Reaction Kinetics

A measure of the rate at which a reaction occurs

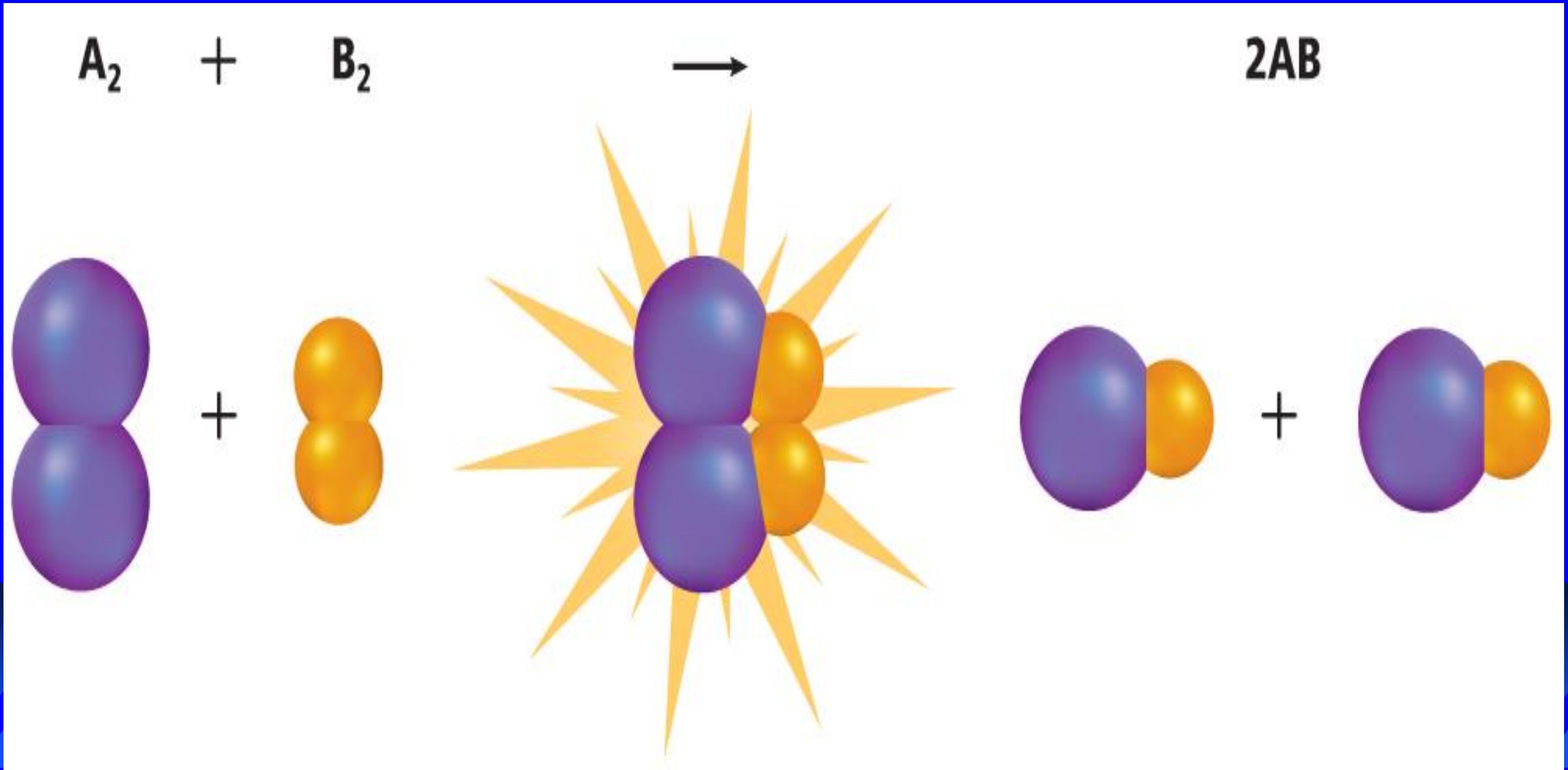
Kinetic Energy

- Related to molecular speed
- KE = $^{1}/_{2}mv^{2}$
  - higher velocity = higher KE

Reaction Rate

- The rate of reaction tells us the **speed** at which a reaction takes place

Collision Theory states that atoms, ions, and molecules must collide in order to react.



