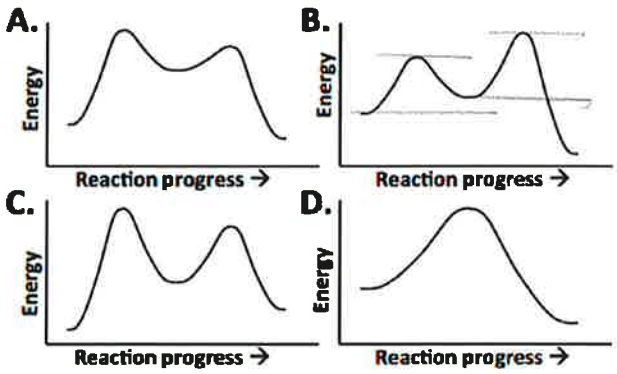
Classroom Activity Chapter 13 Number 4 — 1 February 2017

1. Consider these four reaction profiles:



1a. Which ones are exothermic? (A) (B) C (D)

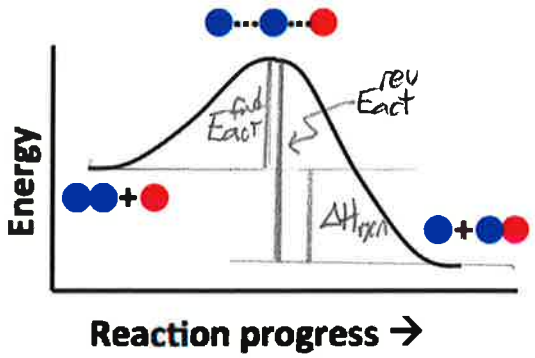
1b. Which reaction is the slowest? A B (C) D

1c. For which of the two step reactions is the first step the slow step? (A) B (C) D

1d. For which reaction is there the best chance of detecting an intermediate? A B (C) D

1e. For which reaction is there no intermediate? A B C (D)

2. Let **A = red circles** and **B = blue circles**.



2a. Is this reaction (exothermic) or endothermic?

2b. Label E_{act}^{fwd}, E_{act}^{rev}, and ΔH_{rxn}.

2c. Write a formula relating ΔH_{rxn} in terms of E_{act}^{fwd} and E_{act}^{rev}.

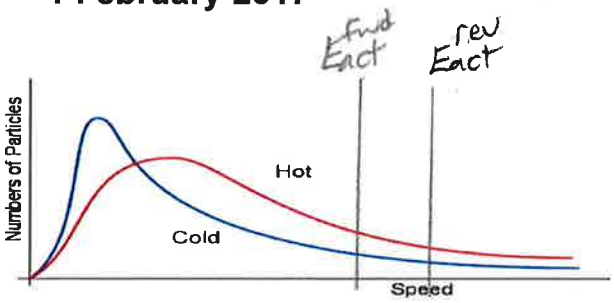
$\Delta H_{rxn} = E_{act}^{fwd} - E_{act}^{rev} < 0$

2d. Write a balanced chemical reaction for this reaction.

$A_2 + B \rightarrow A + AB$

2e. Would increasing the temperature increase the rate of the reaction? (Yes) or No

2f. For this particular reaction, draw a vertical line to represent E_{act}^{fwd} and E_{act}^{rev} on the energy population graph below. Note that "speed" is related to kinetic energy and collision energy.



3. The following mechanism has been proposed for the reaction between nitrogen monoxide and chlorine gases.

Step 1. $NO + Cl_2 \rightarrow NOCl_2$ Slow step

Step 2. $NO + NOCl_2 \rightarrow 2\ NOCl$

3a. What is the overall reaction?

$2\ NO + Cl_2 \rightarrow 2\ NOCl$

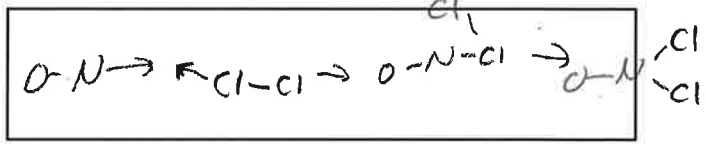
3b. Identify any intermediates.

$NOCl_2$

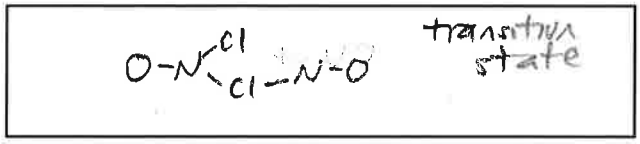
3c. Given the first step is slow, what is the rate law?

$rate = k[NO]^1[Cl_2]^1$

3d. Sketch the reactants → transition state → products for Step 1.



3e. Sketch the reactants → transition state → products for Step 2.



3f. Which of the reaction profiles given in Question 1 best describes this mechanism, given that the overall reaction is exothermic? (A) B C D

3g. Suppose instead that the second step is the slow step. What is the rate law in this case? Remember to substitute for intermediates, for which the concentration cannot be measured in lab and so the rate law cannot be supported by experimental results.

$rate = k[NO][NOCl_2]$

$K_c = [NOCl_2]/[NO][Cl_2]$

$rate = k K_c [NO][Cl_2][NO]$

$= k[NO]^2[Cl_2]$