

## CHEMISTRY 11 REVIEW

### Converting Mass to Moles:

If you have 16.7g of NaOH, convert it to moles.

Molar Mass of NaOH =

$$\frac{\quad}{\quad} =$$

How do you convert mass to moles?

### Converting Mass to Moles:

If you have 16.7g of NaOH, convert it to moles.

Molar Mass of NaOH =

$$\frac{16.7\text{g}}{40.0\text{g}} \times \frac{1\text{ mol}}{1} = 0.418\text{ mol}$$

How do you convert mass to moles?

divide by the molar mass

### Converting Moles to Mass:

If you have 0.756 moles of HCN, what mass of HCN is present?

Molar Mass of HCN =

$$\frac{\quad}{\quad} =$$

How do you convert moles to mass?

### Converting Moles to Mass:

If you have 0.756 moles of HCN, what mass of HCN is present?

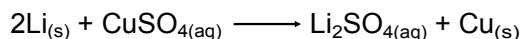
Molar Mass of HCN =

$$\frac{0.756\text{ mol}}{1\text{ mol}} \times \frac{27.0\text{g}}{1} = 20.4\text{g}$$

How do you convert moles to mass?

multiply by the molar mass

### Stoichiometry:



What mass of Cu metal is produced if 14.5g of Li metal reacts?

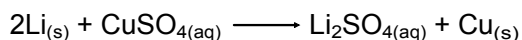
What should you do first?

When can you transfer from one substance to the other?

$$\frac{\quad}{\quad} \times \frac{\quad}{\quad} \times \frac{\quad}{\quad} =$$

# Kinetics

## Stoichiometry:



What mass of Cu metal is produced if 14.5g of Li metal reacts?

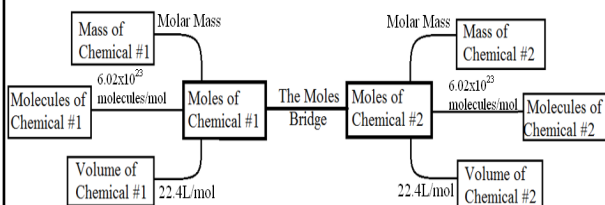
What should you do first? **convert to moles**

When can you transfer from one substance to the other?  
**once you have moles of one substance, use the mole ratio (from the balanced equation) to get moles of the other substance.**

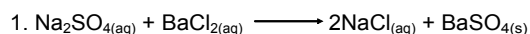
$$\frac{14.5\text{g Li}}{6.9\text{g Li}} \times \frac{1\text{ mol Li}}{2\text{ mol Li}} \times \frac{1\text{ mol Cu}}{1\text{ mol Cu}} \times \frac{63.5\text{g Cu}}{1\text{ mol Cu}} = 67\text{g Cu}$$

Draw the mole map for stoichiometry

Draw the mole map for stoichiometry



## Practice 1: Stoichiometry Exercises

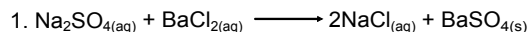


a) How many grams of NaCl would be produced from 80.0g of BaCl<sub>2</sub>?

b) How many grams of BaCl<sub>2</sub> would be required to produce 65.5g of NaCl?

2. When iron reacts with copper II chloride, solid copper and iron III chloride are formed. What mass of copper is produced if 0.594 moles of Fe react?

## Practice 1: Stoichiometry Exercises



a) How many grams of NaCl would be produced from 80.0g of BaCl<sub>2</sub>?

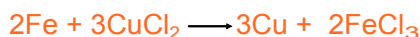
$$\frac{80.0\text{g BaCl}_2}{208.3\text{g BaCl}_2} \times \frac{1\text{ mol BaCl}_2}{1\text{ mol BaCl}_2} \times \frac{2\text{ mol NaCl}}{1\text{ mol BaCl}_2} \times \frac{58.5\text{g NaCl}}{1\text{ mol NaCl}} = 44.9\text{g NaCl}$$

b) How many grams of BaCl<sub>2</sub> would be required to produce 65.5g of NaCl?

$$\frac{65.5\text{g NaCl}}{58.5\text{g NaCl}} \times \frac{1\text{ mol NaCl}}{2\text{ mol NaCl}} \times \frac{1\text{ mol BaCl}_2}{1\text{ mol BaCl}_2} \times \frac{208.3\text{g BaCl}_2}{1\text{ mol BaCl}_2} = 117\text{g BaCl}_2$$

## Kinetics

2. When iron reacts with copper II chloride, solid copper and iron III chloride are formed. What mass of copper is produced if 0.594 moles of Fe react?

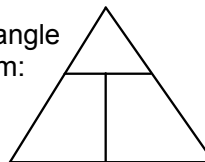


0.594 mol Fe	3 mol Cu	63.5g Cu	=	56.6g Cu
2 mol Fe	1 mol Cu			

### Molarity:

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{litres of solution}} = \frac{\text{mol}}{\text{L}}$$

Triangle form:



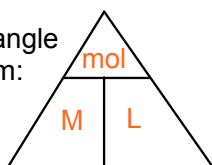
What is molarity a measure of?

What do square brackets mean? ex.  $[\text{NO}_3^-]$

### Molarity:

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{litres of solution}} = \frac{\text{mol}}{\text{L}}$$

Triangle form:



What is molarity a measure of?

concentration of a solution

What do square brackets mean? ex.  $[\text{NO}_3^-]$

square brackets mean 'concentration of'

ex. the concentration of nitrate

### Practice 2: Molarity Exercises

1. What is the molarity of a solution made by dissolving 2.45mol of potassium nitrate in 1.50L of solution?

2. How many moles of NaCl must be dissolved in 400.0mL of solution in order to make a 0.25M solution?

3. What is the volume if 0.555mol of MgS makes a 2.00M solution?

4. Find  $[\text{NaOH}]$  when 0.32mol NaOH is dissolved in  $5.00 \times 10^2$  mL of solution.

### Practice 2: Molarity Exercises

1. What is the molarity of a solution made by dissolving 2.45mol of potassium nitrate in 1.50L of solution?

$$[\text{KNO}_3] = \frac{2.45\text{mol}}{1.50\text{L}} = 1.63\text{M}$$

2. How many moles of NaCl must be dissolved in 400.0mL of solution in order to make a 0.25M solution?  $400.0\text{mL} = 0.4000\text{L}$

$$0.25\text{M} \times 0.4000\text{L} = 0.10 \text{ mol NaCl}$$

3. What is the volume if 0.555mol of MgS makes a 2.00M solution?

$$\frac{0.555\text{mol}}{2.00\text{M}} = 0.278\text{L or } 278\text{mL}$$

4. Find [NaOH] when 0.32mol NaOH is dissolved in  $5.00 \times 10^2$  mL of solution.

$$5.00 \times 10^2 \text{ mL} = 500\text{mL} = 0.500\text{L}$$

$$[\text{NaOH}] = \frac{0.32\text{mol}}{0.500\text{L}} = 0.64\text{M}$$

## Unit 1 - KINETICS

### I) Collision Theory

In a chemical reaction, what actually occurs at the particle level in order for reactants to make products?

This process is called 'collision theory'.

### I) Collision Theory

In a chemical reaction, what actually occurs at the particle level in order for reactants to make products?

reactant particles collide with one another. The collision may cause reactant bonds to break, and then new product bonds can form, making products.

This process is called 'collision theory'.

Particles are always moving as they always have kinetic energy (KE). Solids only vibrate as they have a low amount of KE. Liquid particles have a medium amount of KE and slide past one another. Gases have a large amount of KE and move around very quickly. Because of all this movement, particles are always colliding with one another.

[http://www.harcourtschool.com/activity/states\\_of\\_matter/](http://www.harcourtschool.com/activity/states_of_matter/)

Two things can happen as the result of a collision:

1) The reactant particles collide but either don't have \_\_\_\_\_ or don't have \_\_\_\_\_ to cause reactant bonds to break. Therefore, they bounce off one another and no reaction occurs.

2) The reactant particles collide with the correct geometry and with sufficient kinetic energy to cause reactant bonds to break and product bonds to form. This is called an \_\_\_\_\_.

When a collision has enough kinetic energy to cause the reactant bonds to break, it is said to have reached the **threshold energy**.

[http://www.media.pearson.com.au/schools/cw/au\\_sch\\_derry\\_ibest/1/molecular/Collisions1508.html](http://www.media.pearson.com.au/schools/cw/au_sch_derry_ibest/1/molecular/Collisions1508.html)

## Kinetics

Two things can happen as the result of a collision:

1) The reactant particles collide but either don't have **sufficient kinetic energy** or don't have **correct collision geometry** to cause reactant bonds to break. Therefore, they bounce off one another and no reaction occurs.

2) The reactant particles collide with the correct geometry and with sufficient kinetic energy to cause reactant bonds to break and product bonds to form. This is called an **effective collision**.

When a collision has enough kinetic energy to cause the reactant bonds to break, it is said to have reached the **threshold energy**.

Though it varies for each different reaction, a general rule in chemistry is that only a small percentage of collisions are effective (make products). Why do you think this is the case? Think of the blindfolded walkers' analogy to help.

Though it varies for each different reaction, a general rule in chemistry is that only a small percentage of collisions are effective (make products). Why do you think this is the case? Think of the blindfolded walkers' analogy to help.

Imagine 15 people in an empty classroom. Each is blindfolded. Some are told to walk slow, some faster, and some are told to run. How many collisions would cause a someone to fall over (meaning it is an effective collision)?

Answer: probably about 20% - a small percentage.

This is because you would need good collision geometry - a head on collision, AND sufficient energy in the collision.

It is difficult (in most reactions) to achieve BOTH sufficient KE and excellent collision geometry simultaneously, which is what is necessary for an effective collision. Often, you may have one but not the other.

What could you say about a reaction that has many effective collisions per second?

What could you say about a reaction that has few effective collisions per second?

What could you say about a reaction that has no effective collisions per second?

What could you say about a reaction that has many effective collisions per second?

It is a fast reaction, meaning it has a high reaction rate. The more effective collisions per second the higher the reaction rate.

What could you say about a reaction that has few effective collisions per second?

It is slow and has a low reaction rate.

What could you say about a reaction that has no effective collisions per second?

It does not occur spontaneously (on its own) and has a non-existent reaction rate. Maybe if you heat the reaction up it will start to occur

**The speed of the reaction, otherwise stated as the reaction rate, depends on the amount of effective collisions per unit time. Effective collisions are what produce product particles.**

# Kinetics

## II) Investigating Rate:

Reaction Rate:

Mathematically, rate =

$$\Delta =$$

What are some examples of a 'quantity' that can change in a reaction?