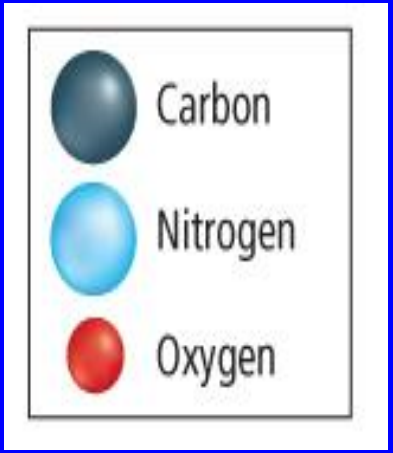
Collision Theory

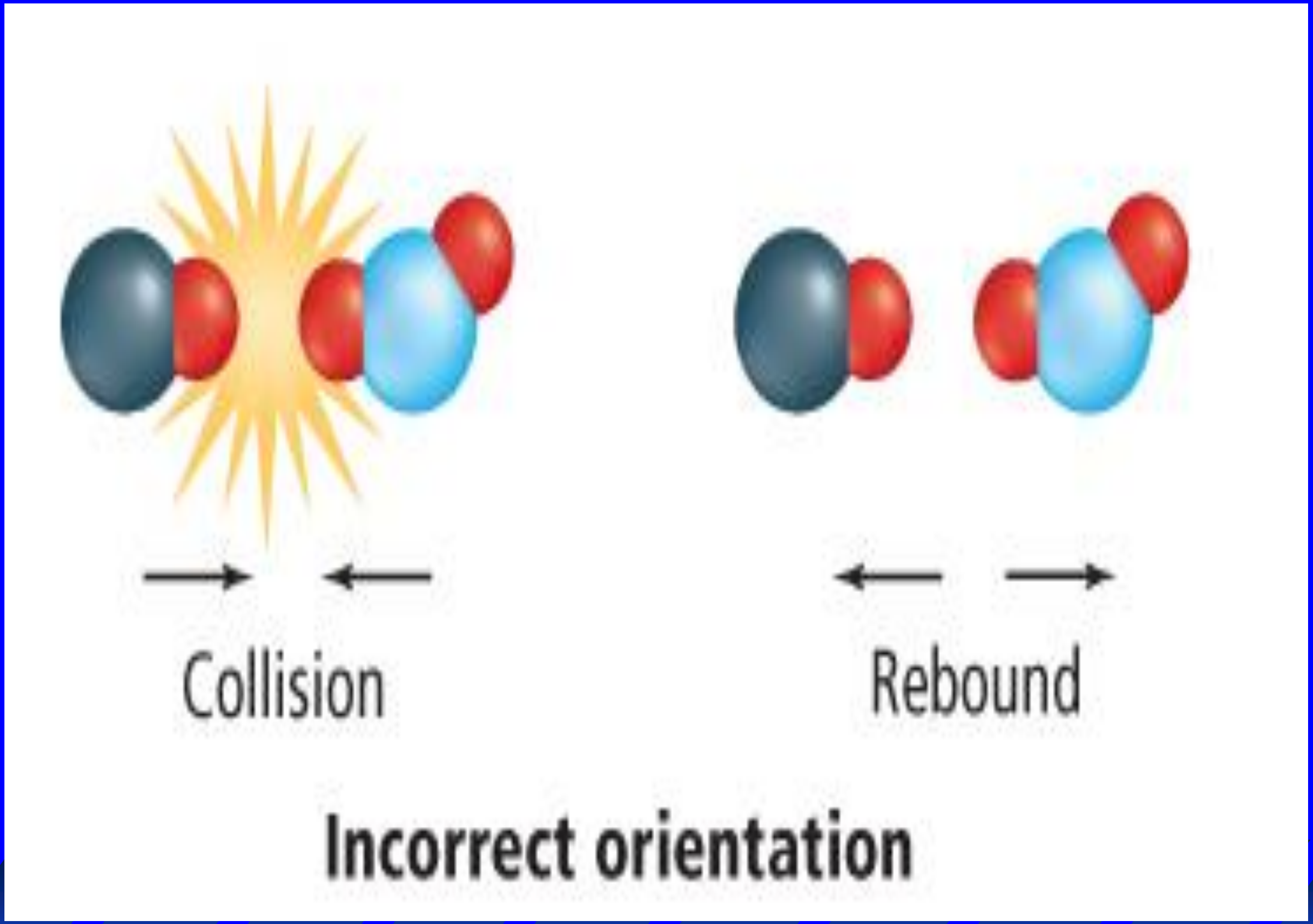
Not every collision between atoms or molecules results in a reaction

1. Orientation is crucial

   ex: darts, train cars

$CO_{(g)} + NO_{2(g)} \rightarrow CO_{2(g)} + NO_{(g)}$





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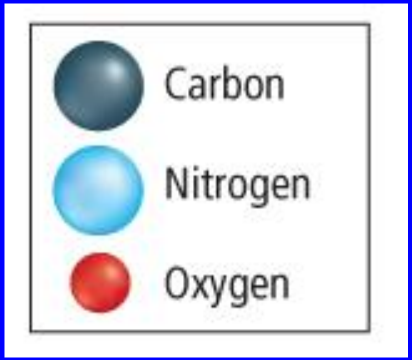

