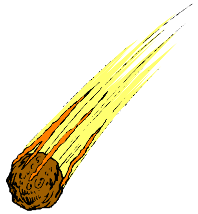
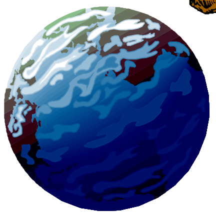
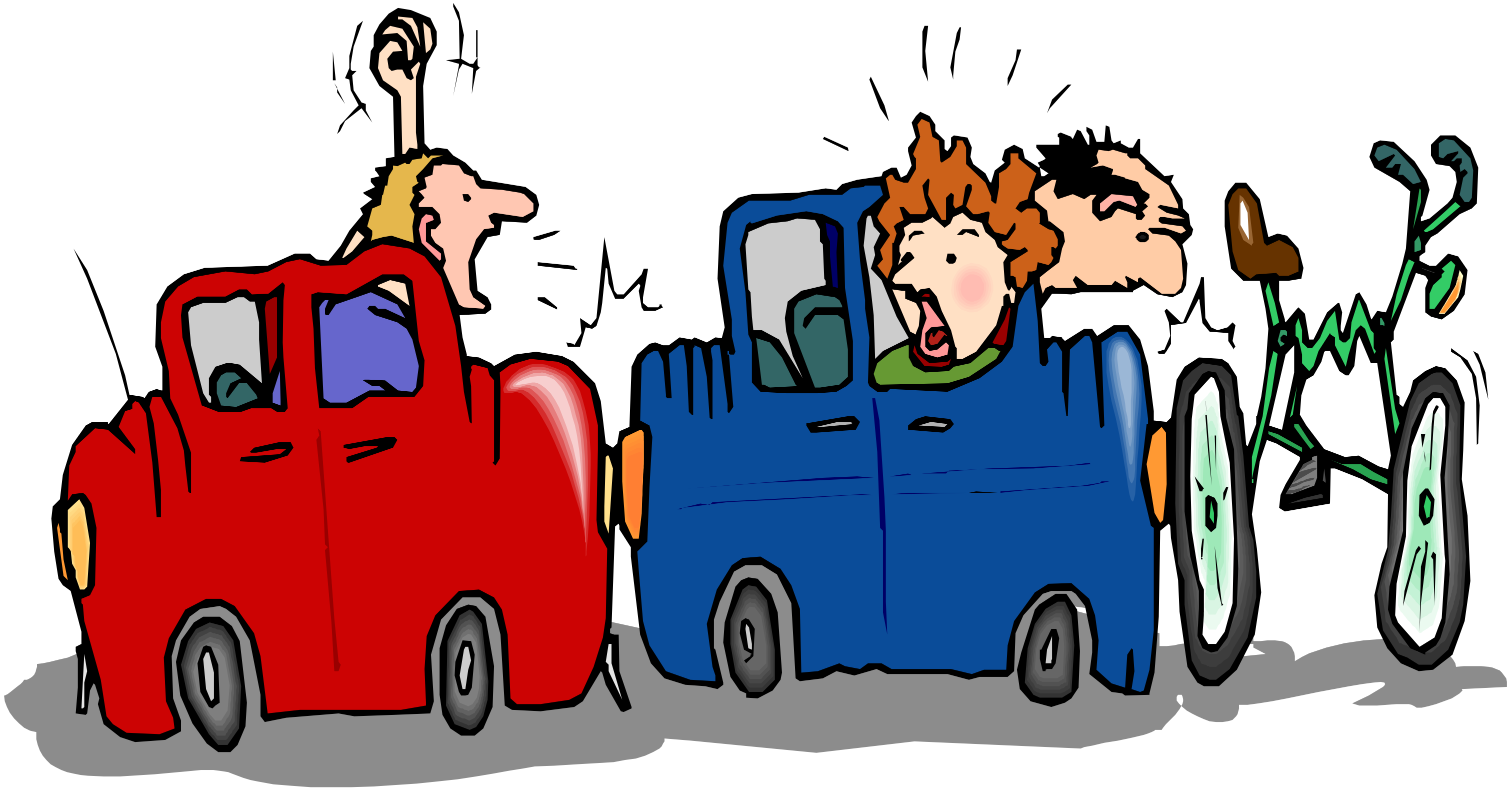
The Collision Theory and Rates of Reactions





Explaining how and why factors affect reaction rates



Elephant toothpaste

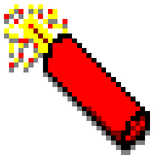
- We are going to look at a reaction named after elephant toothpaste and you’ll see why
- In graduated cylinder # 1 – 20 mL 8% H_2O_2, 1 mL of detergent and a 0.5g of KI
- In graduated cylinder #2 - 20 mL 30% H_2O_2, 1 mL of detergent and a tip of of KI

A Look at The reaction

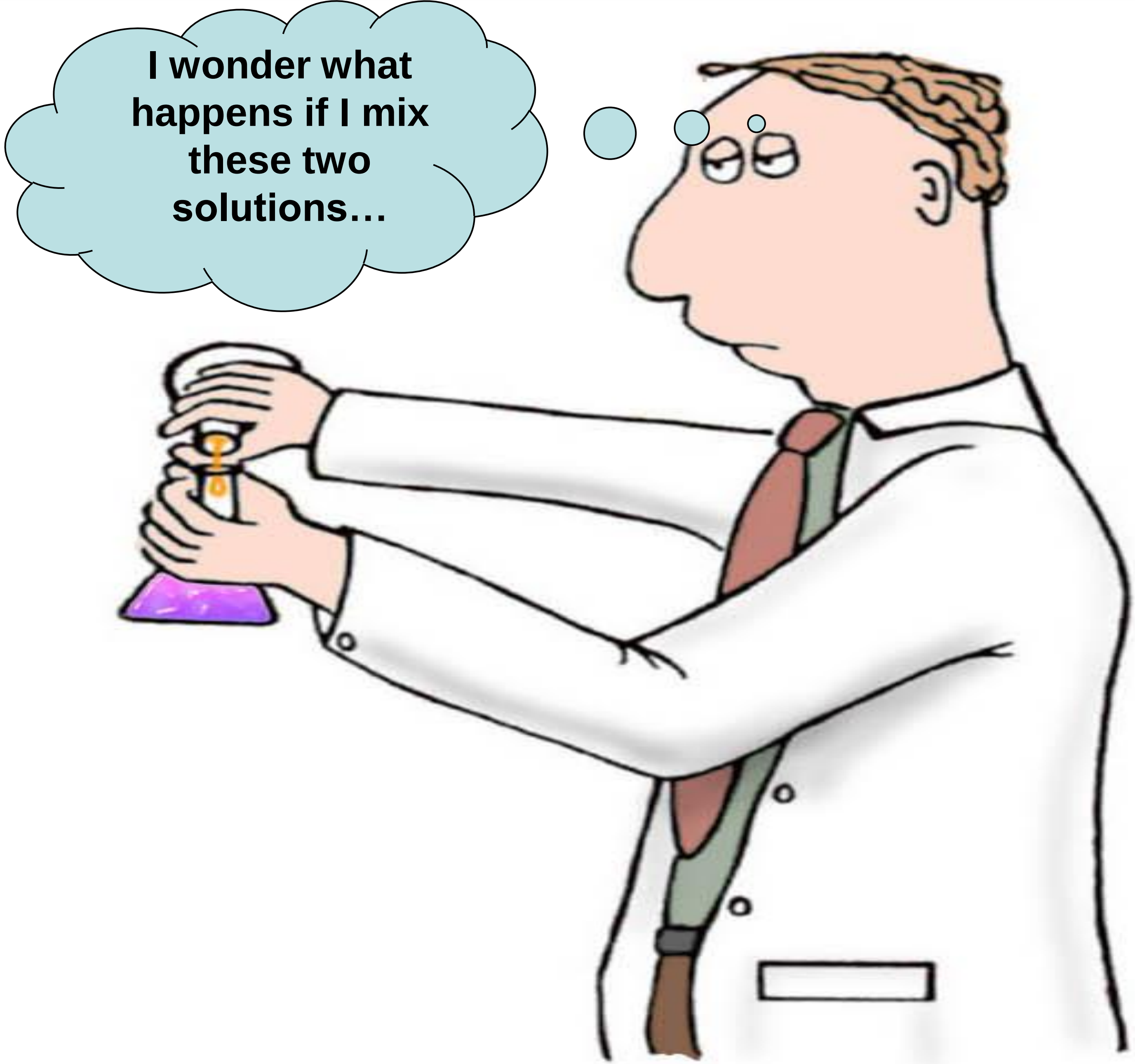
- What we observed was the decomposition of hydrogen peroxide using potassium iodide as a catalyst (the soap is just for fun)
- $2H_2O_{2(l)} \rightarrow 2H_2O_{(l)} + O_{2\,(g)}$
- As you saw the rate of reaction can change quite drastically!