Therefore, what are some valid units for rate?

## II) Investigating Rate:

Reaction Rate: the speed of a reaction

$$\text{Mathematically, rate} = \frac{\Delta \text{ quantity}}{\Delta \text{ time}}$$

$\Delta$  = greek letter 'delta' meaning 'change in'

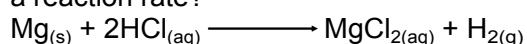
What are some examples of a 'quantity' that can change in a reaction?

mass of reactant (decrease), mass of product (increase), temperature, pressure of a gas, concentration of solution

Therefore, what are some valid units for rate?

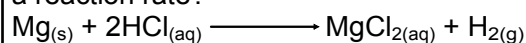
$$\frac{\text{g}}{\text{s}} \quad \frac{\text{mg}}{\text{s}} \quad \frac{\text{g}}{\text{min}} \quad \frac{\text{M}}{\text{s}} \quad \frac{\text{C}}{\text{s}} \quad \frac{\text{atm}}{\text{min}} \quad \frac{\text{kg}}{\text{hr}} \quad \text{etc}$$

Suppose we observe the following reaction in the lab. What are some quantities we could measure over a period of time in order to obtain a reaction rate?



<http://www.absorblearning.com/media/attachment.action?quick=vp&att=2272>

Suppose we observe the following reaction in the lab. What are some quantities we could measure over a period of time in order to obtain a reaction rate?



-mass loss of  $\text{Mg}_{(s)}$  per unit time by weighing the Mg piece and timing the reaction until the Mg disappears

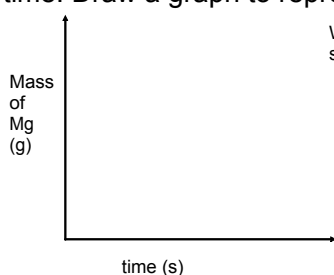
-mass increase of  $\text{H}_2$  gas per unit time by putting the entire 'open' reaction on the mass balance and timing the reaction

<http://www.absorblearning.com/media/attachment.action?quick=vp&att=2272>

-volume increase of  $\text{H}_2$  gas per unit time by trapping the  $\text{H}_2$  gas in a eudiometer tube and timing the reaction

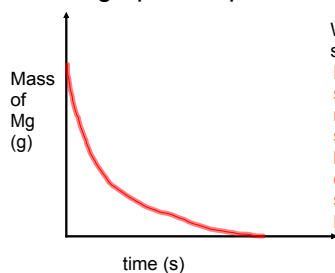


Rate can be measured by measuring the consumption of a reactant, or measuring the production of a product, all the while timing the reaction. Suppose the rate for the reaction under discussion was determined by measuring the mass loss of the reactant  $\text{Mg}_{(s)}$  per unit time. Draw a graph to represent this:



Why does the graph have this shape?

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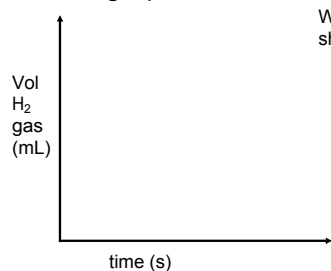


Why does the graph have this shape?

Mg will react very quickly at the start due to the large amount of reactant collisions. Once Mg starts to get used up, there will be less and less reactant collisions, so the reaction will get slower and slower until all of the Mg is gone.

# Kinetics

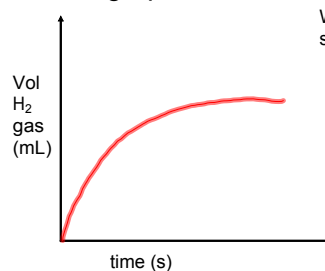
Suppose the rate is determined by measuring the increase in volume of the product H<sub>2</sub> gas. Draw a graph for this:



Why does the graph have this shape?

[http://www.media.pearson.com.au/schools/cw/au\\_sch\\_derry\\_ibcs1\\_1/int/ratesORChange/f1/1301.html](http://www.media.pearson.com.au/schools/cw/au_sch_derry_ibcs1_1/int/ratesORChange/f1/1301.html)

Suppose the rate is determined by measuring the increase in volume of the product H<sub>2</sub> gas. Draw a graph for this:



Why does the graph have this shape?

H<sub>2</sub> gets produced very quickly to start due to the large amount of reactant collisions. As the reactants get used up and the number of collisions decrease, H<sub>2</sub> is not produced as fast.

Remember from math that:

$$\text{slope} = \frac{\text{rise}}{\text{run}}$$

What is the rise (the y axis) in each of the graphs we've just drawn?

A change in \_\_\_\_\_.

What is the run (the x axis) in each?

A change in \_\_\_\_\_.

Therefore, what does the slope of these graphs represent?

$$\text{_____} = \frac{\text{_____}}{\text{_____}} = \text{_____}$$

Remember from math that:

$$\text{slope} = \frac{\text{rise}}{\text{run}}$$

What is the rise (the y axis) in each of the graphs we've just drawn?

A change in quantity.

What is the run (the x axis) in each?

A change in time.

Therefore, what does the slope of these graphs represent?

$$\frac{\text{rise}}{\text{run}} = \frac{\Delta \text{quantity}}{\Delta \text{time}} = \text{rate}$$

What are the units of rate for each of the previous two graphs?

First graph:

Second graph:

Look at the two graphs again. What do you notice about the magnitude of the slope as your reaction proceeds. What does this tell you about the rate as a reaction proceeds?

What are the units of rate for each of the previous two graphs?

First graph:  $\frac{\text{grams}}{\text{second}}$

Second graph:  $\frac{\text{mL}}{\text{second}}$

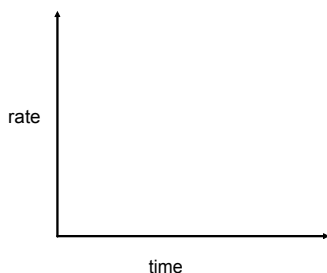
Look at the two graphs again. What do you notice about the magnitude of the slope as your reaction proceeds. What does this tell you about the rate as a reaction proceeds?

The graphs start with large slope magnitudes. Then the slopes continually decrease until they become zero. Thus, for any reaction, rate is always high at the start, and continually decreases as the reaction proceeds. When a reaction is finished, the rate is zero.

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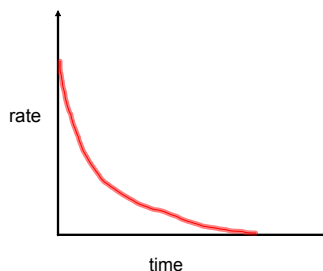
# Kinetics

Draw a graph of rate (y axis) vs. time for any reaction:



As reactants get used up, there are less and less collisions between reactant particles, therefore less and less effective collisions, so reaction rate is always decreasing!

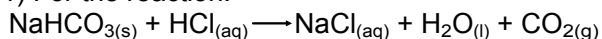
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## Assignment 1:

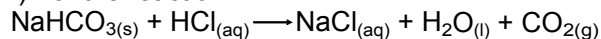
1) For the reaction:



Describe three things that you could measure while timing the reaction to calculate a reaction rate. For each, state the quantity being measured and what substance is being measured.

## Assignment 1:

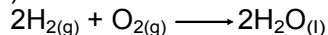
1) For the reaction:



Describe three things that you could measure while timing the reaction to calculate a reaction rate. For each, state the quantity being measured and what substance is being measured.

- mass decrease in  $\text{NaHCO}_3$  (weigh before and then react all of it)
- mass increase in  $\text{CO}_2$  gas (put open system on mass balance)
- volume increase in  $\text{CO}_2$  gas (use a eudiometer tube)
- pressure increase of closed system (use a barometer)
- temperature change in system (every reaction either gains or gives up energy, so use a thermometer)
- would be too difficult to monitor molarity changes in  $\text{HCl}$  or  $\text{NaCl}$

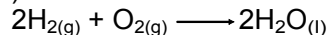
2) For the reaction:



Describe three things that you could measure while timing the reaction to calculate a reaction rate. For each, state the quantity being measured and what substance is being measured.

3) Hebden page 10, question 17

2) For the reaction:



Describe three things that you could measure while timing the reaction to calculate a reaction rate. For each, state the quantity being measured and what substance is being measured.

- decrease in pressure in a closed system (use a barometer)
- temperature change (use a thermometer)
- volume increase of liquid water (use a graduated cylinder)
- mass increase of liquid water (use a mass balance)

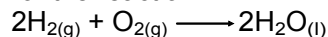
3) Hebden page 10, question 17

- answer in back of Hebden



## Kinetics

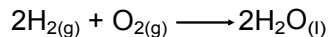
4) For the reaction:



Draw a graph for:

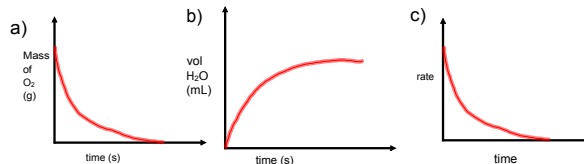
- Mass loss of  $\text{O}_2$  per unit time as the reaction proceeds
- volume gain of  $\text{H}_2\text{O}$  per unit time
- rate per unit time

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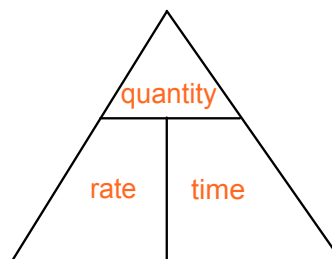


### III) Calculating Rate

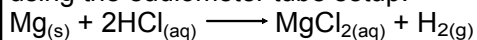
What is the rate equation triangle?

### III) Calculating Rate

What is the rate equation triangle?



Calculate the rate of the demonstration reaction using the eudiometer tube setup.



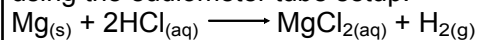
volume of  $\text{H}_2$  gas:

time:

average rate:

Why is the rate we calculated an 'average' rate?

Calculate the rate of the demonstration reaction using the eudiometer tube setup.



volume of  $\text{H}_2$  gas:

time:

average rate:

Why is the rate we calculated an 'average' rate?

Rate in a reaction is ALWAYS decreasing, so it is different from one moment to the next. So any rate that is calculated over a period of time is always an average rate.