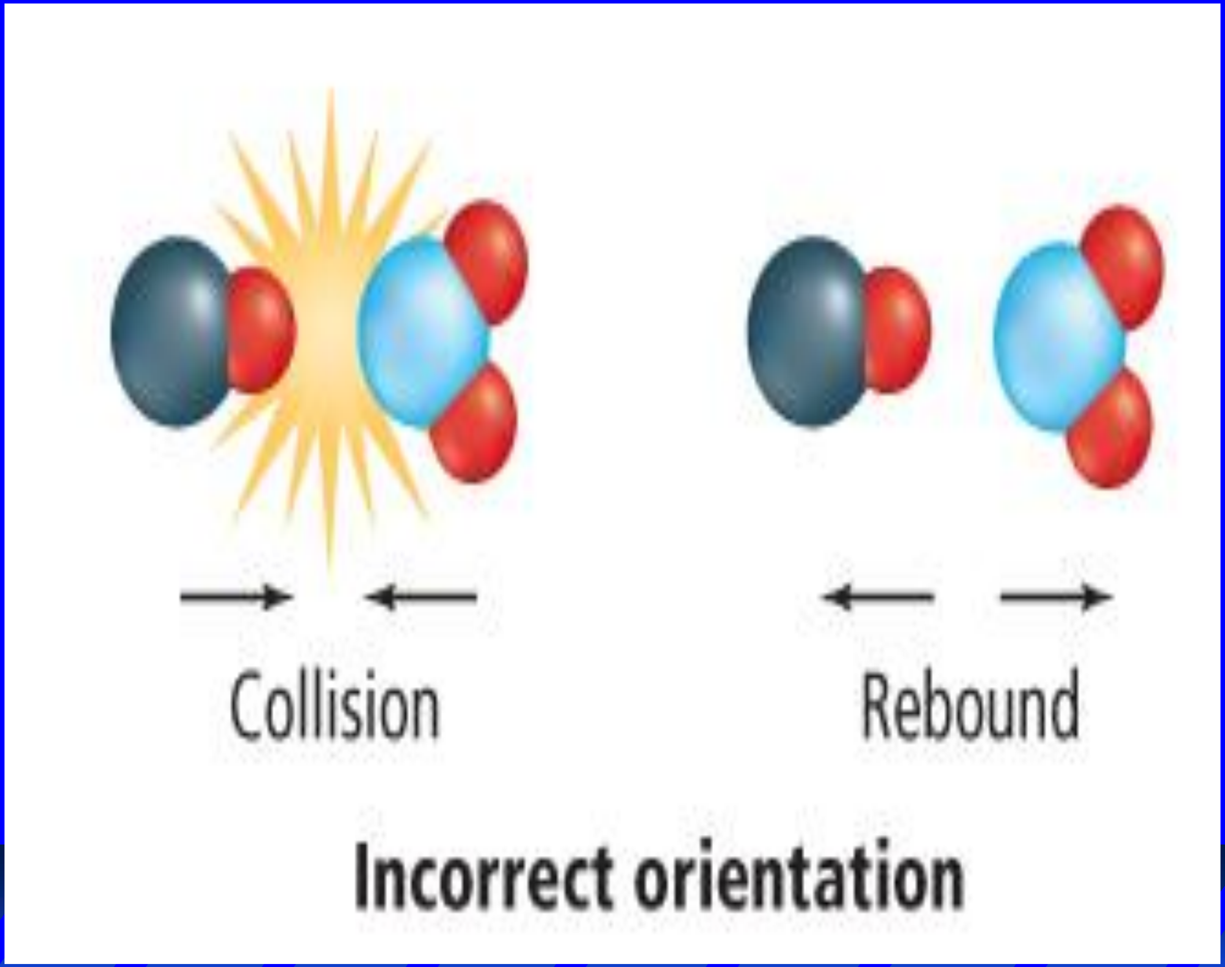

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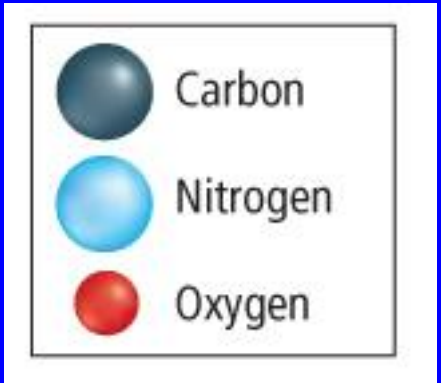