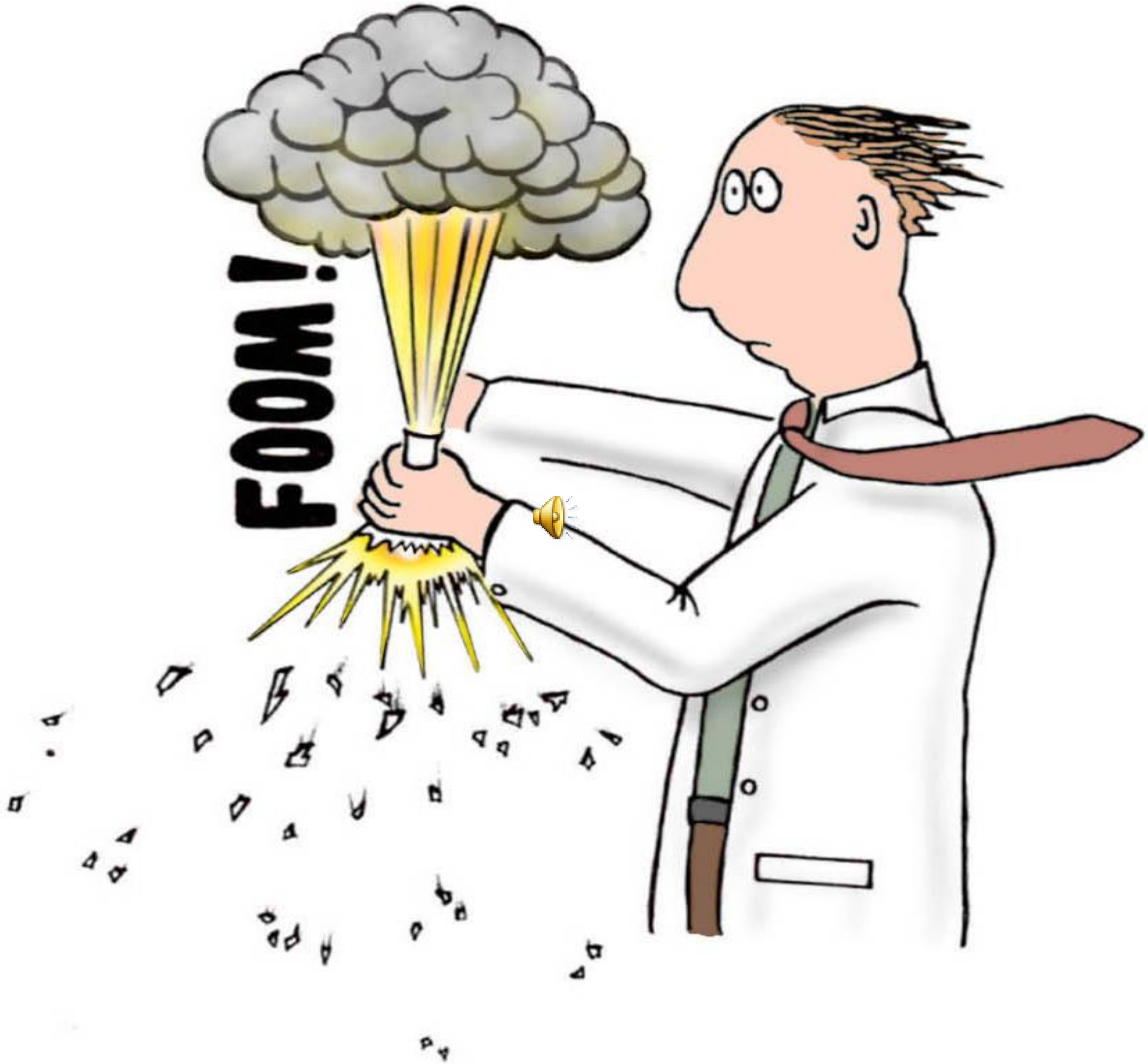
Rates of Reactions

- Reactions can occur:
  - *<u>Very fast</u>* – such as a firecracker
  - *<u>Very slow</u>* – such as the time it took for dead plants to make coal
  - *<u>Moderately fast</u>* – such as food spoilage



- A "<u>rate</u>" is a measure of the speed of any change that occurs within an interval of time
- In chemistry, reaction rate is expressed as the <u>change in amount of reactant (or product) per unit time</u>.
- For example: 3 moles/year or 5 grams/second


I wonder what happens if I mix these two solutions…


FOOM!


WOW, that was really FAST


It was also really FUN


I wonder if I should be wearing my goggles?

What might be some reasons why chemists might want to know about reaction rates?

Designing new medications

Improving techniques to control pollution

Develop more efficient methods in the food processing industry

How do reactions occur?

Collision Theory

Collision Theory

In order for a chemical reaction to take place, the reactants must collide.

The collision transfers kinetic energy needed to break the necessary bonds so that new bonds can be formed.

Collision requirements

Requirement 1

Must have the proper orientation.

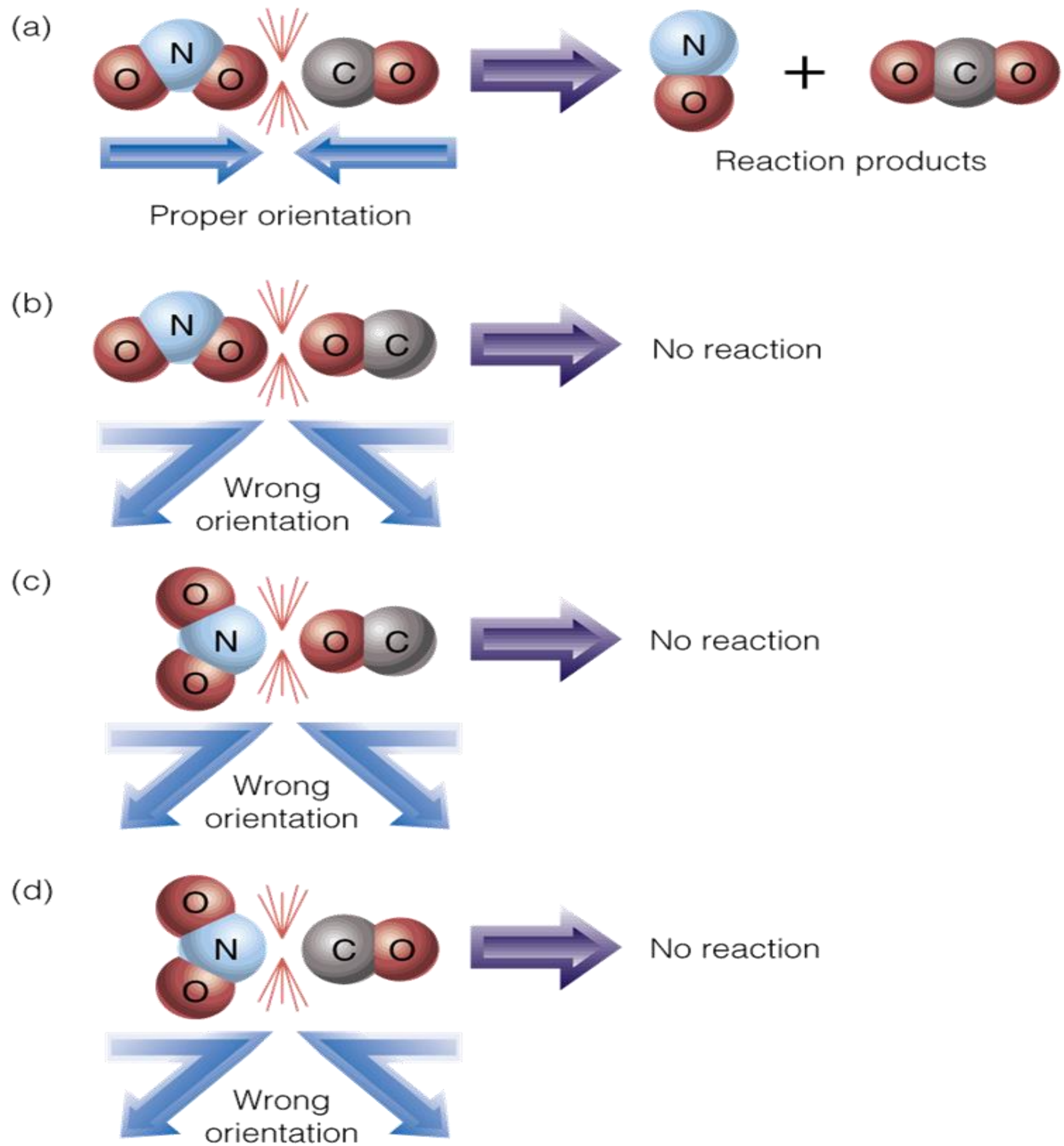
$2HCl + Mg \rightarrow MgCl_2 + H_2$

Mg H---Cl

Wrong Orientation

H-Cl Mg

Correct Orientation



(a)
N
O
O
C
O
N
O
+
O
C
O
Proper orientation
Reaction products
(b)
N
O
O
O
C
No reaction
Wrong orientation
(c)
O
N
O
O
C
No reaction
Wrong orientation
(d)
O
N
O
C
O
No reaction
Wrong orientation