## Kinetics

Calculate the mass of the piece of magnesium used:

We will work through questions 1, 2, and 4 on page 2 of Hebden, and then #19 on page 11.

1)

2)

We will work through questions 1, 2, and 4 on page 2 of Hebden, and then #19 on page 11.

$$1) \frac{5.0 \text{g Mg}}{150 \text{s}} = 0.033333 \frac{\text{gMg}}{\text{s}} \times 60 = 2.0 \frac{\text{gMg}}{\text{min}}$$

$$2) \text{time} = \frac{\text{quantity}}{\text{rate}} = \frac{45.0 \text{g CaCO}_3}{2.35 \frac{\text{g}}{\text{min}}} = 19.1 \text{ min}$$

4)

- 4) a) acceptable  
b) unacceptable  
c) acceptable  
d) unacceptable  
e) acceptable  
f) acceptable

\*Numerator must be a quantity.  
Denominator must be a time unit.

19) (a)

(b)

(c)

## Kinetics

19) (a)  $\text{rate} = \frac{\Delta q}{\Delta t} = \frac{31.0 - 24.6\text{g}}{60 - 0\text{s}} = 0.11 \text{ g/s}$

(b)  $\text{rate} = \frac{\Delta q}{\Delta t} = \frac{20.2 - 17.4\text{g}}{180 - 120\text{s}}$

- (c) Since the reaction rate is always decreasing (due to less and less reactant collisions as reactants get used up), the average rate later in a reaction will always be less than an average rate from earlier in the reaction.

### Assignment 2: Rate Calculations

1)

Time (s)	Mass of Pb (g)
0	65
15	52
30	41
45	32
60	25

The above data for a reaction in which Pb is a reactant.

- a) Calculate the overall rate of the reaction.  
b) Calculate the rate from 0-30s and from 30-60s  
c) Explain why the 30-60 rate is lower.

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c) Explain why the 30-60 rate is lower.

a)  $\frac{\text{quantity}}{\text{time}} = \frac{65 - 25\text{g}}{60 - 0\text{s}} = \frac{40\text{g}}{60\text{s}} = 0.7\text{g/s}$

b)  $\frac{65 - 41\text{g}}{30 - 0\text{s}} = \frac{24\text{g}}{30\text{s}} = 0.8\text{g/s}$

$\frac{41 - 25\text{g}}{60 - 30\text{s}} = \frac{16\text{g}}{30\text{s}} = 0.5\text{g/s}$

- c) Rate is always decreasing. So any average rate later in a reaction is always lower than an average rate from earlier in the same reaction.

- 2) When  $\text{CaCO}_3$  reacts with HCl,  $\text{CO}_{2(g)}$  is produced. If 243mL of  $\text{CO}_2$  is produced in 22s, what is the rate of the reaction?

- 3) A 5.0g sample of Mg reacts completely with HCl in 120s. What is the average rate of this reaction in g Mg per minute?

- 2) When  $\text{CaCO}_3$  reacts with HCl,  $\text{CO}_{2(g)}$  is produced. If 243mL of  $\text{CO}_2$  is produced in 22s, what is the rate of the reaction?

$$\frac{243\text{mL}}{22\text{s}} = 11\text{mL/s}$$

- 3) A 5.0g sample of Mg reacts completely with HCl in 120s. What is the average rate of this reaction in g Mg per minute?

$$\frac{5.0\text{g}}{120\text{s}} = 0.04167\text{g/s} \times 60\text{s} = 2.5\text{g/min}$$

- 4) How long will it take (in s) for 45.0g of Mg to react with HCl, if the average rate of the reaction is 2.30g Mg/min?

## Kinetics

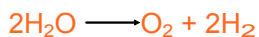
4) How long will it take (in s) for 45.0g of Mg to react with HCl, if the average rate of the reaction is 2.30g Mg/min?

$$\text{time} = \frac{\text{quantity}}{\text{rate}} = \frac{45.0\text{g}}{2.30\text{g/min}} = 19.565\text{min}$$

$$19.565\text{min} \times 60\text{s} = 1174\text{s} = 1170\text{s}$$

5) Electrolysis of  $\text{H}_2\text{O}$  produces  $\text{O}_2$  and  $\text{H}_2$ . If  $\text{O}_2$  gas is produced at an average rate of 28.5mL/min, calculate the consumption of  $\text{H}_2\text{O}$  in grams per second.

5) Electrolysis of  $\text{H}_2\text{O}$  produces  $\text{O}_2$  and  $\text{H}_2$ . If  $\text{O}_2$  gas is produced at an average rate of 28.5mL/min, calculate the consumption of  $\text{H}_2\text{O}$  in grams per second.



0.0285L $\text{O}_2$	1 mol $\text{O}_2$	2 mol $\text{H}_2\text{O}$	18.0g $\text{H}_2\text{O}$
	22.4L $\text{O}_2$	1 mol $\text{O}_2$	1 mol $\text{H}_2\text{O}$

$$= 0.0458\text{g/min divided by } 60 =$$

$$7.63 \times 10^{-4} \text{ g/s}$$

6) Given the reaction:  $\text{H}_{2(g)} + \text{Cl}_{2(g)} \longrightarrow 2\text{HCl}_{(g)}$   
a) If 2.32g of HCl are produced in 4.0min, what is the rate of reaction in mol HCl/second?

b) If  $\text{H}_2$  is used up at a rate of 30.0mol/s, at what rate is HCl produced in g/min?

6) Given the reaction:  $\text{H}_{2(g)} + \text{Cl}_{2(g)} \longrightarrow 2\text{HCl}_{(g)}$   
a) If 2.32g of HCl are produced in 4.0min, what is the rate of reaction in mol HCl/second?

$$\frac{2.32\text{g HCl}}{36.5\text{g HCl}} \times \frac{1 \text{ mol HCl}}{1 \text{ mol HCl}} = 0.06356 \text{ mol HCl in } 4.0\text{min}$$

$$\frac{0.06356\text{mol}}{240\text{s}} = 2.6 \times 10^{-4} \text{ mol/s}$$

b) If  $\text{H}_2$  is used up at a rate of 30.0mol/s, at what rate is HCl produced in g/min?

$$\frac{30.0\text{mol H}_2}{1 \text{ mol H}_2} \times \frac{2 \text{ mol HCl}}{1 \text{ mol H}_2} \times \frac{36.5\text{g HCl}}{1 \text{ mol HCl}} = 2190\text{g/s} \times 60$$

$$= 1.31 \times 10^5 \text{ g/min}$$

7)  $\text{C}_5\text{H}_{12(g)} + 8\text{O}_{2(g)} \longrightarrow 5\text{CO}_{2(g)} + 6\text{H}_2\text{O}_{(g)}$   
If 17.6g of  $\text{C}_5\text{H}_{12}$  is burned in three seconds, calculate the rate of the reaction in grams  $\text{CO}_2$  per second.

## Kinetics

7)  $C_5H_{12(g)} + 8O_{2(g)} \longrightarrow 5CO_{2(g)} + 6H_2O_{(g)}$   
If 17.6g of  $C_5H_{12}$  is burned in three seconds, calculate the rate of the reaction in grams  $CO_2$  per second.

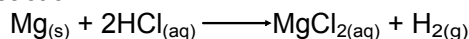
17.6g $C_5H_{12}$	1 mol $C_5H_{12}$	5 mol $CO_2$	44.0g $CO_2$
	72.0g $C_5H_{12}$	1 mol $C_5H_{12}$	1 mol $CO_2$

$$= \frac{53.778g}{3} = 17.9g/s$$

### IV) Factors Affecting Reaction Rates

[http://www.youtube.com/watch?v=XX9Xo6zm\\_kM](http://www.youtube.com/watch?v=XX9Xo6zm_kM)

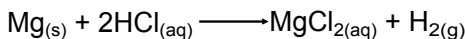
Think back to when we reacted a metal strip of magnesium with 2.0M HCl. What alterations could we make to increase the rate of the reaction?



### IV) Factors Affecting Reaction Rates

[http://www.youtube.com/watch?v=XX9Xo6zm\\_kM](http://www.youtube.com/watch?v=XX9Xo6zm_kM)

Think back to when we reacted a metal strip of magnesium with 2.0M HCl. What alterations could we make to increase the rate of the reaction?



- grind the metal Mg strip into powder to increase the surface area
- increase the temperature
- increase the concentration (molarity) of the HCl
- add a catalyst
- use a more reactive metal than Mg

### 1. The 'nature' of reactants influences reaction rate.

- some substances are simply more reactive than others; usually because their bonds will break at lower energy and/or it is easy to obtain correct collision geometry
- the number of effective collisions, and therefore the reaction rate, depend on the reactivity (the nature) of the reactants
- for example, magnesium is simply a more reactive metal than zinc

### 2. Concentration influences reaction rate.

-the greater the concentration of an aqueous or gaseous reactant, the more reactant particles per unit space, the more collisions there are between reactant particles, therefore the more effective collisions there will be, and hence an increased reaction rate

<http://www.absorblearning.com/media/attachment.action?quick=vy&att=2290>

- for solutions, concentration refers to \_\_\_\_\_.
- for gases, concentration refers to \_\_\_\_\_.

<http://www.absorblearning.com/media/attachment.action?quick=vz&att=2292>

### 2. Concentration influences reaction rate.

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<http://www.absorblearning.com/media/attachment.action?quick=vy&att=2290>

<http://www.kentchemistry.com/links/Kinetics/FactorsAffecting.htm>

scroll down to concentration

- for solutions, concentration refers to molarity.
- for gases, concentration refers to pressure.

<http://www.absorblearning.com/media/attachment.action?quick=vz&att=2292>

### 3. When reactants are heterogeneous, surface area influences reaction rate.

- heterogeneous reactions have reactants that are in different phases, meaning there is an interface between reactants

- if you expand the area of the interface (increase the \_\_\_\_\_), then there are more particles available to \_\_\_\_\_, resulting in more \_\_\_\_\_, therefore \_\_\_\_\_, therefore an \_\_\_\_\_.

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### 3. When reactants are heterogeneous, surface area influences reaction rate.

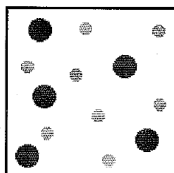
- heterogeneous reactions have reactants that are in different phases, meaning there is an interface between reactants

- if you expand the area of the interface (increase the surface area), then there are more particles available to collide, resulting in more overall collisions, therefore more effective collisions, therefore an increased reaction rate.