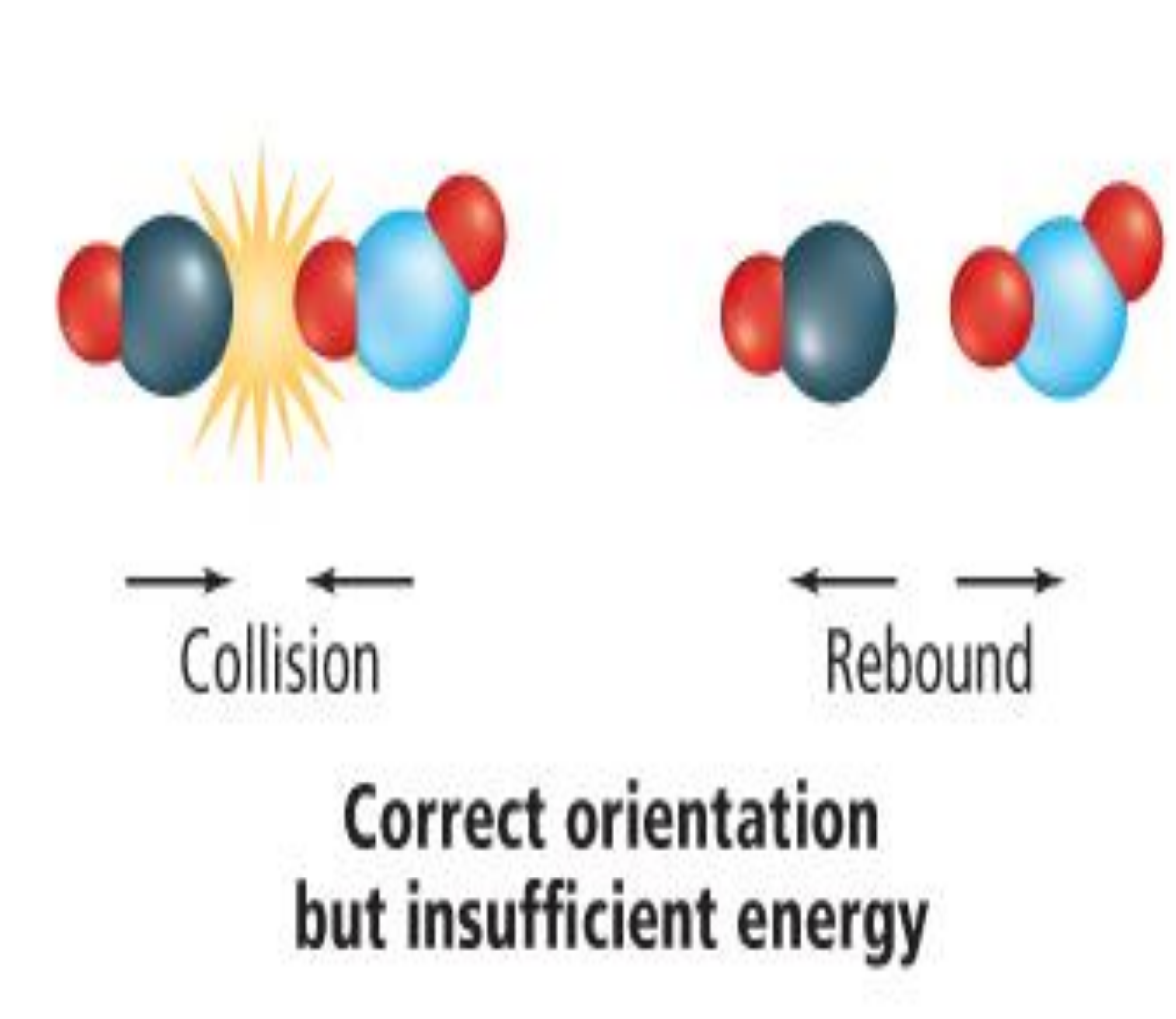

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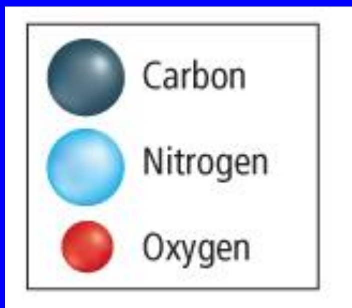
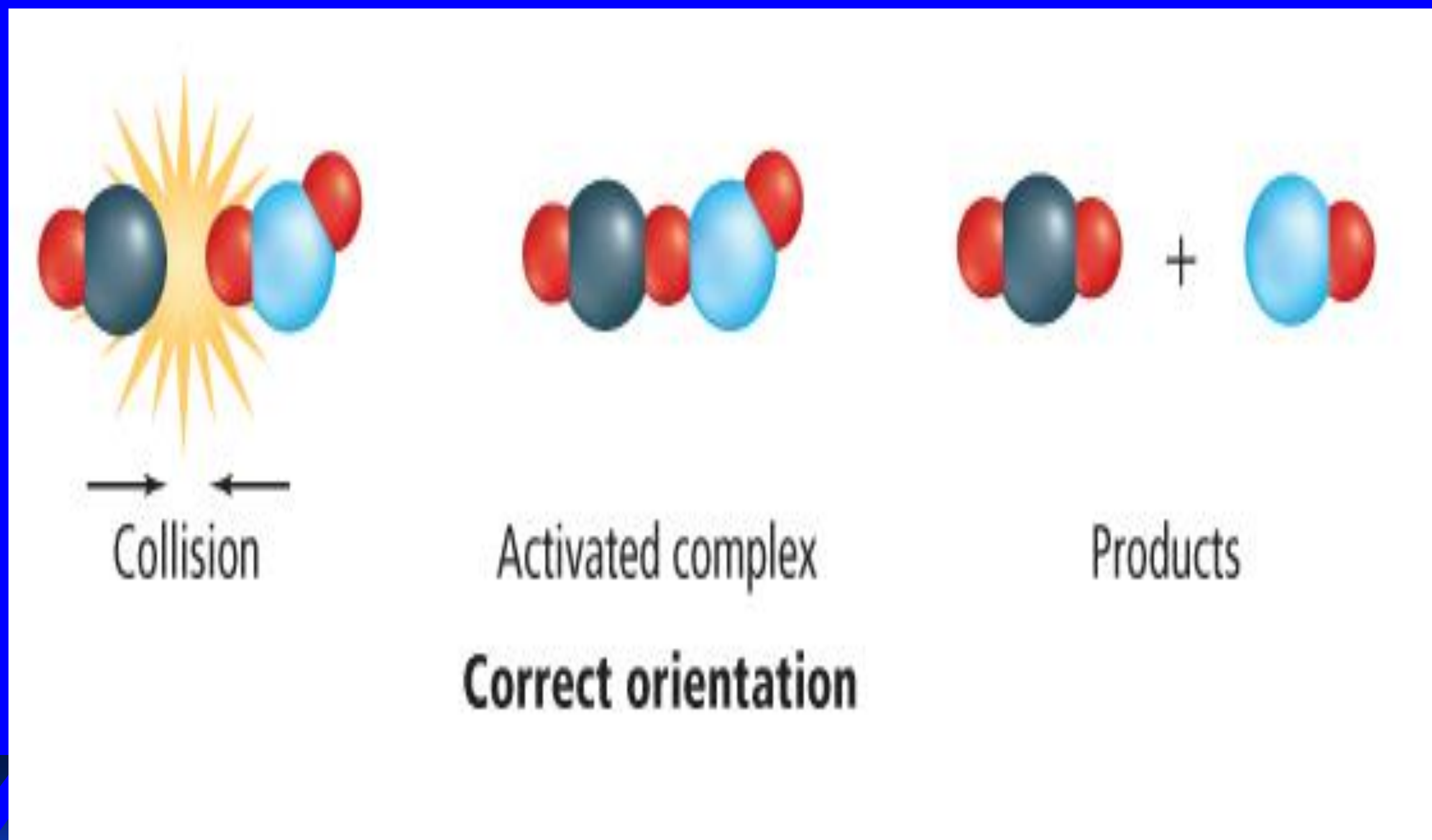