Collision requirements

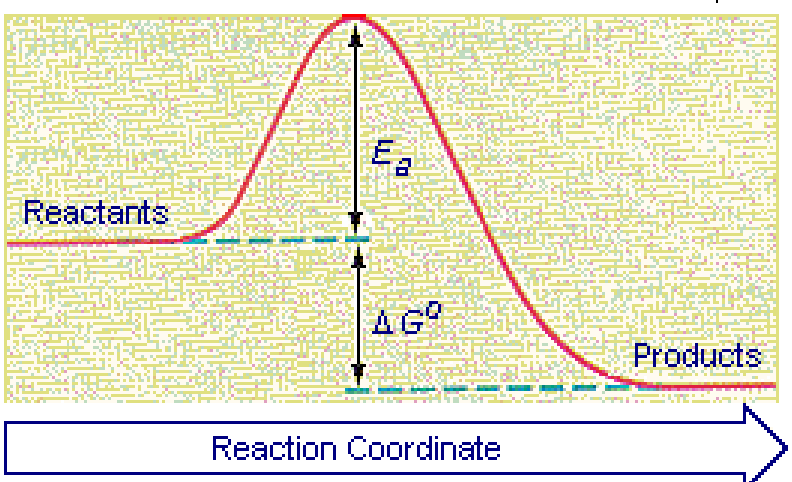
<u>Requirement 2</u>

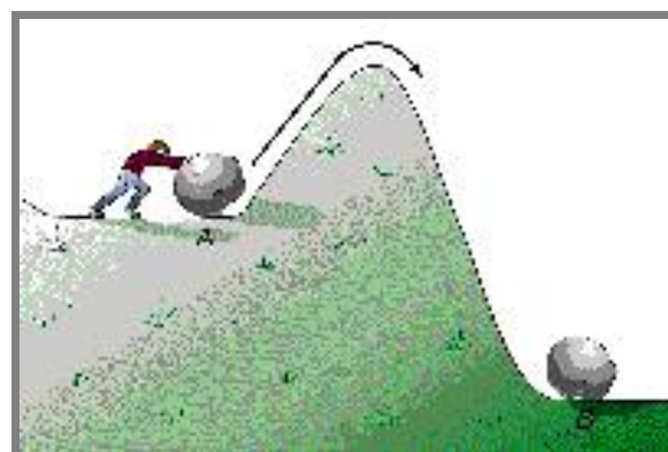
Must have enough kinetic energy to reach a threshold of energy called <u>activation energy</u>

