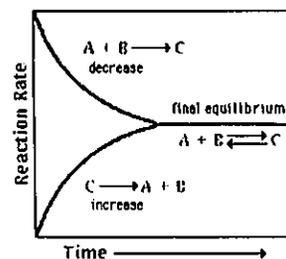
- Solids in liquids exist in equilibrium in a saturated solution
- When the rate of dissolving and recrystallization are equal, equilibrium exists and the solution is saturated.
- Temperature increasing or decreasing will change the rate of the forward or reverse reaction until a new equilibrium is reached.

## Equilibrium

- Recall that some reactions are reversible
- Equilibrium – the forward and reverse reactions occur at the same rate
- State of balance between two opposite processes taking place at the same rate.
- Can only occur in a system in which neither reactants nor products can leave the system.
- The quantities of the reactants and products may not be always be equal, usually not. Just the rates of the forward and reverse reactions
- Equilibrium is involved chemical reactions and also in physical processes such as phase changes and dissolving.

## Chemical Equilibrium

- When you mix reactants at first and there are no products present, only the forward reaction can occur
- As time progresses, the reactants concentrations decrease and the forward reaction slows, while the concentration of the products increase
- This causes the rate of the reverse reaction to increase
- The forward reaction rate will continue to slow as the reverse reaction rate increases proceeds until equilibrium is met. Reaction rates are equal
- That is chemical equilibrium



## Phase Change Equilibrium

- Can exist between the solid and liquid phases of a substance.
- At melting point of solid or freezing point of liquid phase
- Melting can occur at same time as freezing
- Also, in a closed system, rate of evaporation can equal rate of condensation.

## $K_{eq}$ or $K_c$ The Equilibrium Expression

- $K_{eq}$  (also called  $K_c$ ) Equilibrium Expression is a mathematical expression that shows the relationship of reactants and products in a system at equilibrium
- The concentration of products over concentration of reactants expressed in moles per liter
- Each concentration is raised to the power of its coefficient in the balanced equation.
- ( $K_{eq}$ ) remains the same for a particular reaction at a specified temperature
- When calculating  $K_{eq}$  only consider (aq) and (g) phase chemical species as they are the only chemical species whose concentrations will fluctuate
- $K_{eq}$  is known as  $K_c$  when concentrations are expressed in moles / liter.  $K_{eq}$  is known as  $K_p$  when expressed in terms of pressure (used for gases we will not calculate  $K_p$  only  $K_c$  and  $K_{eq}$   $K_c$  will be used even for gasses)
- If the value of  $K_{eq}$  is large, the products are present in a larger concentration than reactants. The products are favored
- If the value is small, the opposite is true, reactants are favored

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad aA + bB \rightleftharpoons cC + dD$$

### Four steps to writing equilibrium expressions

1. Write the balanced equation for the system.
2. Place the products as factors in the numerator of a fraction and the reactants as factors in the denominator. (Only necessary for (aq) and (g) phase chemical species)
3. Place a square bracket around each formula. The square bracket means molar concentration (moles / liter) (Note mol/L used for gasses also).
4. Write the coefficient of each substance as the power of its concentration. The resulting expression is the equilibrium expression, which should be set to equal to the  $K_{eq}$  for that reaction.

### Sample $K_{eq}/K_c$ Expression Writing

- Write the equilibrium expression for the equilibrium system of nitrogen, hydrogen, and ammonia. All these are gasses still can use moles per liters
- Reactants:  $N_2$ ,  $H_2$   
Products:  $NH_3$
- $N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g) + \text{heat}$
- $K_{eq} = \frac{[NH_3]^2}{[H_2]^3[N_2]}$
- Practice Writing the equilibrium expression for:  
 $2NaCl(aq) + CaCO_3(s) \leftrightarrow Na_2CO_3(aq) + CaCl_2(aq)$
- remember don't use (s) or (l) in  $K_{eq}$  expressions

**Class work. Write out the equilibrium expression for the following reactions:**

1.  $PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$
2.  $NH_4HS(s) \leftrightarrow H_2S(g) + NH_3(g)$
3.  $NH_4NO_3(s) \leftrightarrow NH_4^+(aq) + NO_3^-(aq)$
4.  $HCOOH(aq) + H_2O(l) \leftrightarrow H_3O^+ + HCOO^-(aq)$
5.  $Na_2S(s) \leftrightarrow 2Na^+(aq) + S^{2-}(aq)$

Calculations using The Equilibrium Constant

For the general homogeneous, reversible chemical reaction:  $aA + bB \rightleftharpoons cC + dD$

At equilibrium, when the rate of the forward and reverse reactions are equal, we can write the general equilibrium expression:

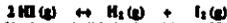


$$K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

If we know the amounts or concentrations of the reactants and products of a chemical system at equilibrium, we can calculate the value of the equilibrium constant  $K_{eq}$ .

Similarly, if we know the value of  $K_{eq}$  at a specified temperature, and the initial concentrations of the reactants, we can calculate the equilibrium concentrations of each species. The stoichiometry (molar coefficients) of the reaction is very important in these calculations.

eg. At 430 °C, the equilibrium constant,  $K_c$ , for the following reaction is  $1.84 \times 10^{-2}$ .