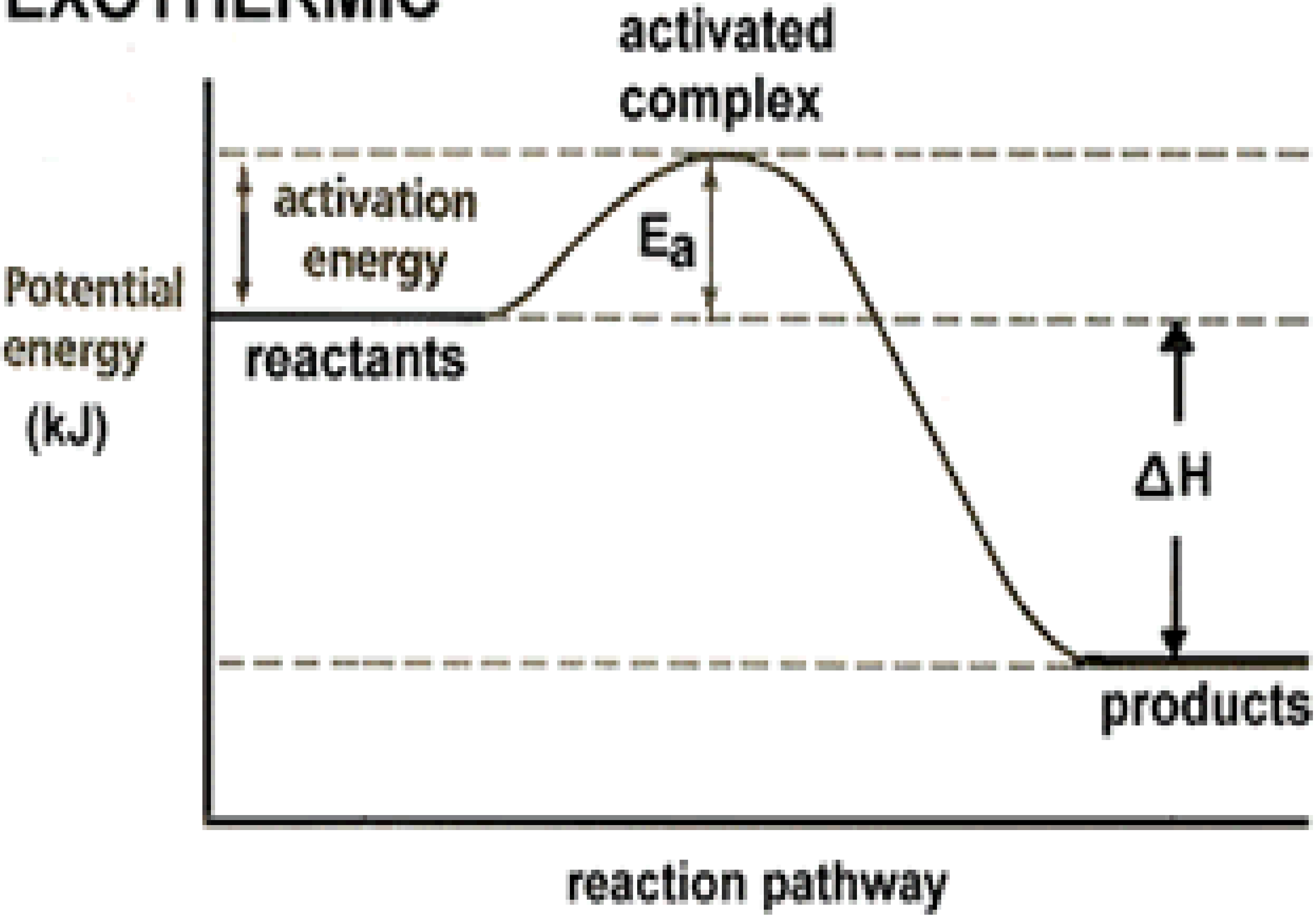
Collision Theory

Not every collision between atoms or molecules results in a reaction

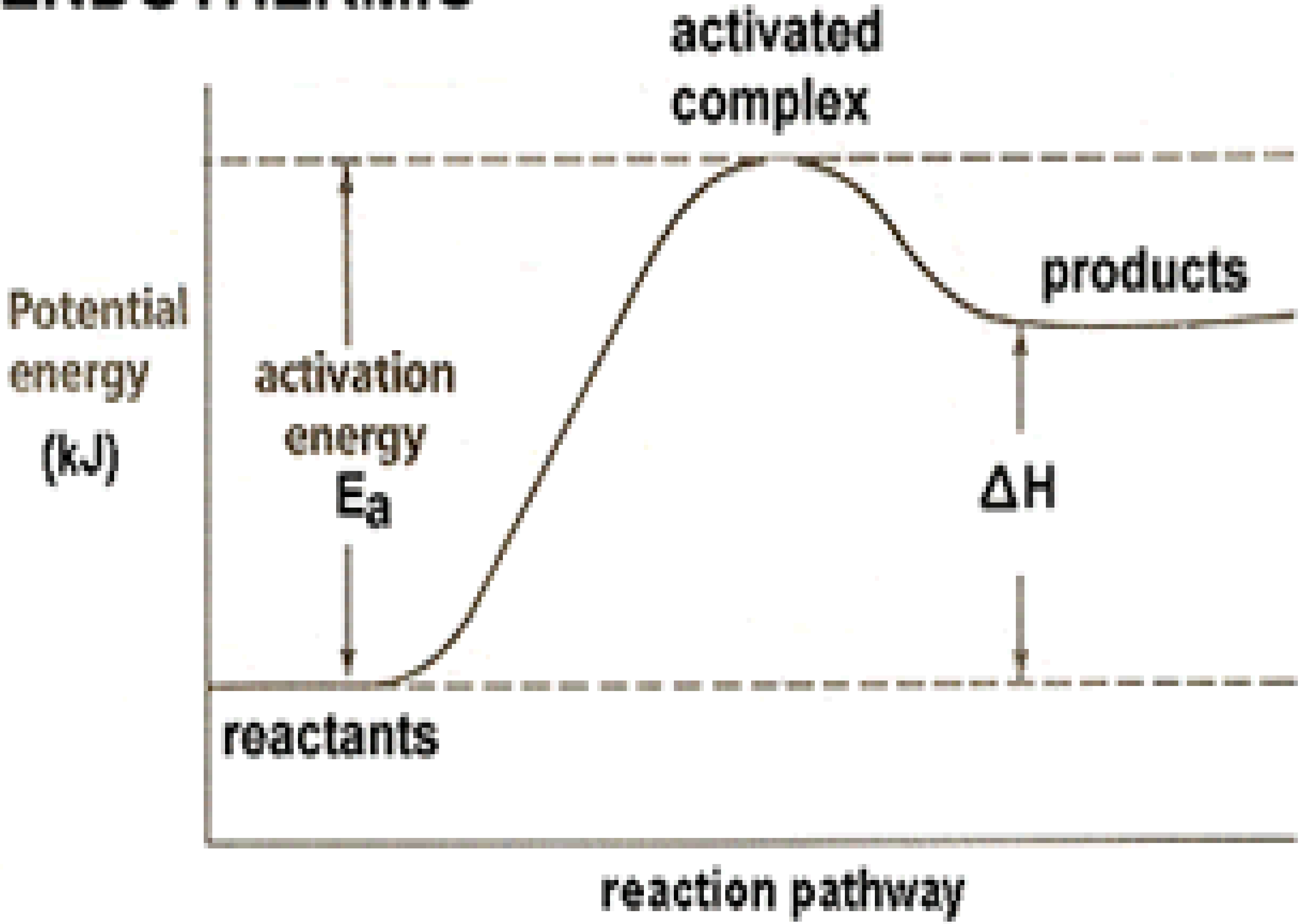
2. They must collide with sufficient energy

   - the "**Activation energy**" (E_a)

EXOTHERMIC


activated
complex
activation
energy
E_a
Potential
energy
(kJ)
reactants
ΔH
products
reaction pathway

ENDOTHERMIC



Collision Theory

Any change that increases the number of collisions should increase reaction rate

Factors in reaction rates…

1. The Nature of the reactants

- what are their "reactivities" / tendencies toward bond formation?
- $Cs + H_2O \rightarrow CsOH + H_2$
  - instantaneous
  - Cs ⇒ very reactive (low IE, EN)
- Fe + H_2O -----> $Fe_2O_3 + H_2$
  - very slow

2. The ability of the reactants to meet (collide)

- reactions usually occur in liquid, gas or aqueous phase
- Related to **surface area** -

  higher surface area means faster reaction in heterogeneous reactions

  - **reaction only occurs at phase interface**
  - **ex: log vs firewood vs kindling vs sawdust**