H---Cl Mg

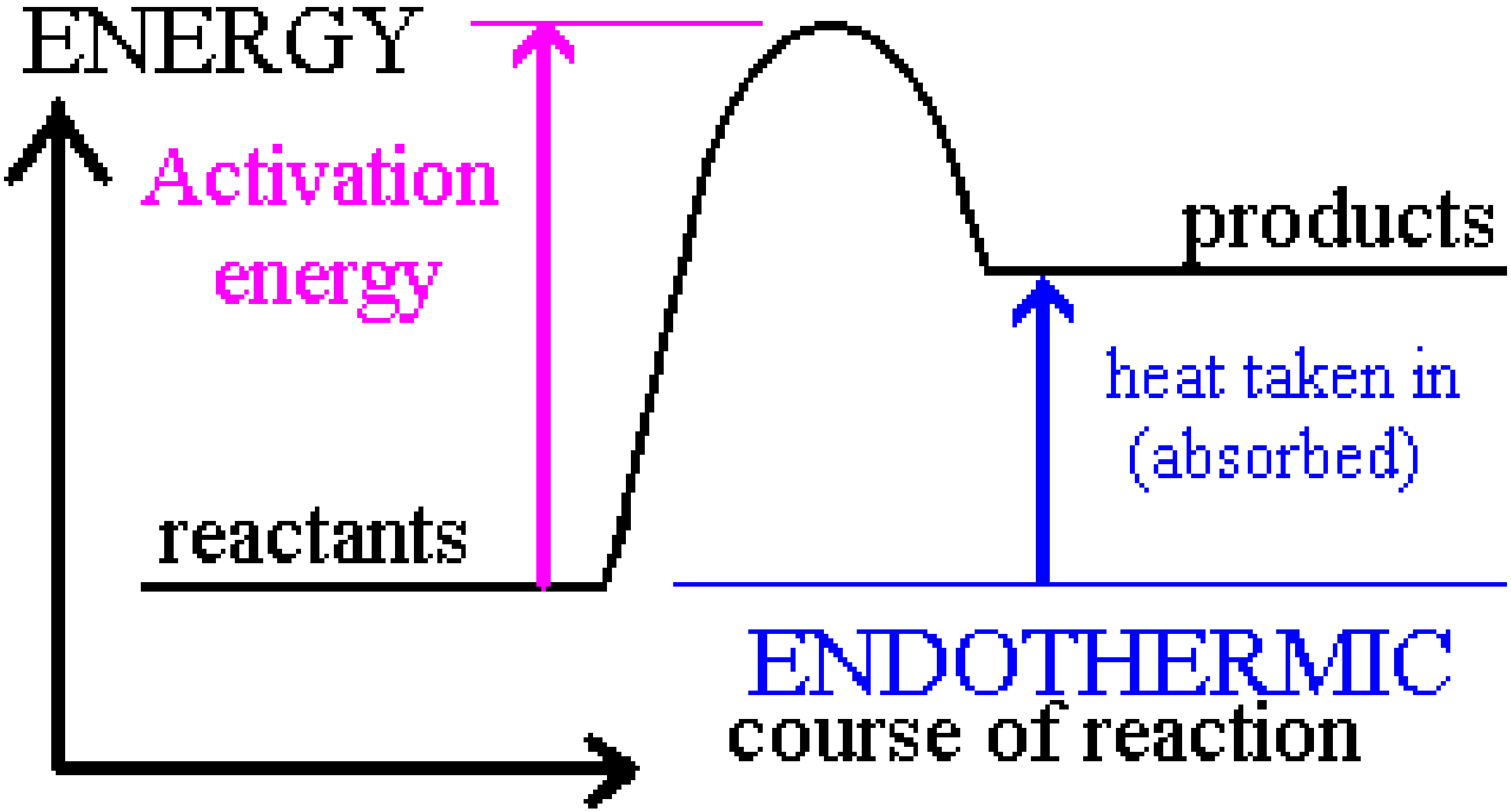
H---Cl Mg

H Cl-Mg

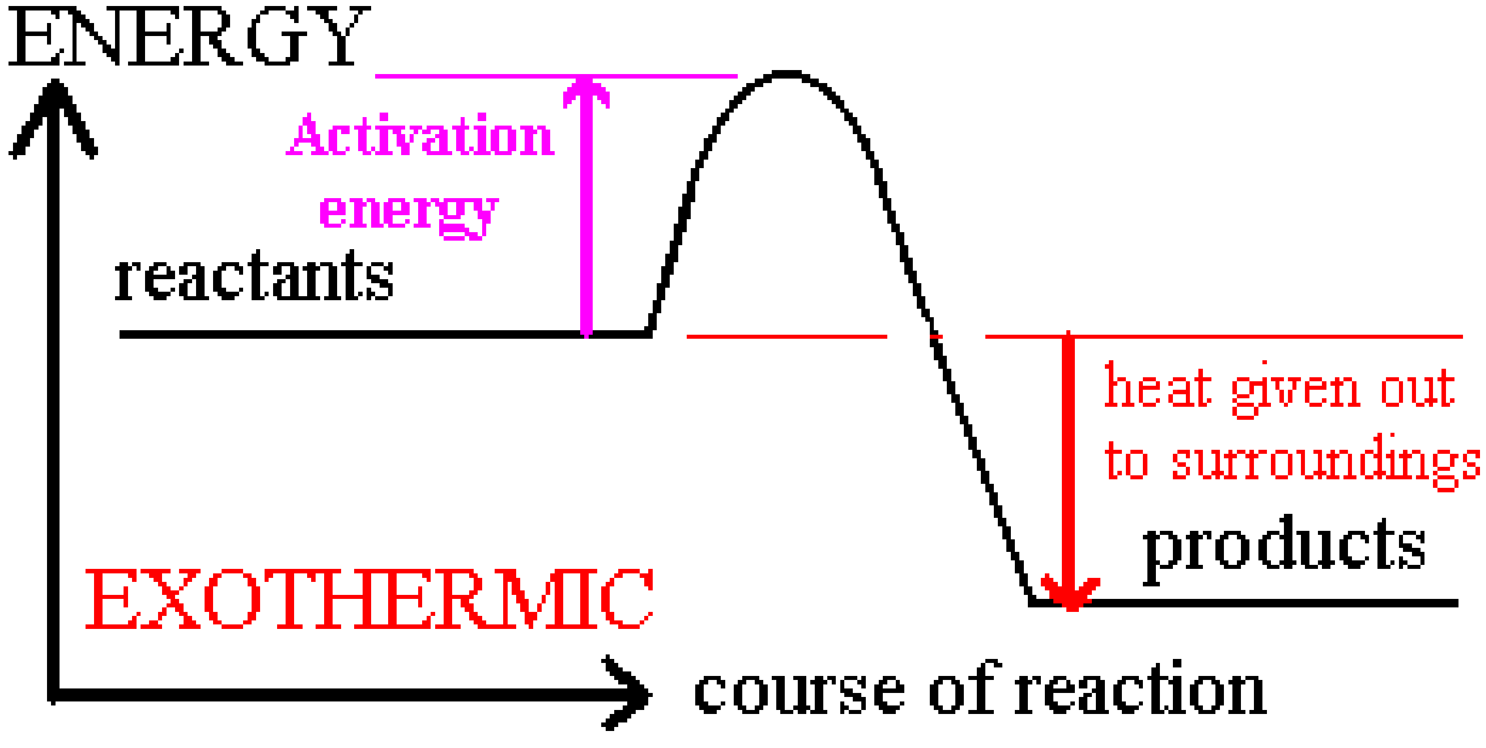




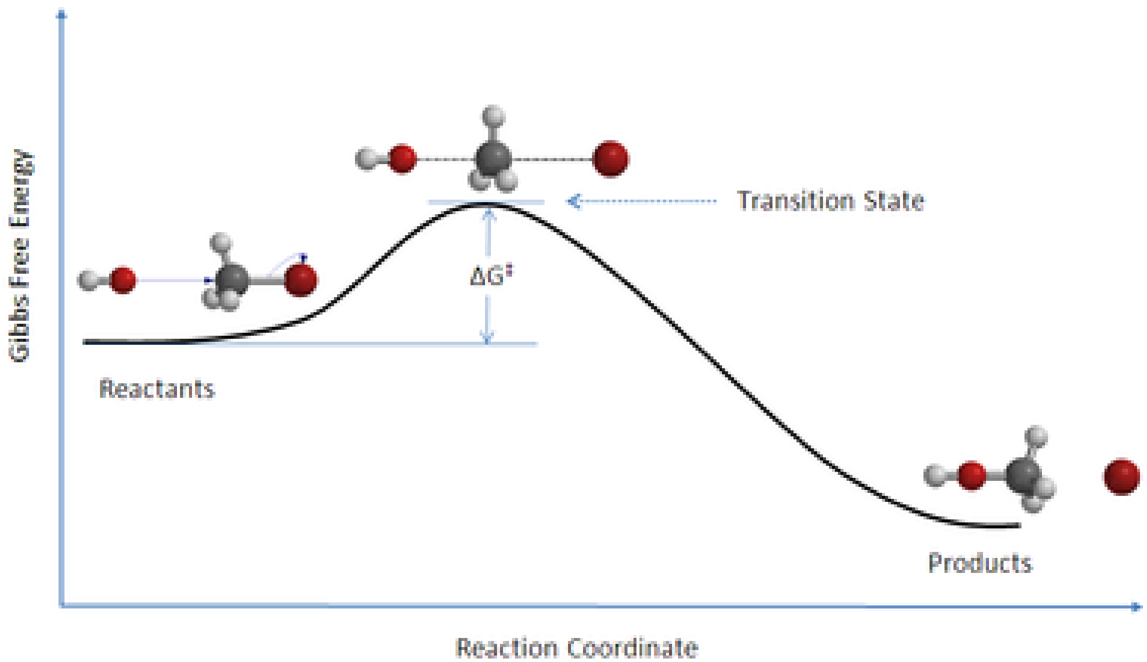
Energy of Activation



Energy of Activation



- An "activated complex" is an *unstable* arrangement of atoms that forms momentarily with partial bonds, at the peak of the activation energy barrier.
- This is sometimes called the ***transition state***



Reaction: $HO^- + CH_3Br \rightarrow [HO\text{---}CH_3\text{---}Br]^\ddagger \rightarrow CH_3OH + Br^-$

ACTIVATED COMPLEX

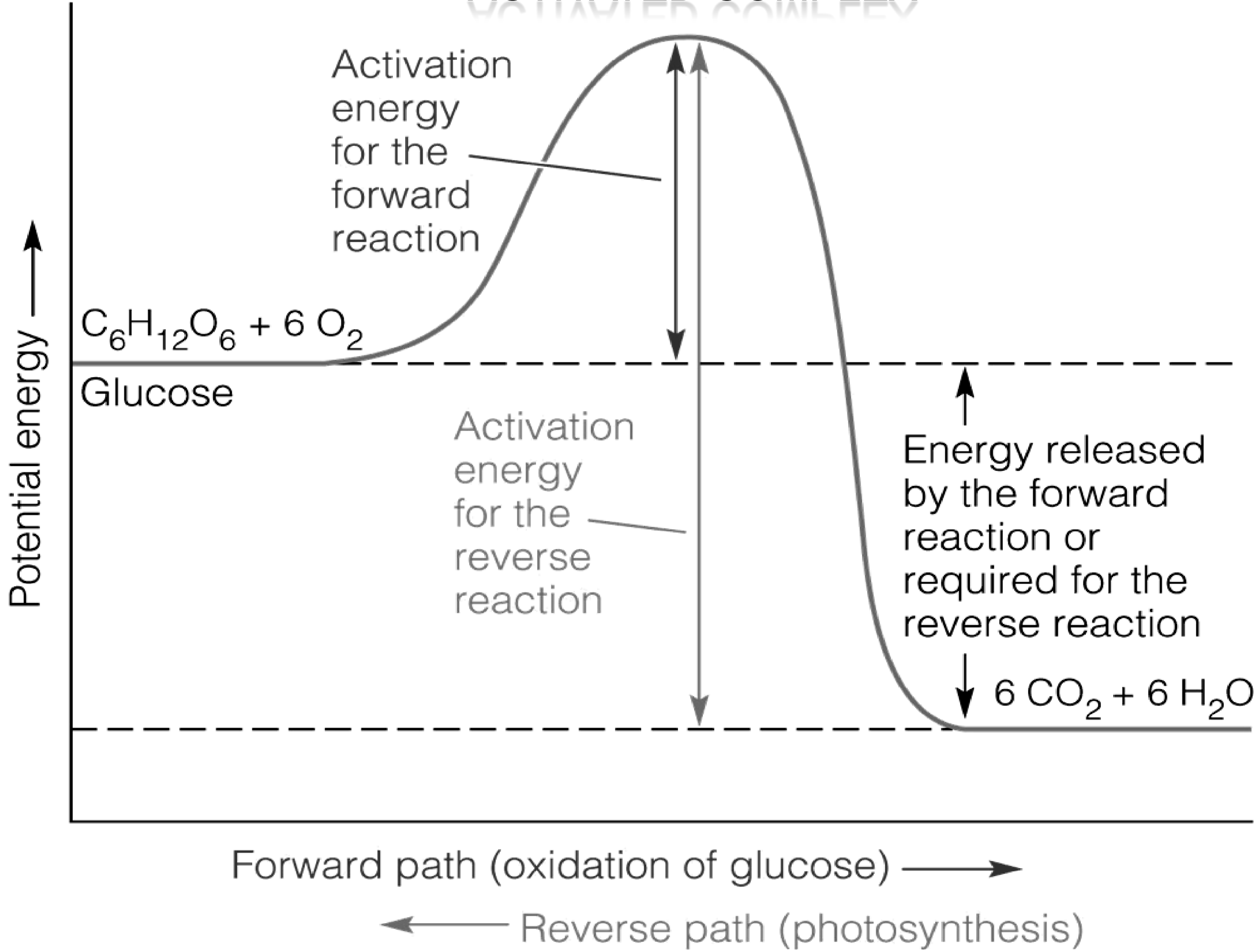
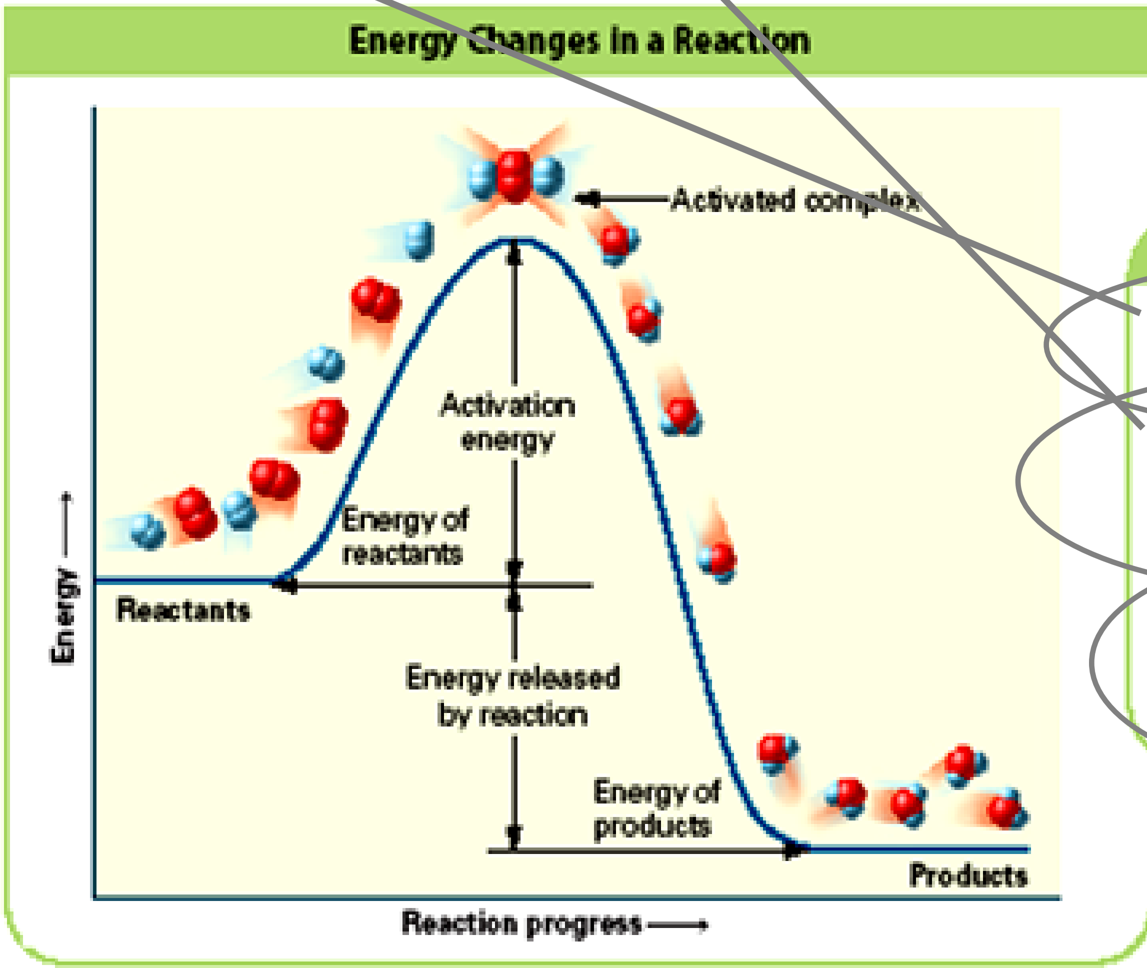


Figure 18.5 Interpreting Graphs

a. Reactants

b. Absorbed

c. No; it could also revert back to the reactants



INTERPRETING GRAPHS

a. Navigate Which are at a higher energy, the reactants or products?

b. Read Is energy absorbed or released in progressing from the reactants to the activated complex?

c. Interpret Once the activated complex is formed, will it always proceed to form products? Explain.