<http://www.absorblearning.com/media/attachment.action?quick=vw&att=2286>

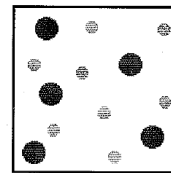
- a homogeneous reaction is one in which there is only \_\_\_\_\_, meaning the reactants are already as mixed as can be, meaning \_\_\_\_\_ has been maximized

- for example, *solutions* are homogeneous mixtures that display maximum surface area (already mixed as much as possible - down to the particle)



- a homogeneous reaction is one in which there is only one phase, meaning the reactants are already as mixed as can be, meaning surface area has been maximized

- for example, *solutions* are homogeneous mixtures that display maximum surface area (already mixed as much as possible - down to the particle)



### 4. Phase of reactants influences reaction rate.

- reactions involving aqueous reactants are instantaneous and are the fastest reactions. What are the possible reasons that this is the case?

- 1)
- 2)

So, in general, aqueous state reactants give the highest reaction rates, followed by \_\_\_\_\_, because of \_\_\_\_\_, then \_\_\_\_\_, and the slowest are generally \_\_\_\_\_ due to \_\_\_\_\_ and \_\_\_\_\_.

### 4. Phase of reactants influences reaction rate.

- reactions involving aqueous reactants are instantaneous and are the fastest reactions. What are the possible reasons that this is the case?

- 1) surface area is maximized
- 2) substantial amount of KE

So, in general, aqueous state reactants give the highest reaction rates, followed by gases, because of large KE, then liquids, and the slowest are generally solids due to low KE and low surface area.

## 5. Catalysts and Inhibitors influence reaction rate.

- catalysts increase reaction rate, not by producing more collisions, but by increasing the **probability of an effective collision**. Catalysts remain unchanged at the end of a reaction.
- inhibitors \_\_\_\_\_ reaction rate by inhibiting correct geometry and may or may not remain unchanged upon completion of the reaction. When are inhibitors useful?

## 5. Catalysts and Inhibitors influence reaction rate.

- catalysts increase reaction rate, not by producing more collisions, but by increasing the **probability of an effective collision**. Catalysts remain unchanged at the end of a reaction.
- inhibitors decrease reaction rate by inhibiting correct geometry and may or may not remain unchanged upon completion of the reaction. When are inhibitors useful?

food preservatives

## 6. Temperature influences reaction rate.

- an increase in temperature \_\_\_\_\_ reaction rate
- suppose a reaction between two reactants has 100 collisions per second, and 10 are effective. What is the % of effective collisions? \_\_\_\_\_
- if temperature is increased, the particles move faster, so there will be more collisions per second. Also, the collisions are harder due to the particles moving faster, so there will be a higher % of effective collisions.
- thus, by increasing temperature, maybe there will now be 150 collisions per second, and now maybe 30 are effective. What is the new % of effective collisions? \_\_\_\_\_

<http://www.absorblearning.com/media/attachment.action?quick=w2&att=2298>

## 6. Temperature influences reaction rate.

- an increase in temperature increases reaction rate
- suppose a reaction between two reactants has 100 collisions per second, and 10 are effective. What is the % of effective collisions? 10%
- if temperature is increased, the particles move faster, so there will be more collisions per second. Also, the collisions are harder due to the particles moving faster, so there will be a higher % of effective collisions.
- thus, by increasing temperature, maybe there will now be 150 collisions per second, and now maybe 30 are effective. What is the new % of effective collisions? 20%

<http://www.absorblearning.com/media/attachment.action?quick=w2&att=2298>

An increase in temperature increases reaction rate due to two factors. Explain each:

1)

2)

<http://www.kentchemistry.com/links/Kinetics/FactorsAffecting.htm>  
scroll down to temp

An increase in temperature increases reaction rate due to two factors. Explain each:

1) The particles will be moving faster, therefore there will be more overall collisions, therefore more effective collisions, and hence an increased reaction rate.

In the example, the overall collisions went from 100 per second to 150 per second.

2) The particles will be moving faster, thus they will collide even harder, therefore a higher percentage of effective collisions, therefore an increased reaction rate.

In the example, the % of effective collisions went from 10% to 20%.

In summary, increasing temperature will increase the rate of the reaction because of two reasons: more frequent and more forceful collisions.

## Kinetics

- a general rule for temperature change is that every 10 degree increase causes the reaction rate to double

### Threshold Energy:

Remember, the threshold energy (TE) is the \_\_\_\_\_ amount of KE needed to cause \_\_\_\_\_.

What types of bonds would have smaller threshold energies?

What types of bonds would have larger TE?

- a general rule for temperature change is that every 10 degree increase causes the reaction rate to double

### Threshold Energy:

Remember, the threshold energy (TE) is the \_\_\_\_\_ amount of KE needed to cause \_\_\_\_\_.

What types of bonds would have smaller threshold energies?

single bonds as they are, in general, easiest to break

What types of bonds would have larger TE?

triple bonds - take the most energy to break

An increase in concentration of a reactant causes an increase in reaction rate. Does an increase in concentration decrease the TE (cause the reactant bond energy to change)?

Does an increase in surface area decrease the TE?

Does an increase in temperature decrease the TE?

An increase in concentration of a reactant causes an increase in reaction rate. Does an increase in concentration decrease the TE (cause the reactant bond energy to change)?

No, but it allows for more collisions to attain the TE, thereby increasing the amount of effective collisions, and hence the reaction rate.

Does an increase in surface area decrease the TE?

No, same as above.

Does an increase in temperature decrease the TE?

No, but an increase in temperature allows for more collisions to attain the TE, as there will be more collisions and they will be harder.

### 'Factors' that can change the TE:

'Nature' of reactants is an actual change in reactant, so the bond energy of the reactants will be different, changing the TE.

Adding a catalyst will lower the TE as well, as a catalyst helps particles collide and bonds to break at a lower energy than usual, either due to a reactant collision with the catalyst, or excellent collision geometry.

### Assignment 3:

*Note: When you are asked to explain something using 'collision theory' you must comment on how overall collisions, effective collisions, and reaction rate are affected.*

1) When a 1.0g cube of Zn reacts with 1.0M HCl, the reaction rate is slower than when a 1.0g cube of Mg reacts with 1.0M HCl. What would be the reason for this?



## Kinetics

### Assignment 3:

*Note: When you are asked to explain something using 'collision theory' you must comment on how overall collisions, effective collisions, and reaction rate are affected.*

1) When a 1.0g cube of Zn reacts with 1.0M HCl, the reaction rate is slower than when a 1.0g cube of Mg reacts with 1.0M HCl. What would be the reason for this?

'Nature' of Reactants is the reason. Mg is simply more reactive than Zn, meaning it has a lower threshold energy that can be attained easier due to collisions. Thus, more of the collisions will be effective, resulting in a higher reaction rate.

2) If 1.0g pieces of copper are placed into beakers of 0.50M, 1.0M, and 1.5M nitric acid, explain using collision theory which reaction will proceed at the highest rate and why.

2) If 1.0g pieces of copper are placed into beakers of 0.50M, 1.0M, and 1.5M nitric acid, explain using collision theory which reaction will proceed at the highest rate and why.

1.5M nitric acid. It has the highest concentration, therefore it will cause the most overall collisions, therefore more effective collisions, and hence a higher reaction rate.

3) If 1.0g pieces of copper are placed into beakers of 0.50M and 1.0M nitric acid, and 1.0g of powdered copper are placed into beakers of 0.50M and 1.0M nitric acid, which reaction will be fastest, and why? And which will be slowest, and why?

3) If 1.0g pieces of copper are placed into beakers of 0.50M and 1.0M nitric acid, and 1.0g of powdered copper are placed into beakers of 0.50M and 1.0M nitric acid, which reaction will be fastest, and why? And which will be slowest, and why?

The fastest would be the 1.0g of powdered copper placed into the 1.0M nitric acid, for two reasons:  
- the 1.0M is the highest concentration nitric acid  
- the powdered copper has a higher surface area than the piece of copper, allowing for more exposure, so more overall collisions, more effective collisions, and a higher rate

The slowest would be the 1.0g piece of copper placed into the 0.50M nitric acid.  
- lowest concentration of nitric acid  
- lowest surface area

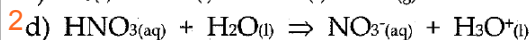
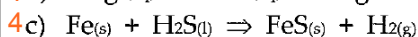
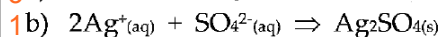
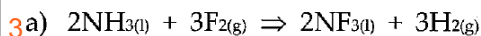
4) Rank each reaction from fastest (1) to slowest (4).

- $2\text{NH}_3(\text{l}) + 3\text{F}_2(\text{g}) \Rightarrow 2\text{NF}_3(\text{l}) + 3\text{H}_2(\text{g})$
- $2\text{Ag}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \Rightarrow \text{Ag}_2\text{SO}_4(\text{s})$
- $\text{Fe}(\text{s}) + \text{H}_2\text{S}(\text{l}) \Rightarrow \text{FeS}(\text{s}) + \text{H}_2(\text{g})$
- $\text{HNO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \Rightarrow \text{NO}_3^-(\text{aq}) + \text{H}_3\text{O}^+(\text{l})$

5) Hebden page 8, number 13

## Kinetics

4) Rank each reaction from fastest (1) to slowest (4).



5) Hebden page 8, number 13

key in the back of Hebden

6) If the rate of a reaction at 20 degrees is 0.040 mol/s, what is the rate at (a) 40 degrees? (b) 10 degrees?

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(a) double twice: 0.16 mol/s

(b) half: 0.020 mol/s

7) Suppose a reactant has a triple bond. How will its threshold energy compare with a reactant that has a single bond? How will this affect reaction rate?

8) Hebden page 12, number 21

7) Suppose a reactant has a triple bond. How will its threshold energy compare with a reactant that has a single bond? How will this affect reaction rate?

The triple bond reactant will have a much higher TE than the single bond reactant, therefore less of the collisions will be effective, resulting in a lower reaction rate.

8) Hebden page 12, number 21

answer in the back of Hebden

### V) Energy Changes in Reactions

What two things are necessary for a collision to be effective (for a reaction to occur)?

# Kinetics

## V) Energy Changes in Reactions

What two things are necessary for a collision to be effective (for a reaction to occur)?