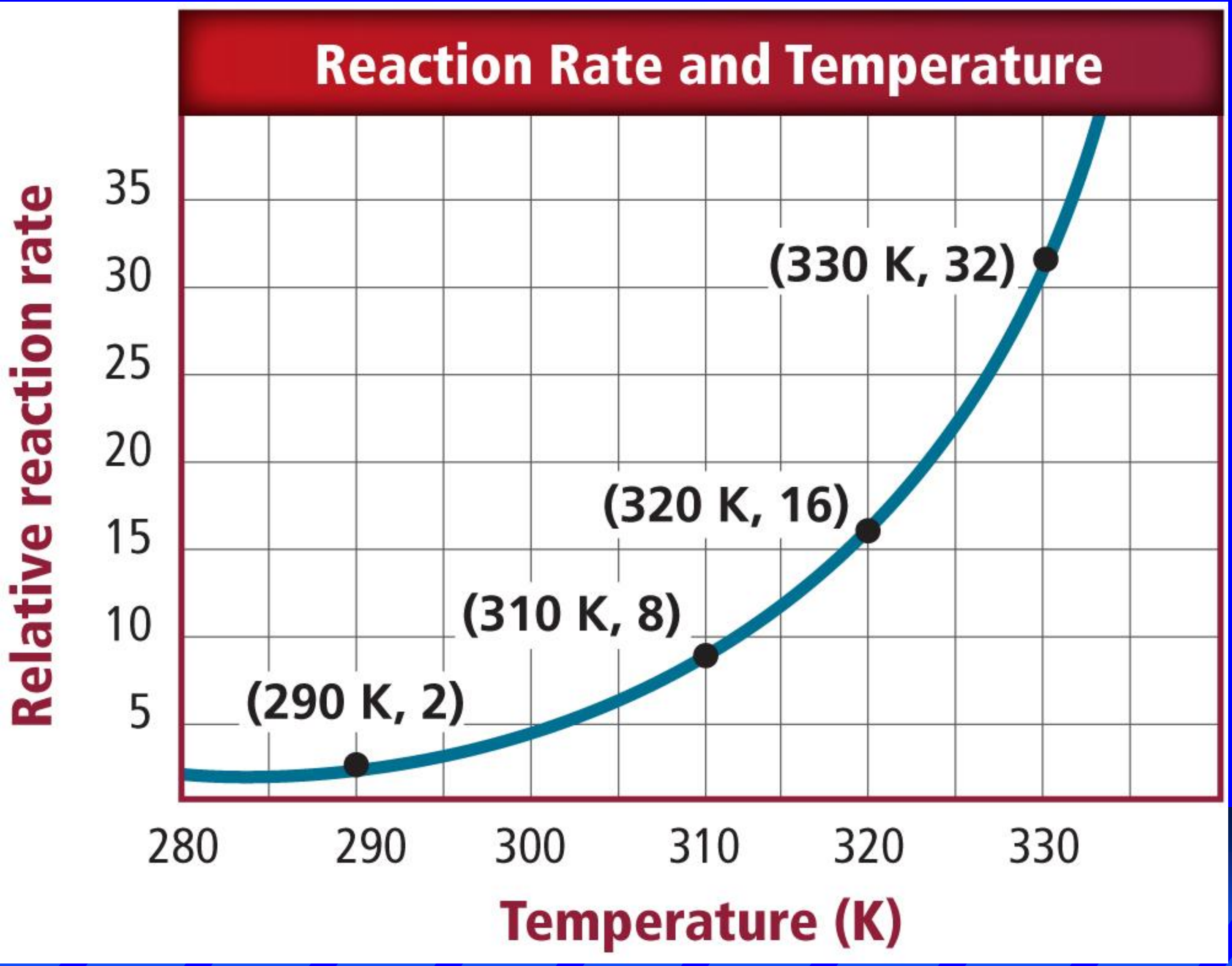
3. Concentration of the reactants

- often listed as molarity (mol/L)
- written as [square brackets]
  - ex: [HCl] = “the M of the HCl”
- more reactants = more collisions
- more collisions = increased reaction rate

4. Temperature

- T is a measure of the average KE
- higher T means higher KE
- higher KE = molecules moving faster
- faster moving molecules means more collisions
- more collisions = faster rate

