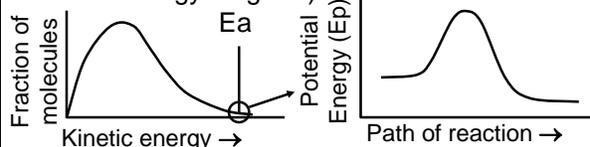


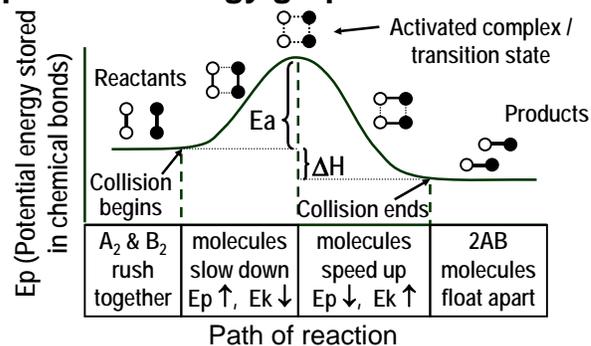
kinetic vs. potential energy diagrams

- Recall the Maxwell-Boltzman distribution (i.e. kinetic energy diagram)



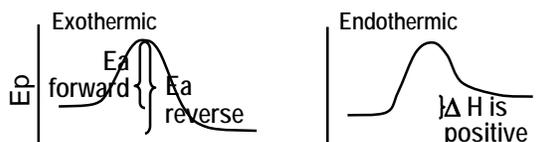
- The E_a is a critical point. To examine it more closely we can use a potential energy graph
- The axes are not the same, thus the E_p graph is not a blow up of the E_k graph; however it does correspond to the part of the E_k graph that is circled

potential energy graph: a closer look



$E_p + E_k = \text{constant throughout}$ Overall $E_p(\text{reactants}) > E_p(\text{products})$
 $E_k(\text{reactants}) < E_k(\text{products})$

E_p graph: Important points



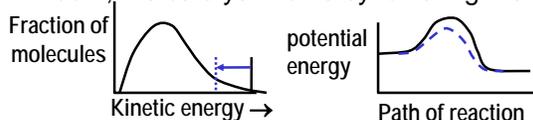
- Forward and reverse reactions are possible
- E_a is the difference between E_p at transition state and initial or final E_p
- The graph depicts an exothermic reaction. Endothermic reactions are also possible
- ΔH is the difference between initial and final E_p . It is -ve for exothermic, +ve for endothermic

The collision theory

- Related to the E_p graph is the "collision theory" - the idea that for molecules to react they must meet with sufficient force
- Factors that affect reaction rate can be explained via the collision theory:
- Increased temperature causes molecules to move faster (increased number of collisions per unit time and greater kinetic energy)
- Increased concentration means more collisions
- Homogenous reactions occur faster because reacting molecules collide more frequently
- Catalysts decrease E_a , decreasing the amount of kinetic energy needed to overcome E_a

Catalysts

- Recall, catalysts speed a reaction
- This can be explained by the E_k or E_p graphs
- In both, the catalyst works by lowering the E_a :

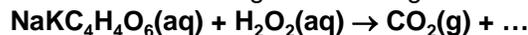


- Catalysts speed forward and reverse reactions
- However, most reactions favour the side that has the lowest potential energy (most stable)
- Catalysts are heterogenous or homogenous
- They provide a substrate (p. 768) for a reaction or they can bond temporarily to a molecule, increasing the odds of a favourable meeting

Transition state lab: purpose

Purpose: 1) to visualize an activated complex, 2) to observe the influence of a catalyst

We will be examining the following reaction:



Procedure:

- Turn hot plates immediately to medium heat
- Get a 10 mL graduated cylinder, a 100 mL beaker, a test tube, and a rubber stopper.
- Weigh 1.7 g $\text{NaKC}_4\text{H}_4\text{O}_6$. Add to beaker along with 10 mL distilled H_2O . Swirl to dissolve.
- Add 4.5 mL of 10% H_2O_2 to beaker. Heat.
- Get 5 mL of CoCl_2 but don't add it yet.

Transition state lab: procedure

Procedure:

- As soon as tiny bubbles start to form and rise, remove the beaker from the hot plate. Add the CoCl_2 at this point.
- Record your observations (in order to answer the questions). Clean up – wash everything down the drain, wipe off your lab bench.

Questions: answer on a separate sheet of paper

- Look at the chemical equation that represents the reaction. What physical sign will there be when a reaction is occurring?
- The products of the reaction are colourless. What colour are the reactants?

Transition state lab: conclusions

Questions: read 18.11 (pg. 767 – 769)

- What was the catalyst in the lab? What colour was it? Is it homogenous or heterogeneous?
- At the beginning of step 5, both reactants were present; why was there no reaction? (Illustrate with a E_k diagram).
- Why is the reaction still slow after heat is added? (illustrate using the E_k diagram)
- Was the catalyst a different colour at the end of the experiment than at the beginning?
- What colour was the activated complex?
- Illustrate the affect the catalyst had on the reaction (using both E_k and E_p diagrams)