If 0.100 mol of hydrogen iodide is placed in a 1.00 L container and allowed to reach equilibrium at this temperature, find the concentrations of all the species at equilibrium.

Solution:

1. Set up a chart called an "ICE" table. "I" stands for "initial", "C" stands for "change" and "E" stands for equilibrium. Write the equation across the top of the table as your heading. Include the  $K_{eq}$  value.

	2 HI (g)	↔	H <sub>2</sub> (g)	+	I <sub>2</sub> (g)	$K_{eq} = 1.84 \times 10^{-2}$
I						
C						
E						

2. Fill in any known initial concentration values. If you are given amounts that are not expressed as concentrations, convert to mol/L before you fill in the ICE table. For this example, we are told that there are 0.100 mol of HI in a 1.00 L flask. This is an initial concentration of 0.100 mol/L. Initially, there is no H<sub>2</sub> or I<sub>2</sub>, because the reaction has not started yet, so fill in 0.0 mol/L for these species.

	2 HI (g)	↔	H <sub>2</sub> (g)	+	I <sub>2</sub> (g)	$K_{eq} = 1.84 \times 10^{-2}$
I	0.100 mol/L		0.0 mol/L		0.0 mol/L	
C						
E						

3. As the reaction occurs and the system reaches equilibrium, the amount of H<sub>2</sub> and I<sub>2</sub> will increase while the amount of HI will decrease. If H<sub>2</sub> increases by an amount "x" then, using the molar coefficients, I<sub>2</sub> will also increase by "x". HI will decrease by "2x" because 2 moles of HI are used up for every one mole of H<sub>2</sub> and I<sub>2</sub> produced. Fill in these changes in molar concentration in the change row.

	2 HI (g)	↔	H <sub>2</sub> (g)	+	I <sub>2</sub> (g)	$K_{eq} = 1.84 \times 10^{-2}$
I	0.100 mol/L		0.0 mol/L		0.0 mol/L	
C	-2x		+x		+x	
E						

6. To find the equilibrium concentrations, substitute the value for "x" back into the equilibrium concentrations in the ICE table. Therefore:

$$[HI] = 0.100 - 2x$$

$$= 0.100 - (2 \times 0.0107)$$

$$= 0.0786 \text{ mol/L}$$

$$[H_2] = [I_2] = x$$

$$= 0.0107 \text{ mol/L}$$

\* the units for concentration are always the same as the initial units for concentration used in the ICE table.

Example #2: Solid ammonium chloride is heated to 400K in a closed one litre flask. The ammonium chloride decomposes to form ammonia and hydrogen chloride gas. The  $K_{eq}$  for this reaction at 400 K is  $6.0 \times 10^{-4}$ . What are the equilibrium concentrations of the gases?

Solution: The reaction is:  $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$   $K_{eq}$  at 400K is  $6.0 \times 10^{-4}$

1. This is a heterogeneous system. The species in the highest entropy phase are  $NH_3(g)$  and  $HCl(g)$ , so we will not be including the concentration of  $NH_4Cl(s)$  in the  $K_{eq}$  equation. Set up a chart called an "ICE" table. Write the equation across the top of the table as your heading. Include the  $K_{eq}$  value. We do not need a value for the  $NH_4Cl$  initial concentration. The concentration of the products will initially be zero.

	$NH_4Cl(s)$	↔	$NH_3(g)$	+	$HCl(g)$	$K_{eq}$ at 400K is $6.0 \times 10^{-4}$
Initial	omit		0.0 mol/L		0.0 mol/L	
Change	omit					
Equilibrium	omit					

2. As the  $NH_4Cl$  decomposes, the concentration of  $NH_3(g)$  and  $HCl(g)$  will increase. The molar coefficient for each gas is one, so if one mole of  $NH_3$  forms, one mole of  $HCl$  will also form. Fill this into the change row of the ICE table. Let "x" be the increase in both gases.

	$NH_4Cl(s)$	↔	$NH_3(g)$	+	$HCl(g)$	$K_{eq}$ at 400K is $6.0 \times 10^{-4}$
Initial	omit		0.0 mol/L		0.0 mol/L	
Change	omit		+x		+x	
Equilibrium	omit					

3. To find the equilibrium concentrations of both gases, add the initial and change values together. Enter this in the equilibrium row.

	$NH_4Cl(s)$	↔	$NH_3(g)$	+	$HCl(g)$	$K_{eq}$ at 400K is $6.0 \times 10^{-4}$
Initial	omit		0.0 mol/L		0.0 mol/L	
Change	omit		+x		+x	
Equilibrium	omit		x		x	

4. Now we have an expression with a single variable (x) that we can solve using the  $K_c$  equilibrium expression, because we know the value of  $K_c$  for the system at equilibrium.

a) Write the  $K_c$  equation including only the species in the highest entropy phase:

$$K_c = [NH_3][HCl]$$

4. At equilibrium, the concentrations of each species will be equal to the initial concentration plus the change in concentration. Add the top two columns to get the equilibrium concentration for each species, and fill this in the equilibrium row.