- 1) sufficient KE (to attain the TE)
- 2) correct geometry

But how does the KE actually contribute to breaking reactant bonds?

As reactant particles approach, their KE (speed) begins to \_\_\_\_\_ due to \_\_\_\_\_ between the two reacting particles. Since energy cannot be destroyed, the KE is \_\_\_\_\_ to \_\_\_\_\_ energy (potential energy - PE), which is stored mainly in the bonds. At particle impact, the KE is at its \_\_\_\_\_ and the PE at its \_\_\_\_\_. If enough PE is built up in the reactant bonds, they will break and the product bonds form, resulting in an effective collision and creation of products.

But how does the KE actually contribute to breaking reactant bonds?  
As reactant particles approach, their KE (speed) begins to decrease due to electron repulsion between the two reacting particles. Since energy cannot be destroyed, the KE is converted to stored energy (potential energy - PE), which is stored mainly in the bonds. At particle impact, the KE is at its lowest and the PE at its highest. If enough PE is built up in the reactant bonds, they will break and the product bonds form, resulting in an effective collision and creation of products.

What happens if not enough PE is built up to break reactant bonds?

Whether products form or not, when particles move away from one another after impact, the KE \_\_\_\_\_, therefore the PE \_\_\_\_\_.

Thus, as KE  $\uparrow$ , PE \_\_\_\_\_ KE and PE are *inversely* related.  
and as KE  $\downarrow$ , PE \_\_\_\_\_

[http://www.wiley.com/college/chem/brady184764/resources/ch01/index\\_ch1\\_bysect.html](http://www.wiley.com/college/chem/brady184764/resources/ch01/index_ch1_bysect.html)

What happens if not enough PE is built up to break reactant bonds?

The collision will be unsuccessful and the reactants will move away from each other after the collision.

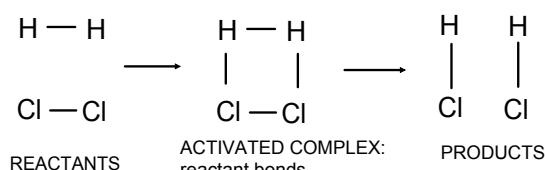
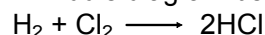
Whether products form or not, when particles move away from one another after impact, the KE increases, therefore the PE decreases.

Thus, as KE  $\uparrow$ , PE  $\downarrow$  KE and PE are *inversely* related.  
and as KE  $\downarrow$ , PE  $\uparrow$

[http://www.wiley.com/college/chem/brady184764/resources/ch01/index\\_ch1\\_bysect.html](http://www.wiley.com/college/chem/brady184764/resources/ch01/index_ch1_bysect.html)

If the collision is effective, there is a moment in time when reactant bonds are breaking, and simultaneously \_\_\_\_\_.

This structure is called the \_\_\_\_\_, as shown in middle diagram below:

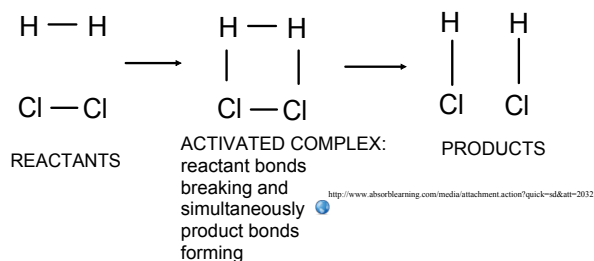
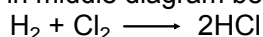


reactant bonds breaking and simultaneously product bonds forming

<http://www.absorblearning.com/media/attachment.action?quick=>

## Kinetics

If the collision is effective, there is a moment in time when reactant bonds are breaking, and simultaneously product bonds are forming. This structure is called the activated complex, as shown in middle diagram below:



Activated complexes are very short lived and very unstable. It is the point in the collision when PE is \_\_\_\_\_.

When discussing reaction energy from a KE standpoint, the threshold energy is the minimum amount of KE needed to cause the reactant bonds to break. But since the KE must transfer to PE first, the **activation energy ( $E_a$ )** is the minimum amount of PE needed to cause reactant bonds to break.

So,  
Threshold Energy (KE) = Activation Energy (PE)

Activated complexes are very short lived and very unstable. It is the point in the collision when PE is is at a maximum.

When discussing reaction energy from a KE standpoint, the threshold energy is the minimum amount of KE needed to cause the reactant bonds to break. But since the KE must transfer to PE first, the **activation energy ( $E_a$ )** is the minimum amount of PE needed to cause reactant bonds to break.

So,  
Threshold Energy (KE) = Activation Energy (PE)

Define Activation Energy ( $E_a$ ):

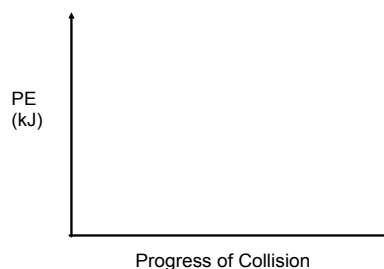
Define Activation Energy ( $E_a$ ):

The minimum PE required for reactant bonds to break

OR

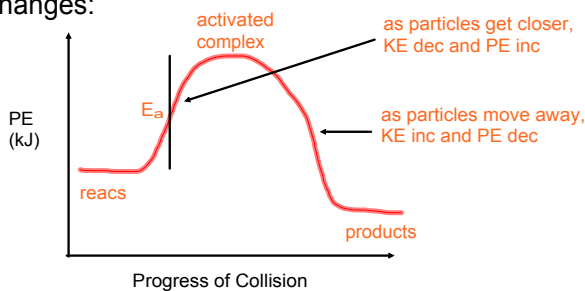
The potential energy difference between the activated complex and the reactants.

As a collision is occurring, sketch how PE changes:



# Kinetics

As a collision is occurring, sketch how PE changes:



## VI) Change in Enthalpy: $\Delta H$

Every reaction involves either an absorption, or a release, of energy, usually in the form of heat. Enthalpy is potential energy that may be evolved (released) or absorbed as heat.

What is an exothermic reaction?

<http://www.absorblearning.com/media/attachment.action?quick=vl&att=2264>

What is an endothermic reaction?

## VI) Change in Enthalpy: $\Delta H$

Every reaction involves either an absorption, or a release, of energy, usually in the form of heat. Enthalpy is potential energy that may be evolved (released) or absorbed as heat.

What is an exothermic reaction?

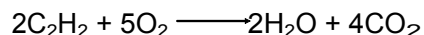
a reaction that involves the release of energy, usually in the form of heat, so product energy is lower than reactant energy

<http://www.absorblearning.com/media/attachment.action?quick=vl&att=2264>

What is an endothermic reaction?

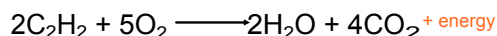
a reaction that involves the absorption of energy, usually by stealing heat from the surroundings. Thus, product energy is higher than reactant energy.

The following reaction is exothermic. Write an 'energy' term on the appropriate side:



Therefore, where is the 'energy' term written for an endothermic reaction?

The following reaction is exothermic. Write an 'energy' term on the appropriate side:

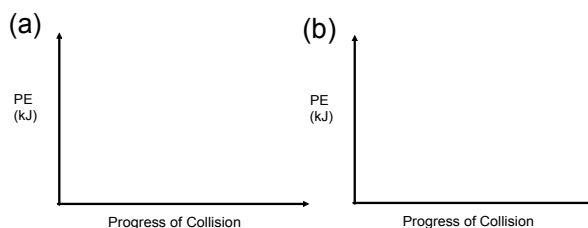


energy is released, so it is a product

Therefore, where is the 'energy' term written for an endothermic reaction?

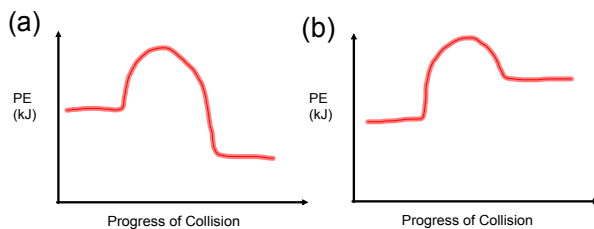
on the reactant side

Draw a PE curve for an exothermic reaction on graph (a), and an endothermic reaction on graph (b):



## Kinetics

Draw a PE curve for an exothermic reaction on graph (a), and an endothermic reaction on graph (b):



Enthalpy (analogous to energy) can be abbreviated as H, and is measured in Joules (J) or kilojoules (kJ). The change in enthalpy,  $\Delta H$ , for a reaction, is always calculated as:

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

So looking at the previous two graphs, what characteristic would  $\Delta H$  have for exothermic reactions?

For endothermic reactions?

[http://employees.oneonta.edu/viningwj/sims/bond\\_energy\\_dh\\_reaction\\_s.html](http://employees.oneonta.edu/viningwj/sims/bond_energy_dh_reaction_s.html)

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So looking at the previous two graphs, what characteristic would  $\Delta H$  have for exothermic reactions?

it is negative for exo reactions

For endothermic reactions?

it is positive for endo reactions

In summary, there are three ways to identify if a reaction is exothermic or endothermic. What are they?

1)

2)

3)

<http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/activa2.swf>

In summary, there are three ways to identify if a reaction is exothermic or endothermic. What are they?

1) the 'energy' term is in the products for exothermic, and reactants for endothermic

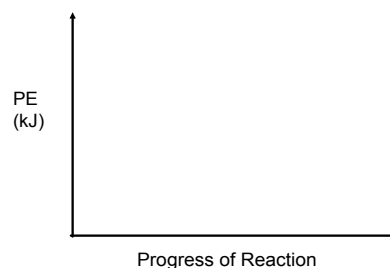
2) on a PE curve, the product PE is lower than reactant PE for exothermic, and higher for endothermic

3) for exothermic, the  $\Delta H$  is negative. For endothermic it's positive

ex. of exothermic:  $2\text{C}_2\text{H}_2 + 5\text{O}_2 \rightarrow 2\text{H}_2\text{O} + 4\text{CO}_2$   $\Delta H = -56\text{kJ}$

<http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/activa2.swf>

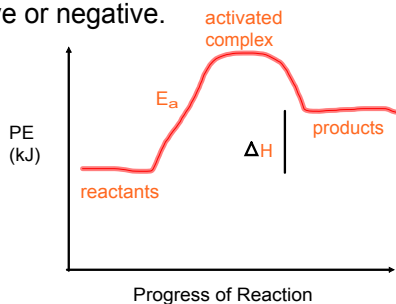
Draw an endothermic PE curve and label the reactants, products, activation energy, activated complex, and  $\Delta H$ . State whether the  $\Delta H$  is positive or negative.