4. Temperature



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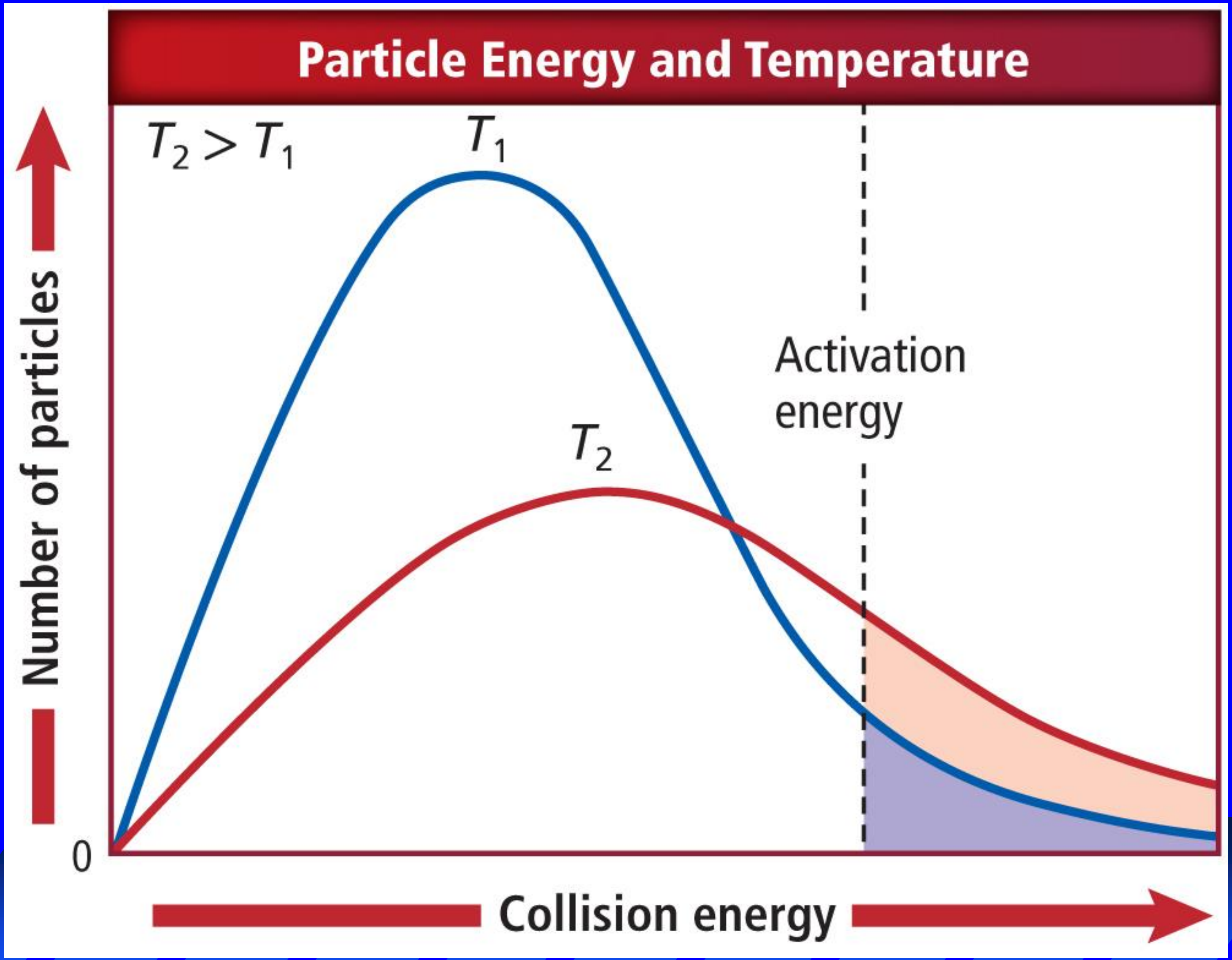
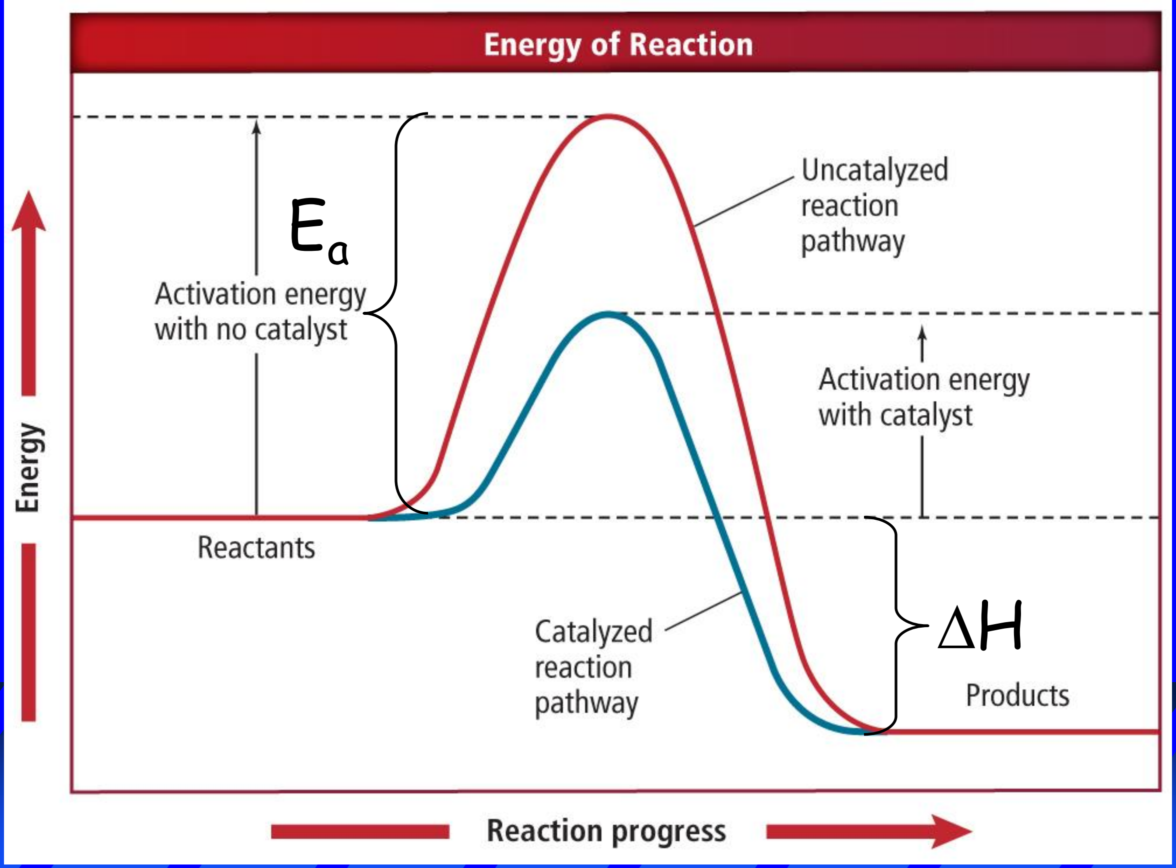


5. Catalysts

- increase the reaction rate without being consumed in the process
- lower the energy barrier needed to overcome in order to react (E_a)

5. Catalysts

- Lower activation energy means more collisions between particles have sufficient energy to react.



5. Catalysts

- A **heterogeneous catalyst** exists in a physical state different than that of the reaction it catalyzes.
- A **homogeneous catalyst** exists in the same physical state as the reaction it catalyzes.
- A **reforming catalyst** is consumed in the reaction, but then re-made
- An **enzyme** is a biologically active catalyst
- An **inhibitor** slows down a reaction by increasing the activation energy

Next:
a look inside the numbers....

Reaction Kinetics

A measure of the rate at which a reaction occurs

To Review....Five factors that change reaction rate

1. **The nature of the reactants**
2. **the ability of the reactants to collide**
3. **concentration of reactants**
4. **temperature of the system**
5. **catalysts**

Measuring Rates

- A rate is something per unit of time
  - interest rate = $ earned / time
  - speed = distance traveled/ time
  - pay = dollars / hour

Measuring Reaction Rates

- as the reaction proceeds, the reactants are "used up"
  - [reactants] goes down
  - [products] goes up
- Reaction rate is measured as : Δ concentration ÷ Δ time

◆ $rate\ unit: \frac{\Delta concentration}{\Delta time} = \frac{M}{s} = \frac{mol/_{L}}{s} = Ms^{-1} = mol\ L^{-1}s^{-1}$

Coefficients and rate

- The coefficients of a chemical equation give us some insight as to the relative rates of consumption of reactants and production of products, but not "absolute" amounts.