	2 HI (g)	↔	H <sub>2</sub> (g)	+	I <sub>2</sub> (g)	$K_{eq} = 1.84 \times 10^{-2}$
I	0.100 mol/L		0.0 mol/L		0.0 mol/L	
C	-2x		+x		+x	
E	0.100 - 2x		x		x	

5. Now we have an expression with a single variable (x) that we can solve using the  $K_c$  equilibrium expression, because we know the value of  $K_c$  for the system at equilibrium.

a) Write the  $K_c$  equation including only the species in the highest entropy phase:

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

b) Substitute the concentration terms with the equilibrium values from the ICE table, and the value of  $K_c$ :

$$1.84 \times 10^{-2} = \frac{x \cdot x}{(0.100 - 2x)^2}$$

c) Now, to solve for x. There are three possibilities:

- If the equation does not include any squared terms, then rearrange and solve for x
- If the equation includes squared terms that are perfect squares, take the square root of both sides of the equation, then rearrange and solve for x
- If the equation includes squared terms that are not perfect squares, you will need to expand out all terms and rearrange the equation into the form  $ax^2 + bx + c = 0$ . Solve for x using the quadratic formula.

For our example, it is a perfect square ( $x \cdot x = x^2$ ), so we will take the square root of both sides (the square root is the same as the number raised to the power 1/2):

$$(1.84 \times 10^{-2})^{1/2} = \frac{x}{(0.100 - 2x)}$$

Now rearrange and solve for x:

$$0.1356 = \frac{x}{(0.100 - 2x)}$$

$$0.1356(0.100 - 2x) = x$$

$$0.01356 - 0.2712x = x$$

$$0.01356 = 1.2712x$$

$$x = 0.0107$$

2

b) Substitute the concentration terms with the equilibrium values from the ICE table, and the value of  $K_c$ .

$$6.0 \times 10^{-4} = x \cdot x$$

c) Solve for x by taking the square root of both sides.

$$6.0 \times 10^{-4} = x^2 \text{ and } x = 7.7 \times 10^{-3} \text{ mol/L}$$

d) Therefore, the concentration of both  $NH_3$  and  $HCl$  gases at equilibrium will be  $7.7 \times 10^{-3} \text{ mol/L}$ .

Does this make sense?  $K_{eq}$  is very small, so the reaction will not proceed very far to the right, so the concentrations of the products will be small. Yes, this answer does make sense.

Example #3: How many moles of HI are present at equilibrium when 2.0 moles of H<sub>2</sub> is mixed with 1.0 mole of I<sub>2</sub> in a 0.50 L container and allowed to react at 448°C. At this temperature  $K_{eq} = 50$

Solution



Then, we have been given the number of moles and volume of the container, so calculate the initial concentrations of each species:

$$[H_2] = n/V = 2.0 \text{ moles} / 0.50 \text{ L} = 4.0 \text{ moles/L}$$

$$[I_2] = n/V = 1.0 \text{ moles} / 0.50 \text{ L} = 2.0 \text{ moles/L}$$

$[HI] = 0 \text{ moles}$  therefore 0 moles/L. (the reaction has not started yet, so no product has formed.)

Fill in the initial concentrations of each species in the ICE table:

	H <sub>2</sub> (g)	+	I <sub>2</sub> (g)	↔	2 HI (g)	$K_{eq} = 50$
I	4.0 mol/L		2.0 mol/L		0.0 mol/L	
C						
E						

As the reaction occurs, for every mole of H<sub>2</sub> that reacts, one mole of I<sub>2</sub> will also react, and two moles of HI will form. Let "x" represent the changes in each species, and fill in the change row of the ICE table:

	H <sub>2</sub> (g)	+	I <sub>2</sub> (g)	↔	2 HI (g)	$K_{eq} = 50$
I	4.0 mol/L		2.0 mol/L		0.0 mol/L	
C	-x		-x		+2x	
E						

Add the first two rows together to find the equilibrium values of each species and complete the "E" row of the ICE table:

	H <sub>2</sub> (g)	+	I <sub>2</sub> (g)	↔	2 HI (g)	$K_{eq} = 50$
I	4.0 mol/L		2.0 mol/L		0.0 mol/L	
C	-x		-x		+2x	
E	4.0 - x		2.0 - x		+2x	

Substitute the equilibrium concentrations of each species and  $K_{eq}$  into the  $K_{eq}$  equation:

$$K_{eq} = \frac{(2x)^2}{[4-x][2-x]}$$

4

**PROBLEMS**

$$50 = (2x)^2$$
$$(4.0 - x)(2.0 - x)$$

We're not so lucky this time. There are no perfect squares here. We will have to expand the equation by multiplying out and get the equation into standard form.

$$50 = 4x^2$$
$$8.0 - 6.0x + x^2 \quad (\text{Careful here: } (2x)^2 \text{ is equal to } 4x^2)$$

$$400 - 300x + 50x^2 = 4x^2$$

$$46x^2 - 300x + 400 = 0 \quad (\text{the standard form of } ax^2 + bx + c = 0)$$

Use the quadratic equation to solve for 'x'.