<http://www.kentchemistry.com/links/Kinetics/FactorsAffecting.htm>

# Kinetics

Draw an endothermic PE curve and label the reactants, products, activation energy, activated complex, and  $\Delta H$ . State whether the  $\Delta H$  is positive or negative.



<http://phet.colorado.edu/en/simulation/reactions-and-rates>

What implications would a large  $E_a$  have on reaction rate?

What about a small  $E_a$ ?

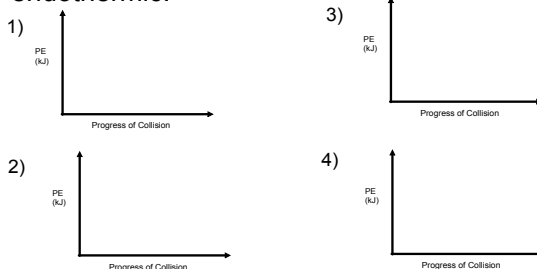
What implications would a large  $E_a$  have on reaction rate?

a large  $E_a$  means a large amount of PE is needed to break reactant bonds, thus implicating a low reaction rate (and possibly not a spontaneous reaction)

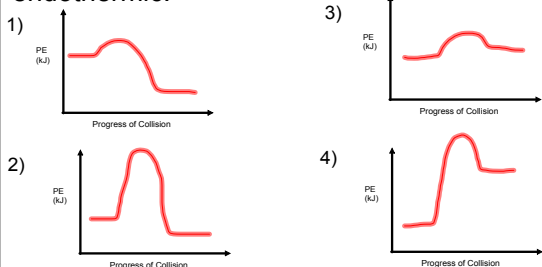
What about a small  $E_a$ ?

a small  $E_a$  means only a small amount of PE is needed to break reactant bonds, thus it will be easier to do so, resulting in a high reaction rate (and most likely a spontaneous reaction)

Draw 4 PE curves below. The first should represent a reaction that is spontaneous and exothermic. The second - non-spontaneous and exothermic. The third - spontaneous and endothermic. The fourth - non-spontaneous and endothermic.



Draw 4 PE curves below. The first should represent a reaction that is spontaneous and exothermic. The second - non-spontaneous and exothermic. The third - spontaneous and endothermic. The fourth - non-spontaneous and endothermic.



## Assignment 4:

*Tip: When drawing PE curves, never put the reactants or product energies at zero. Particles always have some potential energy.*

1) Hebden page 25, numbers 41-45

# Kinetics

## Assignment 4:

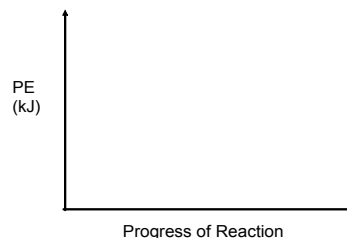
Tip: When drawing PE curves, never put the reactants or product energies at zero. Particles always have some potential energy.

1) Hebden page 25, numbers 41-45

Answers in the back of Hebden

2) Draw a PE curve with labeled axes that has reactant energy of 100kJ, an  $E_a$  of 200kJ, and a  $\Delta H$  of 150kJ. On the y axis, make a scale from 0 to 400kJ with increments of 50kJ.

- Is the reaction endothermic or exothermic?
- Are the products more stable than reactants?
- What is the PE of the activated complex?

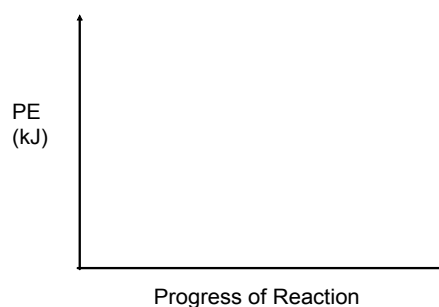


2) Draw a PE curve with labeled axes that has reactant energy of 100kJ, an  $E_a$  of 200kJ, and a  $\Delta H$  of 150kJ. On the y axis, make a scale from 0 to 400kJ with increments of 50kJ.

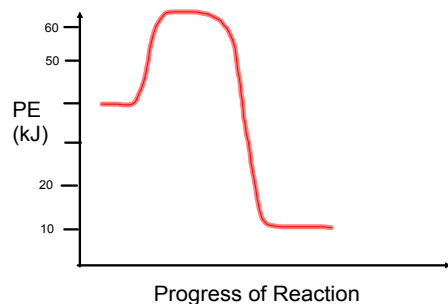
- Is the reaction endothermic or exothermic?
- Are the products more stable than reactants?
- What is the PE of the activated complex?



3) Sketch a PE diagram for a reaction that has an  $E_a = 20$ kJ, a  $\Delta H = -30$ kJ, and a product energy of 10kJ



3) Sketch a PE diagram for a reaction that has an  $E_a = 20$ kJ, a  $\Delta H = -30$ kJ, and a product energy of 10kJ



4) Does reaction rate depend on activation energy? Why or why not?

5) Does reaction rate depend on  $\Delta H$ ? Why or why not?



## Kinetics

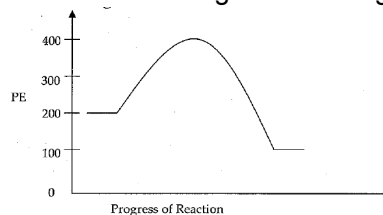
4) Does reaction rate depend on activation energy? Why or why not?

Yes. The higher the  $E_a$ , the lower the reaction rate. The lower the  $E_a$ , the higher the rate.

5) Does reaction rate depend on  $\Delta H$ ? Why or why not?

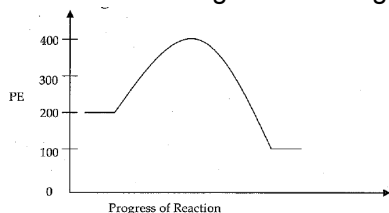
No. Reaction rate only depends on  $E_a$  (the uphill). It doesn't matter how far 'downhill' the energy goes after.

6) Read pages 24 & 25 in Hebden, then answer the question below using the following PE curve:



- Find the  $E_a$  of the forward reaction.
- Find the  $E_a$  of the reverse reaction.
- Find  $\Delta H$  for the fwd reaction. Endo or exo?
- Find  $\Delta H$  for the rev reaction. Endo or exo?
- What would happen to the  $E_a$  if the temperature is increased?

6) Read pages 24 & 25 in Hebden, then answer the question below using the following PE curve:



- Find the  $E_a$  of the forward reaction. 200kJ
- Find the  $E_a$  of the reverse reaction. 300kJ
- Find  $\Delta H$  for the fwd reaction. Endo or exo? -100kJ exo
- Find  $\Delta H$  for the rev reaction. Endo or exo? 100kJ endo
- What would happen to the  $E_a$  if the temperature is increased? nothing - but more particles would attain the  $E_a$

7) Explain in terms of energy changes (kinetic and potential) what occurs when two particles approach each other, collide, and move away as products. Be very specific and use correct vocabulary.

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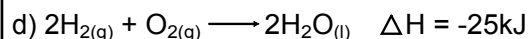
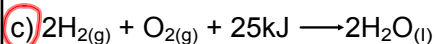
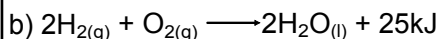
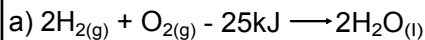
As particles move toward each other, their KE starts to decrease due to electron repulsion, thus their PE starts to increase. At collision, their KE is at its lowest and PE at its highest. Whether they react to make products or not, as the particles move away from each other after the collision, the KE starts to increase and the PE starts to decrease.

8) Which of the following is endothermic?

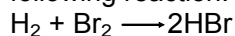
- $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) - 25\text{kJ} \longrightarrow 2\text{H}_2\text{O}(\text{l})$
- $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2\text{H}_2\text{O}(\text{l}) + 25\text{kJ}$
- $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) + 25\text{kJ} \longrightarrow 2\text{H}_2\text{O}(\text{l})$
- $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2\text{H}_2\text{O}(\text{l}) \quad \Delta H = -25\text{kJ}$

## Kinetics

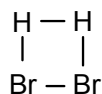
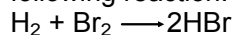
8) Which of the following is endothermic?



9) Draw a picture of an activated complex for the following reaction:

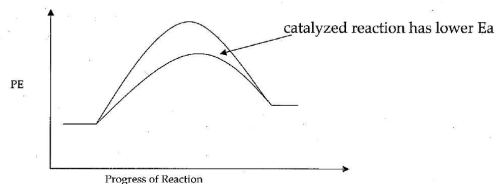


9) Draw a picture of an activated complex for the following reaction:



### VII) Catalysts and Inhibitors

Every type of reaction follows a specific and consistent pathway with a unique activated complex. When a catalyst is introduced to the reaction, the pathway changes and a different activated complex is formed with a lower  $E_a$ . Because of this lower  $E_a$ , a higher % of collisions are effective, thereby increasing the reaction rate.



Catalysts are involved in creating a different, lower energy activated complex, but remain unaltered at the end of the reaction. An inhibitor forms a new activated complex that has a larger activation energy, thereby decreasing reaction rate.

Catalysts lower the  $E_a$ , thereby increasing the rate, but do they alter the  $\Delta H$  for the reaction (look at the curve on the last slide)?

Catalysts are involved in creating a different, lower energy activated complex, but remain unaltered at the end of the reaction. An inhibitor forms a new activated complex that has a larger activation energy, thereby decreasing reaction rate.

Catalysts lower the  $E_a$ , thereby increasing the rate, but do they alter the  $\Delta H$  for the reaction (look at the curve on the last slide)?

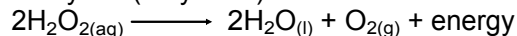
No. The catalyst or inhibitor affects the rate of the reaction, and the way the reaction occurs, but the reactants and products remain the same, so their energies remain the same, so the  $\Delta H$  is not affected.

## Kinetics

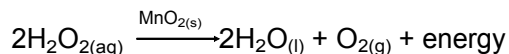
A catalyst is written above the reaction arrow in a reaction. However, sometimes, it's written as a reactant and a product (as it does not get used up in a reaction).