Imagine a machine where you put in 1 dime, 2 nickels, and 5 pennies in a slot on the top, and a quarter drops out the other side

$$1\ D + 2\ N + 5\ P = 1\ Q$$

If you had a pile of 100 dimes, a pile of 100 nickels, and a pile of 100 pennies, are the piles being used up at the same rate?

$1\ D + 2\ N + 5\ P = 1\ Q$

How does the "rate of disappearance" of pennies compare to the "rate of appearance" of quarters?

If the nickels are disappearing at a rate of 3 per second, how fast are the pennies disappearing?

$3\ N\ s^{-1} \times 5\ P/2\ N = 7.5\ P\ s^{-1}$

consider the reaction

$$2\,A + B \rightarrow 3\,C + 2\,D$$

- it takes 2 A's for every single B that reacts, so A is "disappearing" twice as fast as B is

consider the reaction

$$2\,A + B \rightarrow 3\,C + 2\,D$$

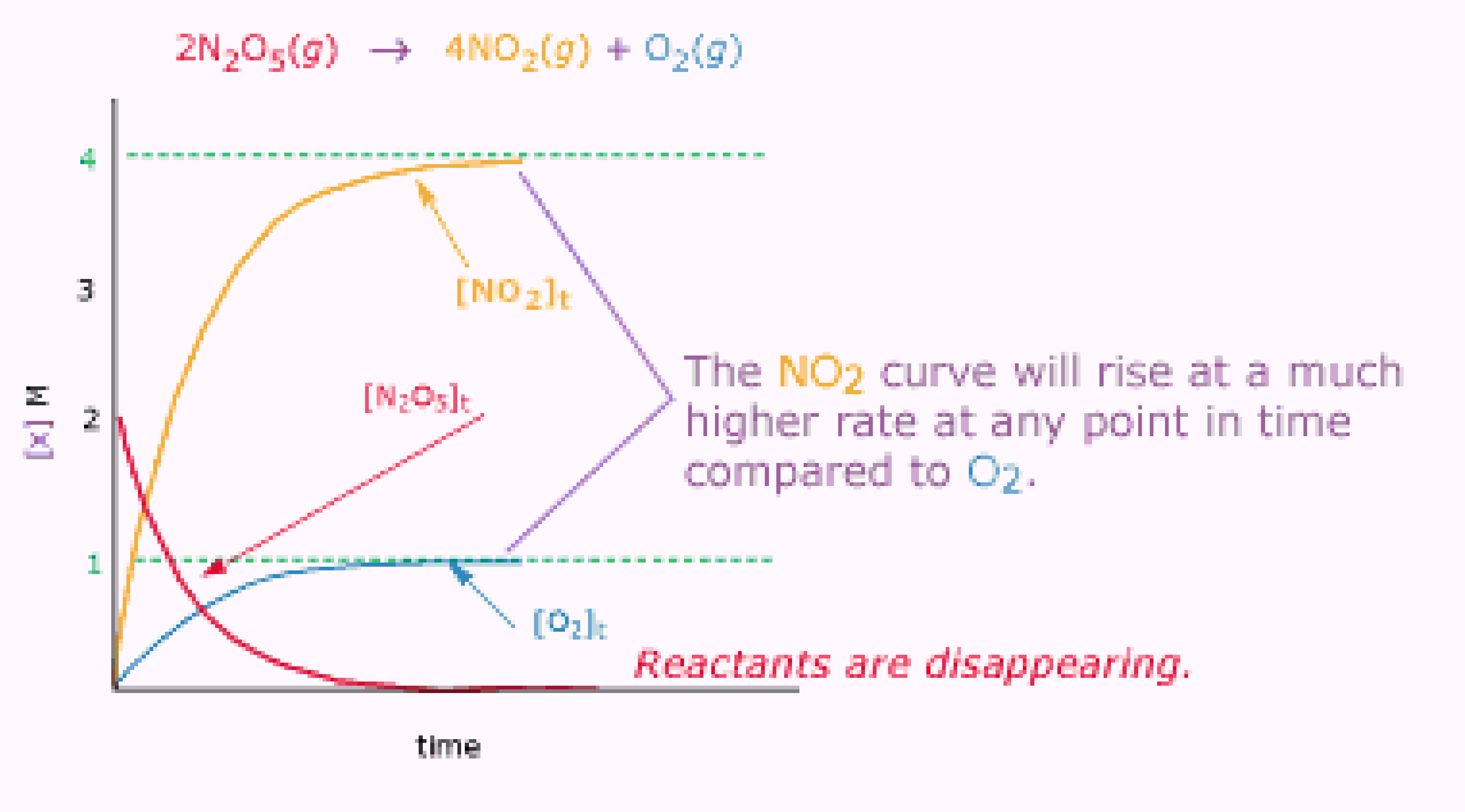
- there are 2 D’s produced for every B and every 2 A’s that are “consumed”, so D is appearing at the same rate A is disappearing and twice as fast as B is disappearing

consider the reaction

$2\,A + B \rightarrow 3\,C + 2\,D$

- C is being produced 3 times faster than B is being used up, and 3/2 times faster that A is being used up

Graphing the rate of ΔM for $2N_2O_{5(g)} \rightarrow 4NO_{2(g)} + O_{2(g)}$



http://www.chem.latech.edu/~upali/chem102/groupactivity/Ch13a.pdf

A **<u>Rate Law</u>** is an equation that describes how a <u>change in concentration</u> affects the reaction rate.

- for the reaction: **A + B → products**

- The rate law would be:

$$\textbf{rate} = k[A]^m[B]^n$$

$rate = k[A]^m[B]^n$

- **k** = the <u>*rate constant*</u> → depends on the **temperature**, different for each reaction
- **m =** the "*order of reaction*" with respect to A
- **n** = the "*order of reaction*" with respect to B

$rate = k[A]^m[B]^n$

- **m** and **n** have to be experimentally determined; they are not the same as the reaction coefficients except in a “one step mechanism” (pretty rare).

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