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

In this case  $a = +46$ ,  $b = -300$  and  $c = +400$ .

The quadratic equation is a mathematical relationship for solving the equation for a parabola to find its roots. In chemistry the 'roots' represent real life values. One will be realistic and the other will not be.

$$x = \frac{+300 \pm \sqrt{(-300)^2 - 4(46)(400)}}{2(46)}$$

$$x = \frac{+300 \pm \sqrt{90000 - 73600}}{92}$$

$$x = \frac{+300 \pm 127}{92}$$

$$x = 1.7 \text{ and } 1.9$$

Only one of these values will be possible. Linking back at our data, we started with only 4.0 mol/L of  $H_2$ . If we subtract 4.7 mol/L, we will get a negative value for concentration, which is impossible. Therefore,  $x=4.7$  is not a realistic answer, so  $x=1.9$  mol/L.

Substitute 'x = 1.9 mol/L' back up into the equilibrium concentration calculations of the ICE table and calculate the equilibrium concentrations of each species:

$[HI]$  at equilibrium is  $2x = 2(1.9 \text{ mol/L}) = 3.8 \text{ mol/L}$ .

But, the question asked us for the number of moles of HI, so to convert the concentration back to moles, use the equation  $n = CV = 3.8 \text{ mol/L} \times 0.5 \text{ L} = 1.9 \text{ mol of HI}$ .

And  $[HI]$  equals  $2x = 2(1.9 \text{ mol/L}) = 3.8 \text{ mol/L}$ .  
The moles of HI present at equilibrium would be:  
 $n = CV = 3.8 \text{ moles/L} \times 0.5 \text{ L}$   
 $= 1.9 \text{ moles}$

Does this make sense? Yes, the  $K_{eq}$  for this reaction is large ( $K_{eq} = 50.0$ ), so the reaction will go almost to completion, and it did.

5

### K<sub>eq</sub> Problems

Write out your answers to the following on a separate sheet of paper.

- 1) At certain temperature the reaction below has a  $K_{eq}$  value of 25.1
- $$N_2(g) + 3 H_2(g) \leftrightarrow 2 NH_3(g)$$

If 5.50 moles of  $NH_3(g)$  is placed in a 20.0 L flask at a constant temperature and allowed to come to equilibrium. What are the concentrations (in mol / liter) of all the chemical species in this reaction?

- 2) At specific temperature the reaction below has a  $K_{eq}$  value of  $6.89 \times 10^{-4}$
- $$N_2(g) + 2 O_2(g) \leftrightarrow 2 NO_2(g)$$

Suppose 2.56 moles of  $O_2(g)$  is combined with 5.67 moles of  $N_2(g)$  in a 55.0 L flask at a constant temperature. What are the concentrations (in mol / liter) of all the chemical species when the system reaches equilibrium?

- 3) At a temperature of 275K the reaction below has a  $K_{eq}$  value of 5.12
- $$CaCl_2(s) \leftrightarrow Ca^{2+}(aq) + 2 Cl^{-}(aq)$$

Suppose a piece of  $CaCl_2(s)$  is dissolved in sufficient water at 275K. What will the concentrations (in mol / liter) of  $Ca^{2+}(aq)$  and  $Cl^{-}(aq)$  when the system reaches equilibrium?

- 4) At specific temperature the reaction below has a  $K_{eq}$  value of  $4.99 \times 10^{-4}$
- $$PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$$

Suppose 760. g of  $PCl_5(g)$  is placed in a 30.5 L flask at a constant temperature. What are the concentrations (in mol / liter) of  $PCl_3(g)$  +  $Cl_2(g)$  when the system comes to equilibrium?

- 5) At a temperature of 425K the reaction below has a  $K_{eq}$  value of  $2.68 \times 10^{-4}$
- $$NH_4HS(s) \leftrightarrow H_2S(g) + NH_3(g)$$

Suppose a piece of  $NH_4HS(s)$  is placed in a 1.00 L flask at 425K. What will the concentrations (in mol / liter) of  $H_2S(g)$  and  $NH_3(g)$  when the system reaches equilibrium?

- 6) At specific temperature the reaction below has a  $K_{eq}$  value of  $3.76 \times 10^{-3}$
- $$I_2(g) \leftrightarrow 2I(g)$$

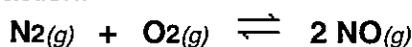
Suppose 1.50 moles of  $I_2$  is placed in a 15.5 L flask at a constant temperature. What are the concentrations (in mol / liter) of  $I_2$  and I when the system comes to equilibrium?