Increasing the Rate of Reactions

What needs to happen in order for the rate of the chemical reaction to increase?

More collisions = Faster reaction rate

5 Factors that affect the rate of a reaction:

- Temperature
- Concentration
- Surface area/particle size
- Chemical nature of reactants
- Catalyst

1. Temperature

- Temperature is a measure of the average kinetic energy of the particles
- The higher the temperature, the faster the particles are moving; which will increase the chance for collisions, as well as the force at which collisions occur

Examples:

Eggs cook faster at higher temperatures.

Cookies bake faster at higher temperatures.

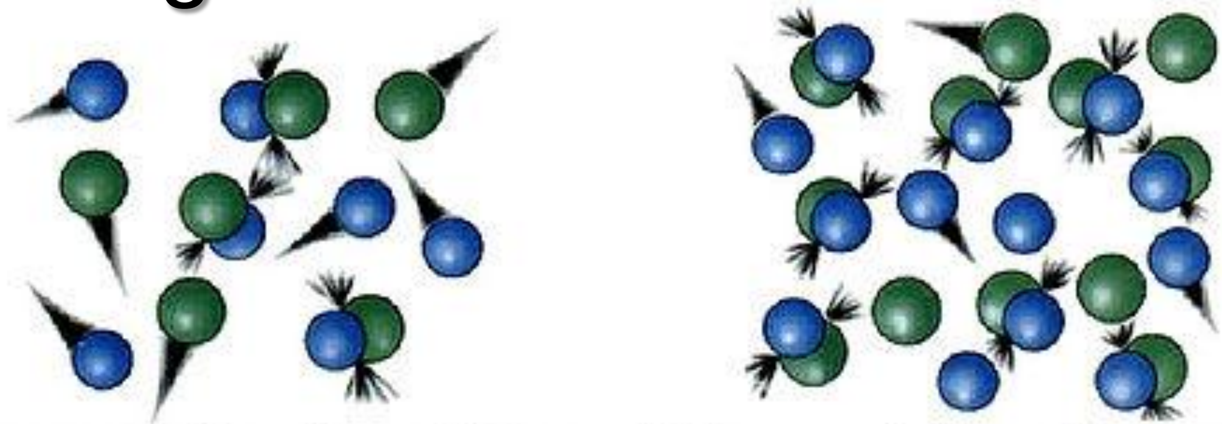
Bread dough rises more quickly in a warm place than in a cool one.

Low body temperatures slow down metabolism. In fact, warm-blooded animals regulate body temperature so that their biochemical reactions run at the correct rate.

Lightsticks produce light via a chemical reaction. Dropping a lightstick into hot water makes it glow more intensely, demonstrating that the reaction is faster at a higher temperature.

2. Concentration

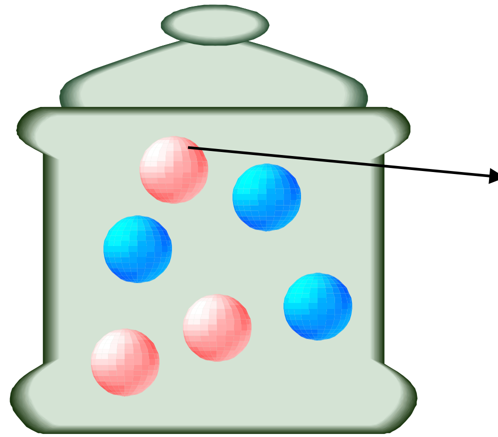
- The more reacting particles you have in a given volume, the greater the total number of collisions and the higher the rate of reaction.
- NOTE: As the reaction proceeds, it tends to slow down i.e. fastest at the beginning



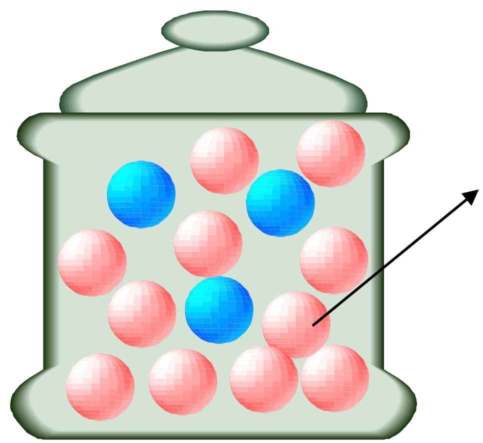
Low concentration = Few collisions High concentration = More collisions

Concentration

- If we make one reactant more concentrated then there are more particles in the same volume to react
- So the reaction goes faster.



There are less red particles in the same volume so there is less chance of a collision



There are more red particles in the same volume so there is more chance of a collision so the reaction goes faster

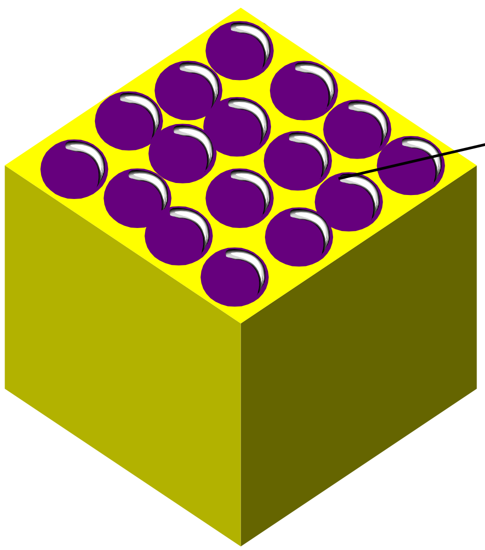
Examples:

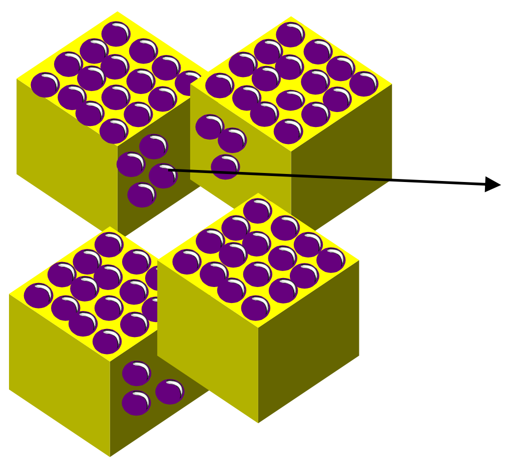
- Two antacid tablets will neutralize a given amount of acid faster than one tablet will.

- Higher concentrations of acid in rain erode marble faster than lower concentrations.

3. Surface Area or Particle Size

- If we make the pieces of the reactants smaller we increase the number of particles on the surface which can react.
- Therefore more space available for collisions to occur.
- This makes the reaction faster.