### Example

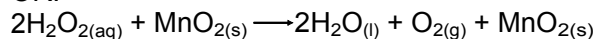
Uncatalyzed (very slow):



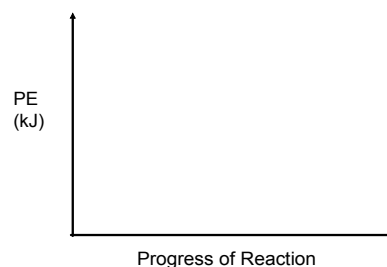
Catalyzed (fast):



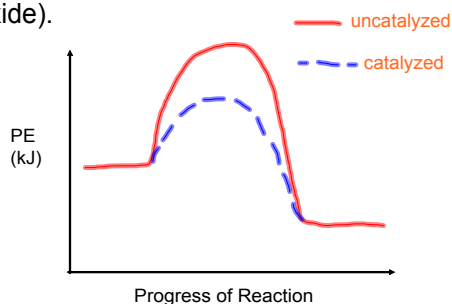
OR:



Draw a PE diagram for the uncatalyzed and catalyzed decomposition of  $\text{H}_2\text{O}_2$  (hydrogen peroxide).



Draw a PE diagram for the uncatalyzed and catalyzed decomposition of  $\text{H}_2\text{O}_2$  (hydrogen peroxide).



A heterogeneous catalyst is in a different phase than reactants, and usually adsorbs the reactant (holds it on its surface) to allow for better geometry with other reactants.

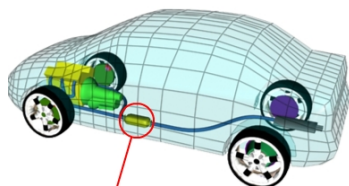
<http://www.kentchemistry.com/links/Kinetics/FactorsAffecting.htm>  
scroll down to Pt catalyst animation

A homogeneous catalyst is in the same phase as reactants, and usually alters the regular reaction pathway to a new, lower energy reaction pathway.

## IX) Specific Examples of Catalysts

### Automobile Catalytic Converters

Vehicles produce NO (nitric monoxide) as a result of the synthesis reaction of  $\text{N}_2$  and  $\text{O}_2$  from air in the high temperature engine where fuel combusts. Some NO then oxidizes to produce  $\text{NO}_2$  (nitrogen dioxide) which can form nitric acid ( $\text{HNO}_3$ ) when it reacts with water vapour in the atmosphere. CO (carbon monoxide) is also released in exhaust, a poisonous gas.



Catalytic Converter

In 1981, automobiles were equipped with a catalytic converter in the exhaust system, which consists of two chambers:

Chamber 1 - contains platinum and palladium catalysts that help convert CO to less harmful  $\text{CO}_2$

Chamber 2 - contains platinum and rhodium catalysts which help convert NO back to unharmed  $\text{N}_2$  &  $\text{O}_2$

<http://www.absorblearning.com/media/attachment.action?quick=ve&att=2250>

# Kinetics

## Autocatalysis

*Autocatalysis* occurs when the product of a reaction serves as a catalyst for the same reaction, creating a chain reaction. The reaction starts slowly but picks up speed once the catalyst begins to form.

Ex. reaction of permanganate ion ( $\text{MnO}_4^-$ ) by the oxalate ion ( $\text{C}_2\text{O}_4^{2-}$ ) in acid with  $\text{Mn}^{2+}$  autocatalysis



## Hydrogen Peroxide

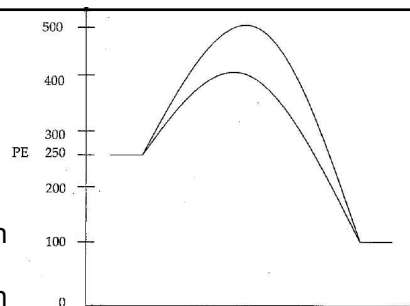
$\text{H}_2\text{O}_2$  is a harmful product formed mainly due to the breakdown of fatty acids in the liver of animals and via photorespiration of glucose in plants.  $\text{H}_2\text{O}_2$  kills cells non-specifically by disrupting cell membranes, thus it is used as a wound cleaner to kill bacteria. Within organisms, excess  $\text{H}_2\text{O}_2$  can be harmful.

Two known catalysts to decompose hydrogen peroxide into water and oxygen:

1. Manganese (IV) oxide ( $\text{MnO}_2$ ) - used in the lab
2. The enzyme (biological catalyst) *catalase*, found in the liver of animals. Used to decompose excess  $\text{H}_2\text{O}_2$  in the body.
3. *Catalase* is also present in plants such as potatoes.

## Assignment 5

- 1) Label and calculate the  $E_a$  for the:
  - a) uncatalyzed forward reaction
  - b) catalyzed forward reaction
  - c) uncatalyzed reverse reaction
  - d) catalyzed reverse reaction



- 2) Calculate the PE of the:
  - a) uncatalyzed activated complex
  - b) catalyzed activated complex

## Assignment 5

- 1) Label and calculate the  $E_a$  for the:

- a) uncatalyzed forward reaction **250kJ**
- b) catalyzed forward reaction **150kJ**
- c) uncatalyzed reverse reaction **400kJ**
- d) catalyzed reverse reaction **300kJ**

- 2) Calculate the PE of the:

- a) uncatalyzed activated complex **500kJ**
- b) catalyzed activated complex **400kJ**

- 3) Calculate the difference in activation energy for the uncatalyzed reaction and the catalyzed reaction.

- 4) Find the  $\Delta H$  for the:
  - a) uncatalyzed forward reaction
  - b) catalyzed forward reaction
  - c) What do you notice about the two?

## Kinetics

3) Calculate the difference in activation energy for the uncatalyzed reaction and the catalyzed reaction.

$$250\text{kJ} - 150\text{kJ} = 100\text{kJ}$$

4) Find the  $\Delta H$  for the:

a) uncatalyzed forward reaction  $-150\text{kJ}$

b) catalyzed forward reaction  $-150\text{kJ}$

c) What do you notice about the two?

They are the same.  $\Delta H$  is not affected by adding a catalyst because reactants and products don't change.

5) Can a catalyst cause an endothermic reaction to become exothermic? Why or why not?

6) A catalyst increases reaction rate in a different way than increasing temperature. Explain how each are different using collision theory.

5) Can a catalyst cause an endothermic reaction to become exothermic? Why or why not?

No. A catalyst does not affect  $\Delta H$  in any way.

6) A catalyst increases reaction rate in a different way than increasing temperature. Explain how each are different using collision theory.

A catalyst doesn't increase the number of overall collisions. However, it increases the percentage of effective collisions, thereby increasing reaction rate. It does this by lowering the activation energy.

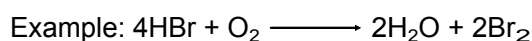
An increase in temperature increases the number of overall collisions AND increase the percentage of effective collisions, not by lowering activation energy, but due to harder collisions, more collisions attain the activation energy.

### PE Curve - Reversible Reactions Summary:

- $\Delta H$  for the forward and reverse reactions have the same magnitude but opposite signs
- $E_a$  is always larger for the endothermic reaction
- altering temperature, pressure, concentration, or surface area will have no effect on a potential energy diagram. Only the addition of a catalyst or inhibitor can do this by altering the  $E_a$ .

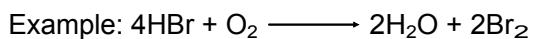
### VIII) Reaction Mechanisms

Reactions are expressed using a reaction equation. The equation gives information about the reactants and products, but very little information about the process that occurs to get from reactants to products. An equation suggests a one-step process, which is in reality seldom the case. Most reactions involve a series of steps, called a **reaction mechanism**.



The chemical reaction above actually has a three-step mechanism. Why doesn't this reaction occur in one step? (think about the collisions that would have to occur)

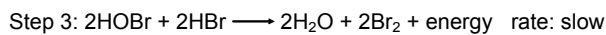
# Kinetics



The chemical reaction above actually has a three-step mechanism. Why doesn't this reaction occur in one step? (think about the collisions that would have to occur)

The probability of 5 particles (four HBr and one O<sub>2</sub>) colliding simultaneously with correct geometry and sufficient KE is extremely low. Thus, this reaction happens in three steps.

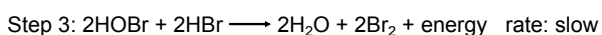
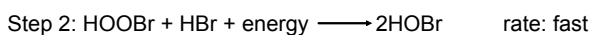
The three step mechanism for HBr reacting with O<sub>2</sub>:



Overall:  $\Delta\text{H} = \text{negative}$

a) Which step is the 'rate-determining step' and why?

The three step mechanism for HBr reacting with O<sub>2</sub>:



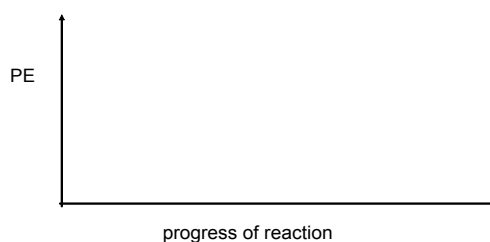
Overall:  $4\text{HBr} + \text{O}_2 \longrightarrow 2\text{H}_2\text{O} + 2\text{Br}_2$   $\Delta\text{H} = \text{negative}$

a) Which step is the 'rate-determining step' and why?

The 3rd step because it's the slowest. The slowest step always determines the speed of the entire mechanism.

b) Why are HOOBr and HOBr not part of the overall reaction equation?

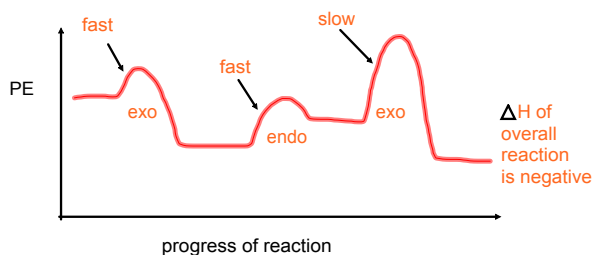
c) Draw a PE curve for the 3-step mechanism.



b) Why are HOOBr and HOBr not part of the overall reaction equation?

They are produced in one step and then subsequently used. Thus, they are not present at the beginning and not present at the end of the reaction.

c) Draw a PE curve for the 3-step mechanism.



HOOBr and HOBr are called \_\_\_\_\_.

Define **Reaction Intermediate**:

Notice that reaction intermediates are not part of the overall reaction because they 'cancel out'.