6

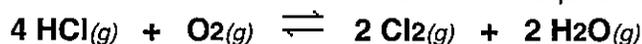
## Equilibrium

1. When a reaction has reached equilibrium, the \_\_\_\_\_ and \_\_\_\_\_ rates of reaction are \_\_\_\_\_. During dynamic equilibrium, molecules of \_\_\_\_\_ and \_\_\_\_\_ are still being formed, however their overall \_\_\_\_\_ are the same. At \_\_\_\_\_, it is not necessary to have \_\_\_\_\_ concentrations of \_\_\_\_\_ and \_\_\_\_\_. The equilibrium constant, known as \_\_\_\_\_, is specific for a given \_\_\_\_\_. It is calculated as the ratio of the concentrations of \_\_\_\_\_ to reactants, with each concentration raised to a \_\_\_\_\_ equal to the \_\_\_\_\_ of that substance. If  $K_{eq}$  is greater than 1, then \_\_\_\_\_ are favored. If  $K_{eq}$  is less than one, then reactants are \_\_\_\_\_. When calculating  $K_{eq}$ , only substances which are \_\_\_\_\_ or \_\_\_\_\_ are considered. Solids and liquids are \_\_\_\_\_ since their concentrations generally won't \_\_\_\_\_.

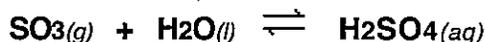
2. In a 4.0 L vessel, 0.015 mol  $N_2$ , 0.022 mol  $O_2$ , and 0.0065 mol  $NO$  have reached equilibrium at 890 K. Calculate  $k_{eq}$  for this reaction:



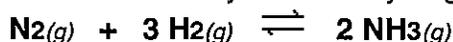
3. For the following reaction at equilibrium,  $[HCl] = 0.010 M$ ,  $[O_2] = 0.014 M$ ,  $[Cl_2] = 0.016 M$ , and  $K_{eq} = 214$  at 560 K. Calculate the concentration of the water at equilibrium.



4. In a 1.0 L tank, 0.95 mol  $SO_3$ , 8.95 mol  $H_2O$ , and 1.35 mol  $H_2SO_4$  are at equilibrium at 25°C. Calculate  $k_{eq}$ .



5. A 15.0 L tank at equilibrium is found to contain 4.50 mol of  $N_2$  and 2.11 mol ammonia ( $NH_3$ ). Given that  $K_{eq} = 3.2 \times 10^{-4}$ , calculate how many *moles* of hydrogen are present.

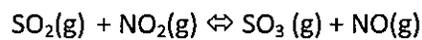


**Ans (IAO+1):** aqueous change concentrations equal equal equilibrium forward favored gases ignored  $K_{eq}$   
marshmallows moles power products products products products reactants reactants reverse temperature

**Ans (IRO+3):** 0.011 0.13 1.42 1.98 2.5 88.7 110.5

Assume that the reaction for the formation of gaseous hydrogen fluoride from hydrogen and fluorine has an equilibrium constant of  $1.15 \times 10^2$  at a certain temperature. In particular experiment 3.00 mol of each component was added to a 1.50 L flask. Calculate the equilibrium concentration of all species.

A 1.00 L flask was filled with 2.00 mol gaseous  $\text{SO}_2$  and 2.00 mol gaseous  $\text{NO}_2$  and heated. After equilibrium was established, it was found that 1.30 mol gaseous  $\text{NO}$  was present. Assume the reaction



Occurs under these conditions. Calculate the value of the equilibrium constant,  $K$ , for this reaction

## Making Use of Le Chatelier's Principle

Le Chatelier's Principle is very useful in predicting how a system at equilibrium will respond to a change. It states that when a system at equilibrium is stressed (disturbed), the equilibrium shifts to alleviate the effect of the stress (disturbance).

There are three common ways equilibrium may be stressed/disturbed:

- Change in the \_\_\_\_\_ (or partial pressure) of one of the reactants or products.
- Change in the temperature.
- Change in the \_\_\_\_\_ of the container.

### Effect of Changes in Concentration (or Partial Pressure)

If a system at equilibrium is disturbed by the addition of a reactant (or the removal of a product), Le Chatelier's principle predicts that the equilibrium will shift \_\_\_\_\_. Shifting right will use up some of the added reactant (or replace some of the removed product), and therefore alleviate the stress/disturbance.

Similarly, if the disturbance is the removal of a reactant (or the addition of a product), Le Chatelier's principle predicts that the equilibrium will shift \_\_\_\_\_. Shifting left will replace some of the removed reactant (or use up some of the added product), and therefore alleviate the stress/disturbance.

Since concentrations of solids and liquids are constants and do not appear in expression  $K_{eq}$ , removing or adding some solid or liquid does not cause equilibrium shifts. However, shifts in the equilibrium do change the amount of solid present if it is a reactant or product!

### Effect of Changes in Temperature

If a reaction is endothermic ( $\Delta H > 0$ ), heat is absorbed in the forward reaction and released in the backward reaction; thus in endothermic reactions, heat behaves like a \_\_\_\_\_. Increasing the temperature (adding heat) shifts the reaction right, since that is the direction which absorbs heat and alleviates the stress/disturbance. Similarly, decreasing the temperature of an endothermic reaction shifts the reaction left.

Exothermic reactions ( $\Delta H < 0$ ) release heat in the forward direction, and absorb heat in the reverse direction; thus heat acts like a \_\_\_\_\_ in exothermic reactions. Increasing the temperature (adding heat) shifts the reaction left since that is the direction that absorbs heat and alleviates the stress/disturbance. Conversely, lowering the temperature shifts it right.