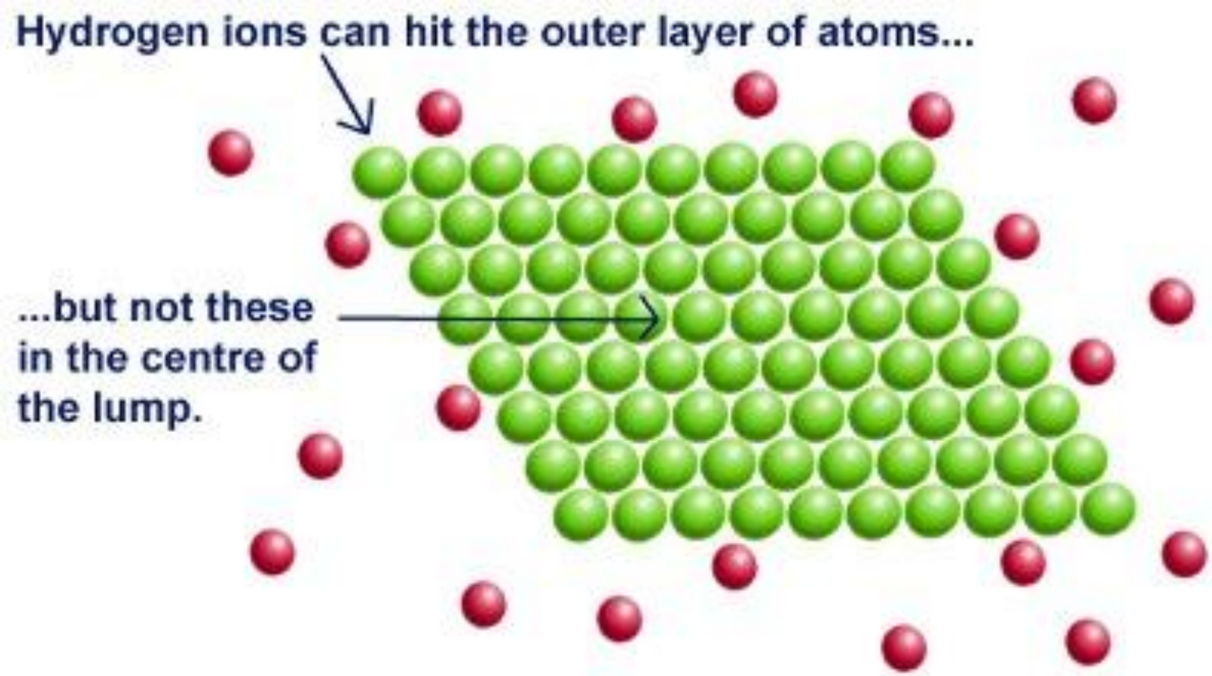


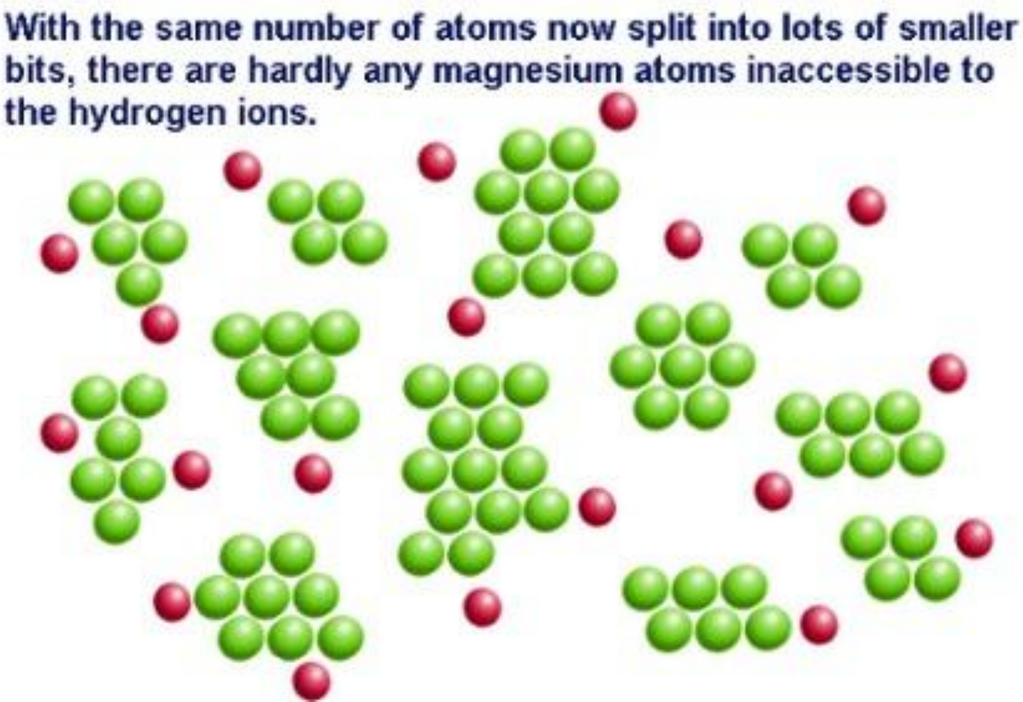


Surface Area or Particle Size

- One way to increase the surface area of solid reactants is to dissolve them
  - which separates the particles and makes them more accessible to other reactants.

- Grinding solids into a fine powder also increases the surface area of reactants
  - Small dust-like particles can be very dangerous, can be highly explosive


Hydrogen ions can hit the outer layer of atoms...
...but not these in the centre of the lump.


With the same number of atoms now split into lots of smaller bits, there are hardly any magnesium atoms inaccessible to the hydrogen ions.

4. Nature of Reactants

- **Individual properties of substances also affect reaction rates.**
- The scope of these properties is broad and there are few generalizations that you can apply consistently.
- **Some of the properties in this category are state of matter, molecular size, bond type and bond strength.**
- **Example: K is more reactive than Li**

- **A) State of Matter**
- ***Gases tend to react faster than solids or liquids:*** It takes energy to separate particles from each other. In order to burn candle wax, the solid wax has to be melted and then vaporized before it reacts with oxygen. Methane gas is already in the gas state so it burns faster than wax.
- ***Aqueous ions tend to react faster than species in other states of matter:*** Solid lead(II) nitrate will react with solid potassium iodide, but the reaction is really, really slow. That's because the ionic bonding in each reactant is strong and the ions in each compound are hard to separate from each other. When aqueous solutions of these compounds are mixed, the formation of lead(II) iodide is rapid. In aqueous solutions, the ions of each compound are dissociated. When the two solutions are mixed together, all that is required for a reaction to occur is contact between the lead(II) ions and the iodide ions.

- Reactions where all reactants are in the same state (homogeneous reactions) will occur at a faster rate than reactions where reactants are in different states (heterogeneous reactions).
- This occurs because reactants will have a greater opportunity of colliding.
- NOTE: stirring also increases the rate of reaction

- **B. Bond Type**
- Reactions involving ions tend to proceed faster than reactions involving molecular compounds.

- **C. Bond Strength**
- Reactions involving the breaking of weaker bonds proceed faster than reactions involving the breaking of stronger bonds.
- For example, double carbon to carbon bonds are stronger than single C-C bonds.

- **D. Number of Bonds/Molecular Size**
- **Reactions involving the breaking of fewer bonds per reactant proceed faster than those involving the breaking of a larger number of bonds per reactant.**
- The simple ion Fe^{2+} reacts faster than oxalate ($OOCCOO^{2-}$).
- Kerosene burns more slowly than methane because there are more bonds to be broken per molecule of kerosene than there are per molecule of methane. Kerosene is a larger molecule.

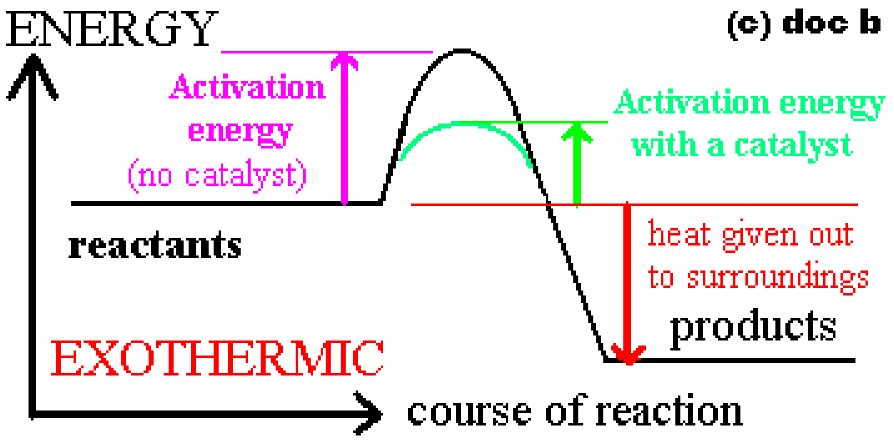
5. Catalyst

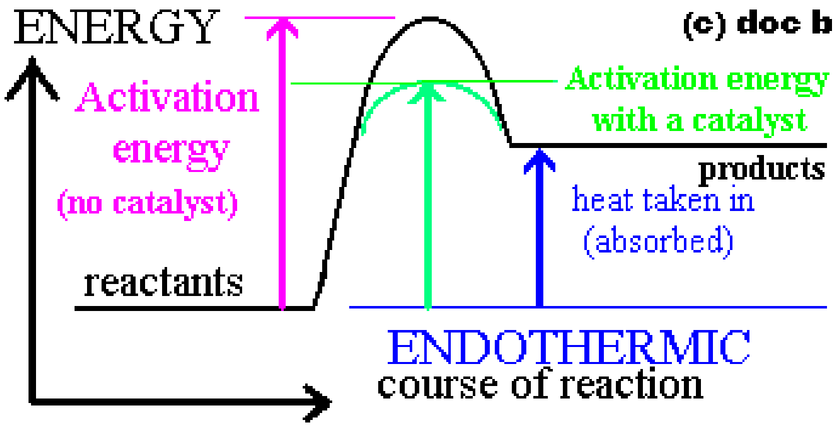
- A catalyst is a substance that increases the rate of a reaction without being changed or used up during the reaction
- They provide an alternative pathway for the reaction which has a lower activation energy
- This allows for more reactants to form products in a given amount of time, since it increases the number of effective collisions