The first place you will see a reaction intermediate is on the \_\_\_\_\_ side and then subsequently on the \_\_\_\_\_ side.

# Kinetics

HOOBr and HOBr are called

reaction intermediates

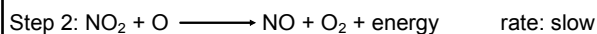
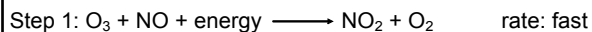
Define **Reaction Intermediate**:

substance(s) that are produced in one step and subsequently used up in a future step

Notice that reaction intermediates are not part of the overall reaction because they 'cancel out'.

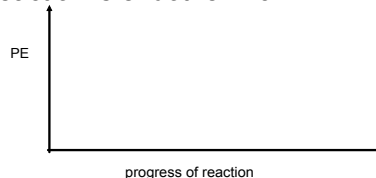
The first place you will see a reaction intermediate is on the product side and then subsequently on the reactant side.

Example:

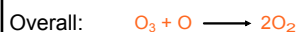
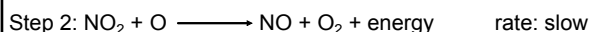
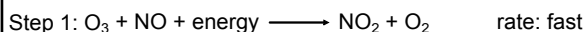


Overall:

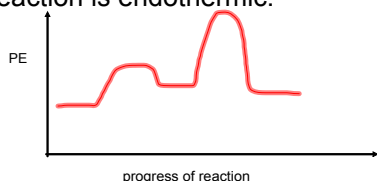
- What is the overall reaction?
- State any reaction intermediates.
- Which is the rate-determining step?
- Sketch a PE curve for the mechanism. The overall reaction is endothermic.



Example:



- What is the overall reaction?
- State any reaction intermediates.  $NO_2$
- Which is the rate-determining step? step 2
- Sketch a PE curve for the mechanism. The overall reaction is endothermic.



What was the role of NO in the last example?

**Catalyst:** a substance that \_\_\_\_\_ the rate without being consumed in the reaction.

The involvement of NO created a different, lower energy mechanism for the reaction. What happened with NO in step 1? What happened with NO in step 2?

Why must many catalyzed reactions involve a two-step mechanism?

What was the role of NO in the last example?

It acted as a catalyst.

**Catalyst:** a substance that increases the rate without being consumed in the reaction.

The involvement of NO created a different, lower energy mechanism for the reaction. What happened with NO in step 1? it was consumed What happened with NO in step 2? it was regenerated

Why must many catalyzed reactions involve a two-step mechanism?

It's used in one step and then reproduced in a subsequent step.

A catalyst is usually not part of the overall reaction (not a reactant or product) because it 'cancels out'.

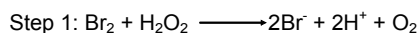
You'll see a catalyst on the \_\_\_\_\_ side first and then on the \_\_\_\_\_ side.

## Kinetics

A catalyst is usually not part of the overall reaction (not a reactant or product) because it 'cancels out'.

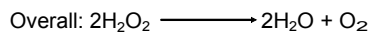
You'll see a catalyst on the reactant side first and then on the product side.

Example:



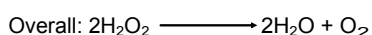
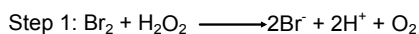
Step 2:

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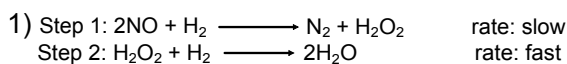
- Determine Step 2.
- Identify any reaction intermediates.
- Identify any catalysts.

Example:



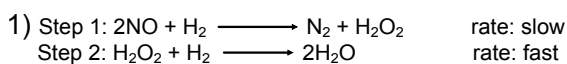
- Determine Step 2.
- Identify any reaction intermediates.  $\text{Br}^-$  and  $\text{H}^+$
- Identify any catalysts.  $\text{Br}_2$

### Assignment 6

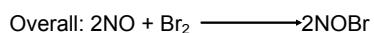
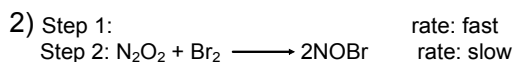


- What is the overall reaction?
- Which is the rate-determining step?
- Identify any reaction intermediates.

### Assignment 6



- What is the overall reaction?  
 $2\text{NO} + 2\text{H}_2 \longrightarrow \text{N}_2 + 2\text{H}_2\text{O}$
- Which is the rate-determining step?  $\text{step 1}$
- Identify any reaction intermediates.  $\text{H}_2\text{O}_2$



- Determine Step 1.
- If it was possible to increase the rate of step 1, how would this affect the overall reaction rate?
- Identify any reaction intermediates.

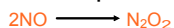


## Kinetics

2) Step 1: rate: fast  
 Step 2:  $\text{N}_2\text{O}_2 + \text{Br}_2 \longrightarrow 2\text{NOBr}$  rate: slow

Overall:  $2\text{NO} + \text{Br}_2 \longrightarrow 2\text{NOBr}$

a) Determine Step 1.

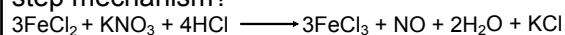


b) If it was possible to increase the rate of step 1, how would this affect the overall reaction rate?

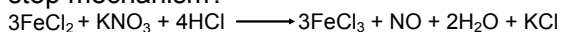
It wouldn't, because step 2 is the rate-determining step.

c) Identify any reaction intermediates.  $\text{N}_2\text{O}_2$

3) Why does the following reaction have a multi-step mechanism?

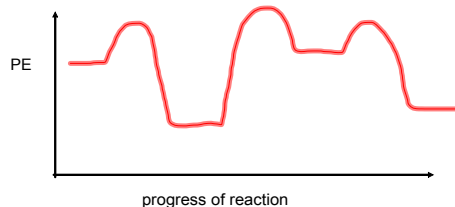


3) Why does the following reaction have a multi-step mechanism?



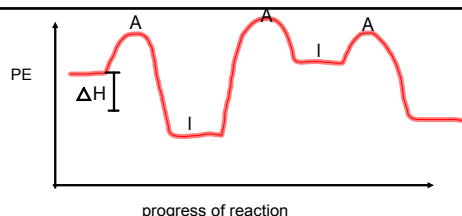
There is an extremely low probability of eight reactant particles colliding simultaneously with correct geometry and sufficient kinetic energy to cause an effective collision.

4)



- How many steps make up the reaction mechanism?
- State whether each step is endothermic or exothermic?
- Is the overall reaction endothermic or exothermic?
- Which is the rate-determining step? How can you tell?
- Label  $\Delta H$  for the overall reaction on the curve.
- On the curve, label an A wherever there is an activated complex and an I wherever there is a reaction intermediate. How do the two differ?

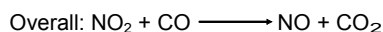
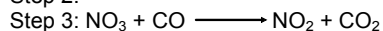
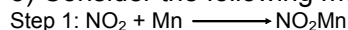
4)



- How many steps make up the reaction mechanism? 3
- State whether each step is endothermic or exothermic?  
 step 1: exo, step 2: endo, step 3: exo
- Is the overall reaction endothermic or exothermic? exo
- Which is the rate-determining step? How can you tell?  
 step 2 as it has the largest  $E_a$
- Label  $\Delta H$  for the overall reaction on the curve.
- On the curve, label an A wherever there is an activated complex and an I wherever there is a reaction intermediate. How do the two differ?

An activated complex is a short-lived, unstable complex, whereas a reaction intermediate is short-lived, but is much more stable.

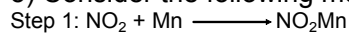
5) Consider the following mechanism:



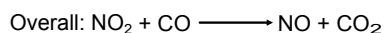
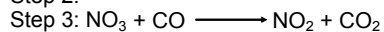
- Determine Step 2.
- Identify a catalyst.
- Identify any reaction intermediates.
- What would be the chemical formula of the activated complex in Step 3?

## Kinetics

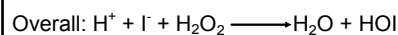
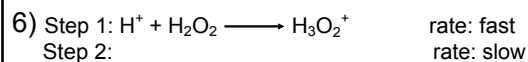
5) Consider the following mechanism:



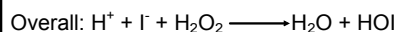
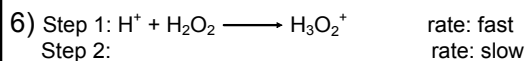
Step 2:



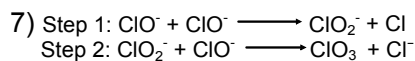
- Determine Step 2.  $\text{NO}_2 + \text{NO}_2\text{Mn} \longrightarrow \text{NO} + \text{Mn} + \text{NO}_3$
- Identify a catalyst.  $\text{Mn}$
- Identify any reaction intermediates.  $\text{NO}_2\text{Mn}, \text{NO}_3$
- What would be the chemical formula of the activated complex in Step 3?  $\text{CNO}_4$  ( $\text{NO}_3$  &  $\text{CO}$  together)



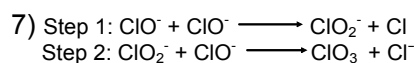
- Determine Step 2.
- Identify any reaction intermediates.
- Which is the rate-determining step?
- What is the formula for the activated complex in Step 2?



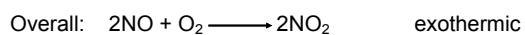
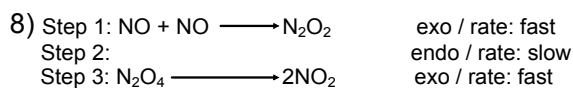
- Determine Step 2.  $\text{H}_3\text{O}_2^+ + \text{I}^- \longrightarrow \text{H}_2\text{O} + \text{HOI}$
- Identify any reaction intermediates.  $\text{H}_3\text{O}_2^+$
- Which is the rate-determining step? 2nd
- What is the formula for the activated complex in Step 2?  $\text{H}_3\text{O}_2\text{I}$



- What is the overall reaction?
- What would be chemical formula be for the activated complex in Step 1?
- Identify any reaction intermediates.



- What is the overall reaction?  
 $3\text{ClO}^- \longrightarrow \text{ClO}_3^- + 2\text{Cl}^-$
- What would be chemical formula be for the activated complex in Step 1?  $\text{Cl}_2\text{O}_2^{2-}$
- Identify any reaction intermediates.  
 $\text{ClO}_2^-$



- Determine Step 2.
- Sketch a PE diagram for the mechanism.