The equilibrium constant  $K$  is temperature dependent and the shifts above change its value.

Reaction Type	Role of heat	Effect of $T \uparrow$	Effect of $T \downarrow$
Endothermic ( $\Delta H > 0$ )	<i>reactants + heat <math>\rightarrow</math> products</i>	$K \uparrow$	$K \downarrow$
Exothermic ( $\Delta H < 0$ )	<i>reactants <math>\rightarrow</math> products + heat</i>	$K \downarrow$	$K \uparrow$

### Effect of Changes in the Volume of the Container

When the volume of a reaction vessel is decreased, the partial pressures of all gases in the container increase so the total pressure increases. Following Le Chatelier's principle, the reaction shifts to reduce the total pressure since that alleviates the stress/disturbance. This means that the shift is in the direction which contains the fewest moles gas.

Similarly, if the volume of the reaction vessel is increased, the total pressure decreases, and the shift is in the direction which contains the most moles gas.

In either case, if both sides of the equation have the same number of moles gas, the change in the volume of the container has no effect on the equilibrium.

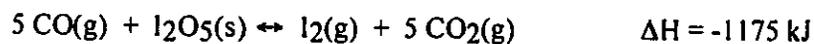
### Effect of the Addition of a Catalyst

Catalysts speed up the rate at which equilibrium is obtained, but have no effect on the magnitude of  $K$ . They increase both the forward and backward rate of reaction. Since catalysts do not appear in the net equation for a reaction, they are not involved in the expressions for  $K$ .

In the problems below, for each *change given in the first column* of the table, use *Le Chatelier's principle* to predict

- the *direction of shift* of the equilibrium.
- the *effect on the quantity* in the third column.

1. For the following reaction



for each change listed, predict the equilibrium shift and the effect on the indicated quantity.

	Change	Direction of Shift ( $\rightarrow$ ; $\leftarrow$ ; or <i>no change</i> )	Effect on Quantity	Effect (increase, decrease, or <i>no change</i> )
(a)	decrease in volume		$K_c$	
(b)	raise temperature		amount of CO(g)	
(c)	addition of I <sub>2</sub> O <sub>5</sub> (s)		amount of CO(g)	
(d)	addition of CO <sub>2</sub> (g)		amount of I <sub>2</sub> O <sub>5</sub> (s)	
(e)	removal of I <sub>2</sub> (g)		amount of CO <sub>2</sub> (g)	

2)

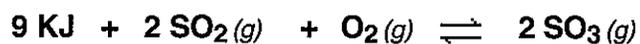


In which direction will the equilibrium shift in response to each change, and what will be the effect on the indicated quantity?

	Change	Direction of Shift ( $\rightarrow$ ; $\leftarrow$ ; or <i>no change</i> )	Effect on Quantity	Effect (increase, decrease, or <i>no change</i> )
(a)	add Ni(s)		Ni(CO) <sub>4</sub> (g)	
(b)	raise temperature		$K_c$	
(c)	add CO(g)		amount of Ni(s)	
(d)	remove Ni(CO) <sub>4</sub> (g)		CO(g)	
(e)	decrease in volume		Ni(CO) <sub>4</sub> (g)	
(f)	lower temperature		CO(g)	
(g)	remove CO(g)		$K_c$	

### 3 - Le Châtelier's Principle

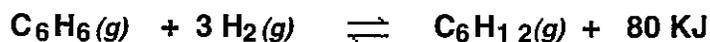
Use arrows (up / down) to indicate the affect of each of these disturbances (stresses) on the concentration of the reactants and products in this equilibrium:



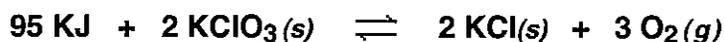
disturbance	affect on [SO <sub>2</sub> ]	affect on [O <sub>2</sub> ]	affect on [SO <sub>3</sub> ]
decrease [SO <sub>2</sub> ]			
increase [O <sub>2</sub> ]			
increase [SO <sub>3</sub> ]			
increase pressure			
decrease temp			



disturbance	affect on [N <sub>2</sub> ]	affect on [H <sub>2</sub> ]	affect on [NH <sub>3</sub> ]
increase [N <sub>2</sub> ]			
decrease [H <sub>2</sub> ]			
increase [NH <sub>3</sub> ]			
increase pressure			
increase temp			



disturbance	affect on [C <sub>6</sub> H <sub>6</sub> ]	affect on [H <sub>2</sub> ]	affect on [C <sub>6</sub> H <sub>12</sub> ]
decrease [C <sub>6</sub> H <sub>6</sub> ]			
increase [H <sub>2</sub> ]			
decrease [C <sub>6</sub> H <sub>12</sub> ]			
decrease pressure			
decrease temp			



disturbance	affect on [KClO <sub>3</sub> ]	affect on [KCl]	affect on [O <sub>2</sub> ]
decrease [KClO <sub>3</sub> ]			
decrease [KCl]			
increase [O <sub>2</sub> ]			
decrease pressure			
decrease temp			