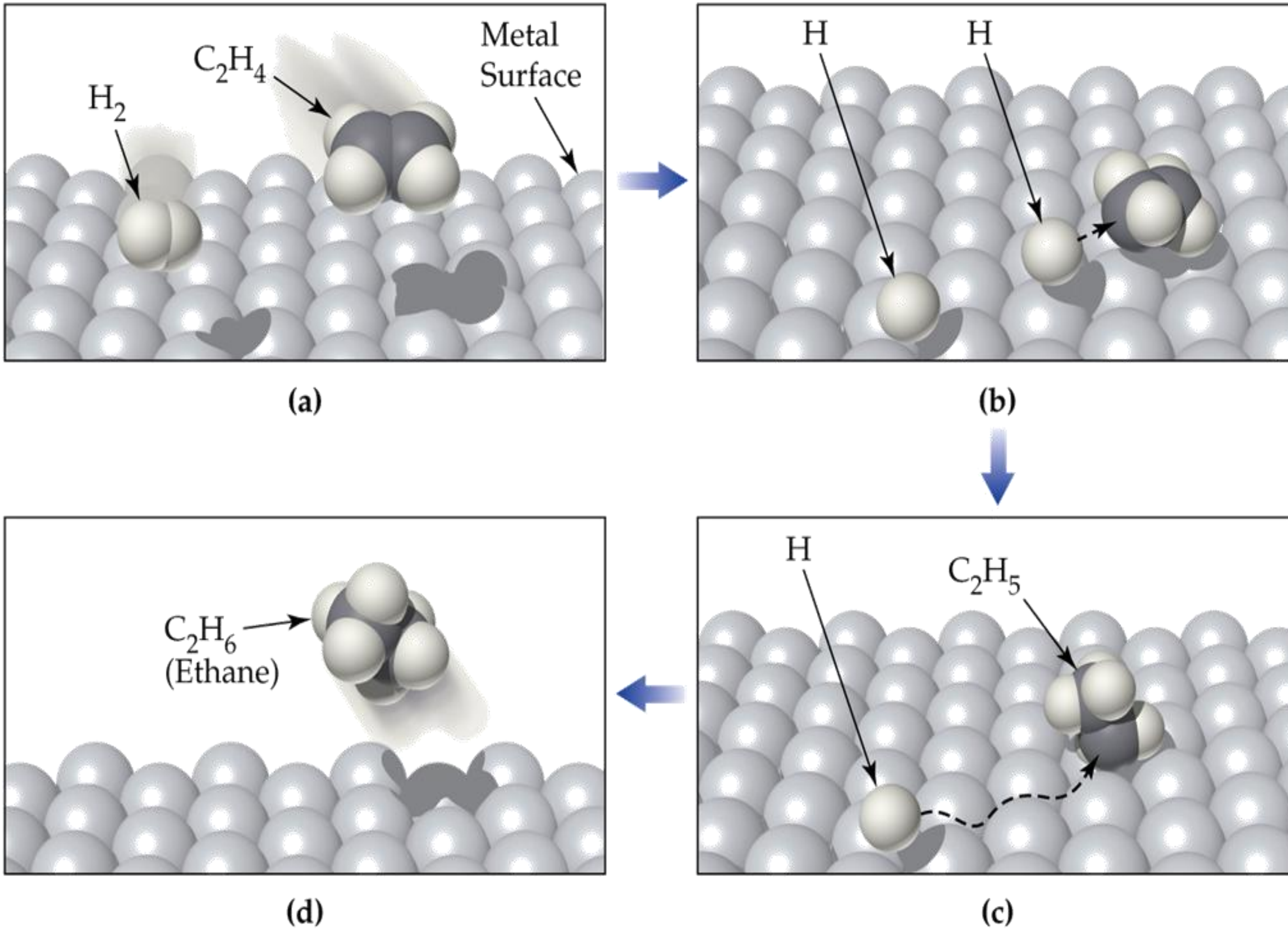
Catalyst

Add a <u>Catalyst</u> = Speeds up a reaction but is not used in the reaction

- ➢Lowers the activation energy

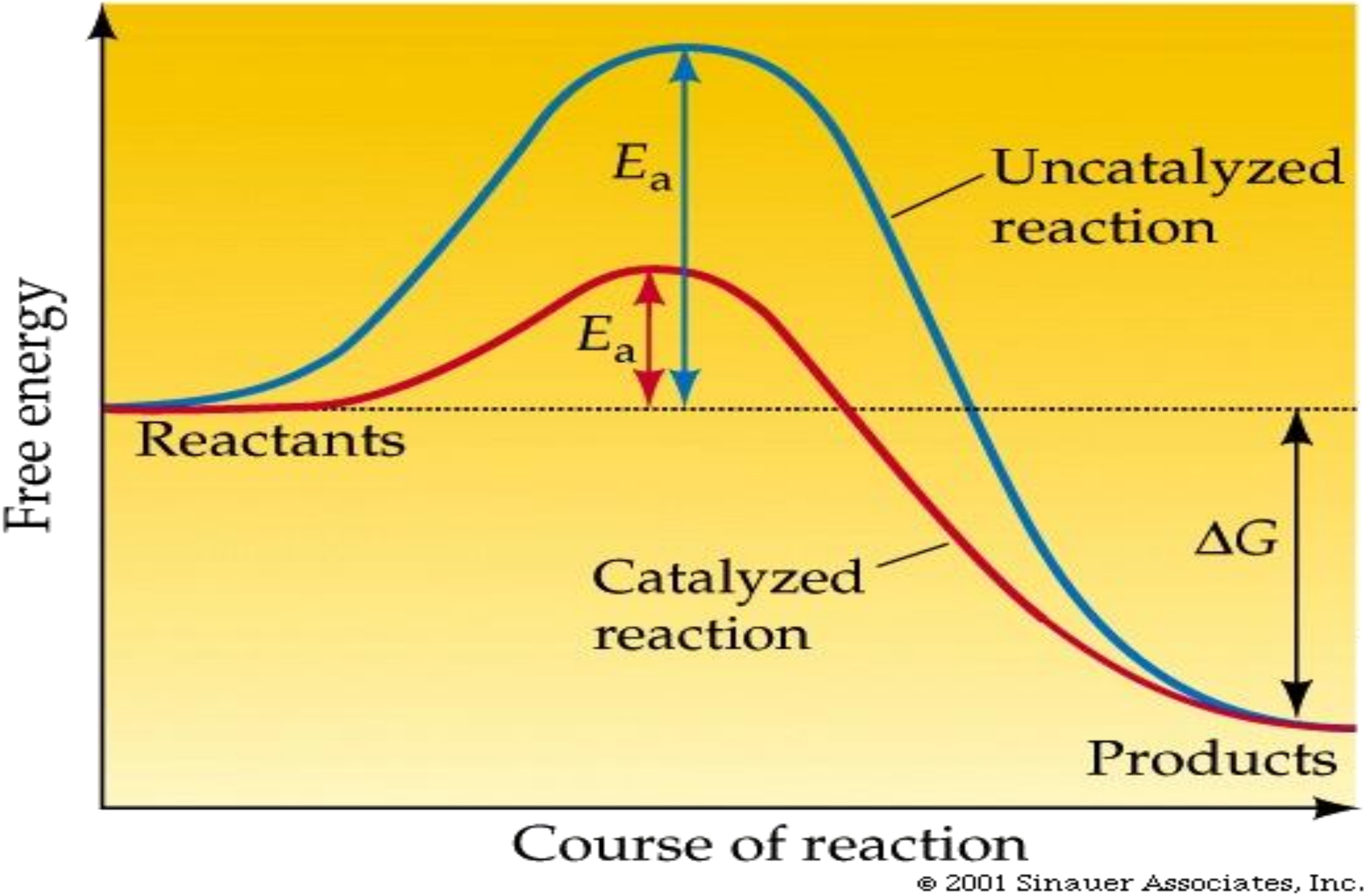






H2
C2H4
Metal Surface
(a)
H
H
(b)
H
C2H5
(c)
C2H6 (Ethane)
(d)