In general, Le Châtelier's principle states that if a system at \_\_\_\_\_ is subjected to a \_\_\_\_\_, the equilibrium is \_\_\_\_\_ in the direction that tends to \_\_\_\_\_ the stress. If there is a change in pressure, this only affects equilibriums in which \_\_\_\_\_ are involved. If the total # of \_\_\_\_\_ of gases differs from reactants to \_\_\_\_\_, then under pressure, the equilibrium will \_\_\_\_\_ in the direction to produce \_\_\_\_\_ gases. The effect of changing the temperature of a \_\_\_\_\_ reaction depends on which of the opposing reaction is \_\_\_\_\_ and which is \_\_\_\_\_. According to the \_\_\_\_\_ principle, the addition of heat \_\_\_\_\_ the equilibrium so that the heat is absorbed (endothermic reaction \_\_\_\_\_). The removal of heat would favor \_\_\_\_\_ reactions.

**ANS (IAO+2):** endothermic, equilibrium, exothermic, exothermic, favored, gases, heavier, Le Châtelier, less, moles, products, relieve, reversible, shift, shifted, shifts, solubility, stress

1. As the number of effective collisions between reacting particles increases, the rate of reaction
- A) decreases                      C) remains the same  
B) increases
2. The energy needed to start a chemical reaction is called
- A) potential energy              C) activation energy  
B) kinetic energy                D) ionization energy
3. As the temperature increases, the rate of an exothermic reaction
- A) decreases                      C) remains the same  
B) increases
4. Given the reaction:
- $$X(g) + Y(g) \leftrightarrow XY(g) + \text{heat}$$
- As the temperature increases at constant pressure, the rate of the forward reaction
- A) decreases                      C) remains the same  
B) increases
5. As the number of moles per liter of a reactant in a chemical reaction increases, the number of collisions between the reacting particles
- A) decreases                      C) remains the same  
B) increases
6. Which statement explains why the speed of some chemical reactions is increased when the surface area of the reactant is increased?
- A) This change increases the density of the reactant particles.  
B) This change increases the concentration of the reactant.  
C) This change exposes more reactant particles to a possible collision.  
D) This change alters the electrical conductivity of the reactant particles.
7. Which factors must be equal in a reversible chemical reaction at equilibrium?
- A) the concentrations of the reactants and products  
B) the potential energies of the reactants and products  
C) the activation energies of the forward and reverse reactions  
D) the rates of reaction of the forward and reverse reactions

8. Which reaction has the equilibrium expression

$$K = \frac{[A][B]^2}{[AB_2]}$$

- A)  $AB_2(g) \leftrightarrow A(g) + 2 B(g)$   
B)  $2 AB(g) \leftrightarrow A(g) + B_2(g)$   
C)  $A(g) + 2 B(g) \leftrightarrow AB_2(g)$   
D)  $A(g) + B_2(g) \leftrightarrow 2 AB(g)$

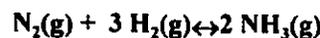
9. Given the reaction at equilibrium:



As the pressure increases at constant temperature, the number of moles of HCl

- A) decreases                      C) remains the same  
B) increases

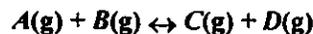
10. Given the reaction at equilibrium:



If the pressure is increased at constant temperature, there will be an increase in the number of moles of

- A)  $NH_3(g)$ , only  
B)  $N_2(g)$ , only  
C)  $H_2(g)$ , only  
D) both  $N_2(g)$  and  $H_2(g)$

11. Given the reaction at equilibrium:



The addition of a catalyst will

- A) shift the equilibrium to the right  
B) shift the equilibrium to the left  
C) increase the rate of forward and reverse reactions equally  
D) have no effect on the forward or reverse reactions

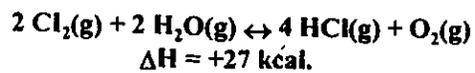
12. A saturated solution is represented by the equation



Which change will cause an increase in the amount of  $AgCl(s)$ ?

- A) a decrease in pressure  
B) an increase in temperature  
C) a decrease in the concentration of  $Ag^+(aq)$   
D) an increase in the concentration of  $Cl^-(aq)$

following system at equilibrium:



If the temperature of the system is increased at a constant pressure, the rate of the forward reaction will

- A) decrease                      C) remain the same  
B) increase

temperature will always

- A) increase the rate of reaction  
B) increase the concentration of products  
C) favor the exothermic reaction  
D) favor the endothermic reaction

15. Which reaction may be represented by the chemical equilibrium expression

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

- A)  $\text{A}(\text{aq}) + 2 \text{B}(\text{aq}) \leftrightarrow \text{C}(\text{s})$   
B)  $2 \text{C}(\text{s}) + \text{A}(\text{aq}) \leftrightarrow \text{B}(\text{aq})$   
C)  $2 \text{C}(\text{s}) \leftrightarrow \text{A}(\text{aq}) + 3 \text{B}(\text{aq})$   
D)  $\text{C}(\text{s}) + \text{B}(\text{aq}) \leftrightarrow 2 \text{A}(\text{aq})$



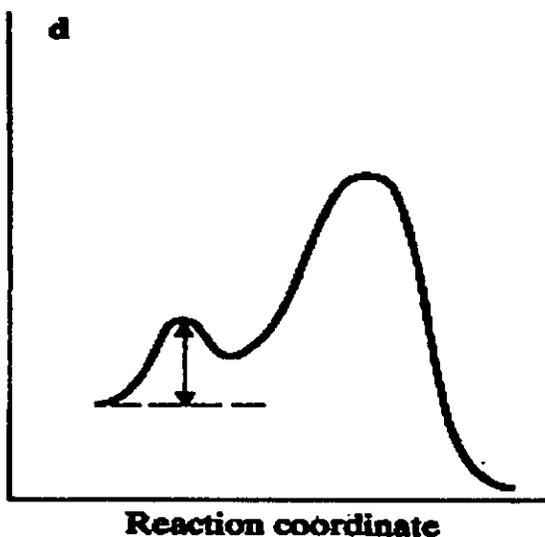
**THIS WILL BE A CUMULATIVE EXAM. UNIT 6 MATERIAL WILL PREDOMINATE.**

***From "Introduction to Reaction Rates, Equilibria" Notes:***

1. In the LEFT box below, draw a generic reaction coordinate diagram for an EXOTHERMIC reaction. In the RIGHT box below, draw a generic reaction coordinate diagram for an ENDOTHERMIC reaction. Label the following on each diagram: axes, location of reactants, transition state, and products, activation energy ( $E_A$ ), and overall energy change.

--	--

2. According to collision theory, what are TWO main requirements for a successful collision between particles?
3. List FOUR main ways to affect the rate of a chemical reaction, and FOR EACH, describe how the affect takes place.
4. Complete the following for the reaction coordinate diagram below:



- Label the Y-axis.
- Label the elementary steps and the intermediate
- Label the activation energies for each elementary step
- Label the total energy change
- Which elementary step is RATE-DETERMINING and WHY?

5. Please categorize the following statements as true (T) or false (F). Correct all false statements to make them true.

- At equilibrium, the forward reaction rate equals the reverse reaction rate
- At equilibrium, the amount of reactant present equals the amount of product present
- In order for equilibrium to be achieved, a closed system is required
- Equilibria can be achieved with both reversible and irreversible processes
- At equilibrium, concentrations of reactants and products do not necessarily remain constant

6. List, define, and provide an example of each of the three types of equilibria discussed in class.
7. Write the expressions for  $K_c$  for the following reactions:
- $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$
  - $Co(s) + 2H^+(aq) \rightleftharpoons Co^{2+}(aq) + H_2(g)$
  - $Ti(s) + 2Cl_2(g) \rightleftharpoons TiCl_4(l)$
8. Why do we normally exclude pure solids and pure liquids from equilibrium constant expressions?
9. A 1.20-L flask at a constant temperature contains an equilibrium mixture of 0.0168 mol of  $N_2$ , 0.2064 mol of  $H_2$ , and 0.0143 mol of  $NH_3$ . Calculate the equilibrium constant ( $K_c$ ) for the reaction:  $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

Are reactants or products favored in this equilibrium at this particular temperature? How do you know?

**From "Equilibrium Formation and Disturbance" Notes:**

**PLEASE SHOW WORK FOR # 8 – 9 ON SEPARATE PAGE AND ATTACH**

10.  $CaCl_2(s) \rightleftharpoons Ca^{2+}(aq) + 2Cl^-(aq)$   
 At a temperature of 275 K, the above reaction has a  $K_c$  value of 5.12. Suppose a piece of  $CaCl_2(s)$  is fully dissolved in water at 275 K. What will be the molarities of  $Ca^{2+}(aq)$  and  $Cl^-(aq)$  when the system reaches equilibrium?
- ANSWER:**
11.  $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$   
 How many moles of HI are present at equilibrium when 2.0 moles of  $H_2$  are mixed with 1.0 mole of  $I_2$  in a 0.50 L container and allowed to react at 448°C. At this temperature,  $K_c = 50.0$ .

**ANSWER:**

12.  $5CO(g) + I_2O_5(s) \rightleftharpoons I_2(g) + 5CO_2(g) \quad q = -1175 \text{ kJ}$   
 For each change listed below, predict the equilibrium shift and the effect on the indicated quantity

	<u>Change</u>	<u>Direction of shift</u> ( $\rightarrow$ ; $\leftarrow$ ; or no effect)	<u>Quantity</u>	<u>Effect</u>
(a)	Decrease in volume		$K_c$	
(b)	Increase in temperature		CO (g) amount	
(c)	Decrease in temperature		$K_c$	
(d)	Addition of $I_2O_5(s)$		CO (g) amount	
(e)	Addition of $CO_2(g)$		$I_2O_5(s)$ amount	
(f)	Addition of a catalyst		$K_c$	
(g)	Removal of $I_2(g)$		$CO_2$ amount	
(h)	Addition of CO (g)		$K_c$	