Catalyst



Free energy
E_a
E_a
Uncatalyzed reaction
Reactants
Catalyzed reaction
ΔG
Products
Course of reaction
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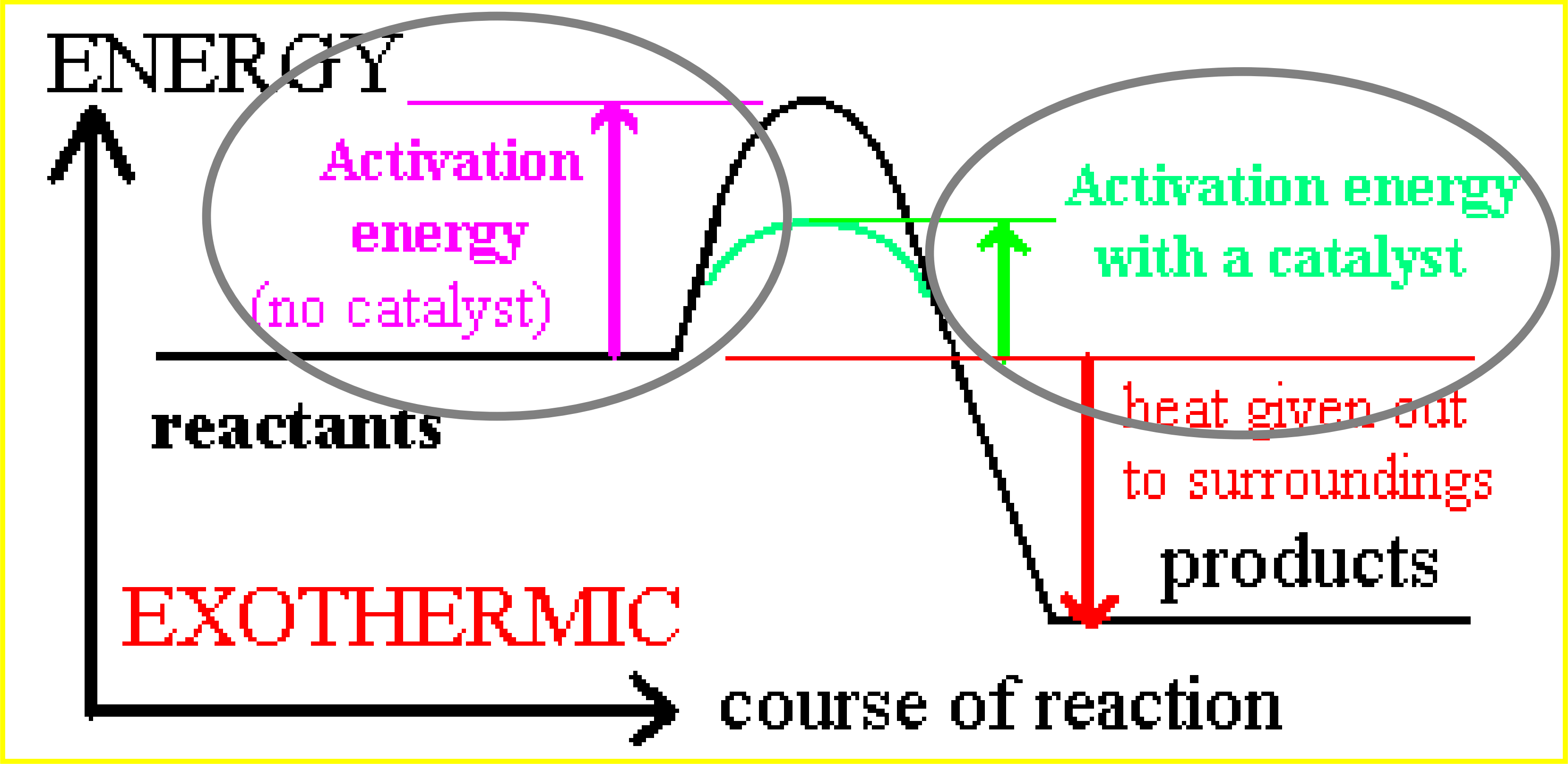
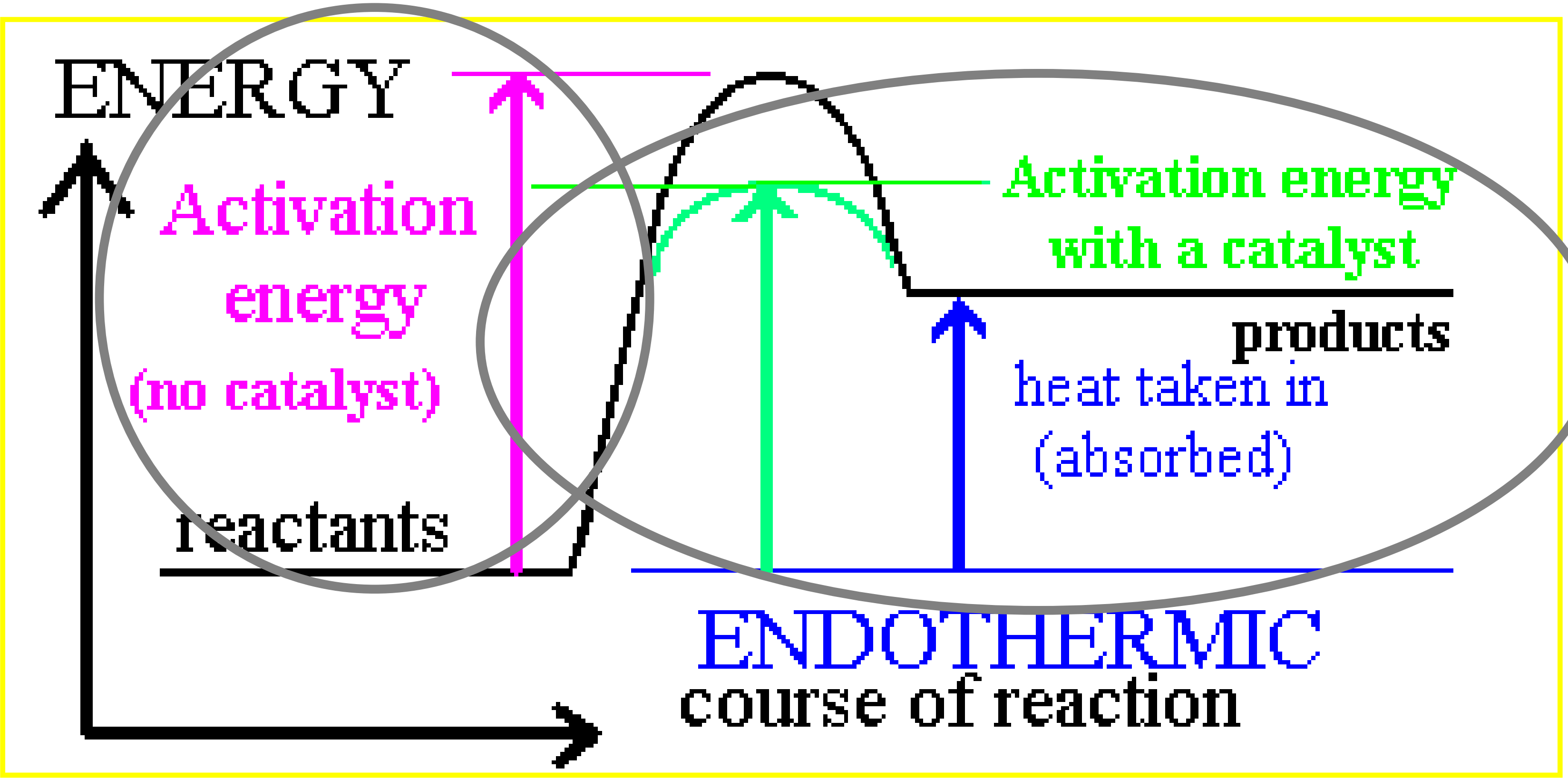
Exothermic Reaction with a Catalyst



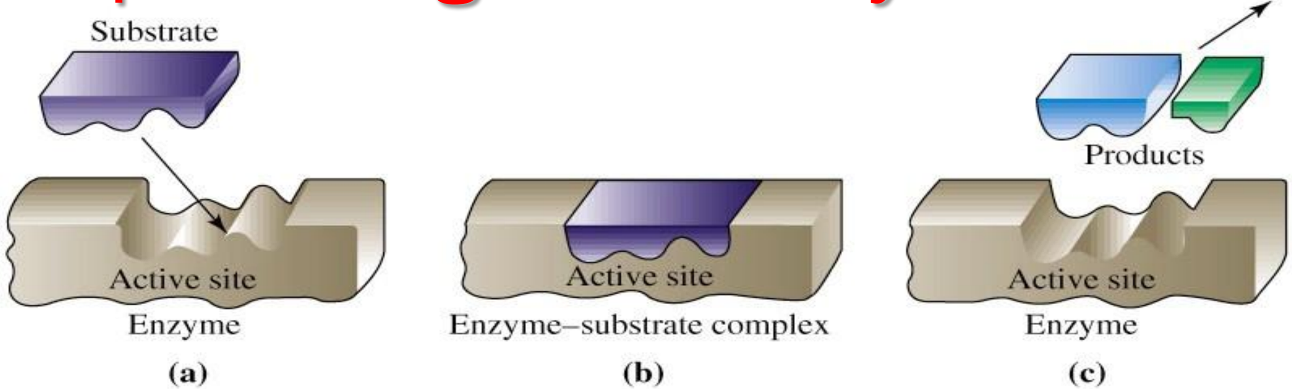
Endothermic Reaction with a Catalyst



Catalyst

- Since catalysts are not consumed during a reaction, they do not appear as reactants or products in the chemical equation
  - written above the reaction's arrow
- Catalysts are crucial for many life processes.
  - Your body temp is only 37 °C and cannot be raised significantly without danger
- Without catalysts, few reactions in the body would proceed fast enough at that temperature
  - Enzymes, biological catalysts, increase the rates of biological reactions

- When you eat a meal containing protein, enzymes in your digestive tract break down the protein molecules in a few hrs..
    - Without enzymes, the digestion of proteins at 37°C takes years
- An inhibitor is a substance that interferes with the action of a catalyst
    - An inhibitor could work by reacting with or “poisoning” the catalyst itself



Common examples of catalysts include:

- MnO_2 in the decomposition of H_2O_2
- Fe in the manufacture of NH_3
- Pt in the conversion of NO and CO to N_2 and CO_2

Summary

- Molecules must collide with the correct orientation and enough energy in order to react.

More collisions = Faster reaction rate.

Increasing Temperature, Concentration, Pressure, and Surface Area results in more collisions therefore faster reactions.

Adding a catalyst lowers the energy of activation thus faster reaction rate.

The Rate Law

The ***rate law*** expresses the relationship of the rate of a reaction to the rate constant and the concentrations of the reactants raised to some powers.

$$aA + bB \longrightarrow cC + dD$$

$$\text{Rate} = k\,[A]^{x}[B]^{y}$$

reaction is ***x*th order** in A

reaction is ***y*th order** in B

reaction is **(x +y)th order overall**

$F_2 (g) + 2ClO_2 (g) \longrightarrow 2FClO_2 (g)$

$$\text{rate} = k [F_2]^x[ClO_2]^y$$

Table 13.2 Rate Data for the Reaction between F_2 and ClO_2

	$[F_2](M)$	$[ClO_2](M)$	Initial Rate (M/s)
1.	0.10	0.010	1.2×10^{-3}
2.	0.10	0.040	4.8×10^{-3}
3.	0.20	0.010	2.4×10^{-3}

Double $[F_2]$ with $[ClO_2]$ constant

Rate doubles

Therefore $x = 1$

Quadruple $[ClO_2]$ with $[F_2]$ constant

Rate quadruples

Therefore $y = 1$

rate = $k [F_2][ClO_2]$

Reaction is 1st order for each of the reactants and 2nd order overall

Run #	Initial [A] ($[A]_0$)	Initial [B] ($[B]_0$)	Initial Rate (v_0)
1	1.00 M	1.00 M	1.25×10^{-2} M/s
2	1.00 M	2.00 M	2.5×10^{-2} M/s
3	2.00 M	2.00 M	2.5×10^{-2} M/s

What is the order with respect to A? 0

What is the order with respect to B? 1

What is the overall order of the reaction? 1