

Exam I – CHM 152/154 – Fall 2005 – Dr. Smith – Form A

Crib

- Which of the following statements is FALSE?
 - The units for a reaction's rate are M/time.
 - A reaction rate depends on the product concentrations.**
 - As the reactant concentration increases, the rate increases.
 - As temperature increases, the rate increases.
 - The half life is the time it takes for 50% of the reactant to react.
- What is the average rate if the concentration of reactant at 15.0 seconds is 0.983M, and then the concentration of reactant at 47.0 seconds is 0.777M?

$$\text{Ave rate} = \Delta M / \Delta t = (0.777M - 0.983M) / (15.0s - 47.0s) = -0.206M / 32.0s = -6.44 \times 10^{-3} \text{ M/s}$$

- For a reaction the rate law is found to be $\text{rate} = k[\text{SO}_2]^2[\text{O}_2]$. What is the overall reaction order? **2 + 1 = 3, third order**
- What percentage of a sample would remain after 588 seconds if the corresponding rate constant for this first order reaction were $3.53 \times 10^{-3} \text{ s}^{-1}$?

$$\ln(x/100) = (-3.53 \times 10^{-3} \text{ s}^{-1})(588\text{s})$$

$$\ln x = -2.0756 + \ln 100 = 2.5295$$

raise both sides to e

$$x = 12.5\%$$

- What is the half-life for a first order reaction if the rate constant is 0.335 s^{-1} ?

$$t_{1/2} = 0.693 / 0.335 \text{ s}^{-1} = 2.07\text{s}$$

- To find the rate constant for a first order reaction graphically one would plot:
 - $[\text{A}]_t$ versus time
 - $\ln k$ versus E_a
 - $\ln k$ versus $1/T$
 - $1/[\text{A}]_t$ versus time
 - $\ln[\text{A}]_t$ versus time**

- A second order reaction has a half-life of 24.5 seconds. If the initial reactant concentration is 0.505M, how long will it take for the reactant concentration to reach 0.450M?

$$24.5\text{s} = 1 / (0.505\text{M} \times k)$$

$$k = 0.80824\text{M}^{-1}\text{s}^{-1}$$

$$1/0.450\text{M} = (0.80824\text{M}^{-1}\text{s}^{-1})t + 1/0.505\text{M}$$

$$2.222\text{M}^{-1} = (0.80824\text{M}^{-1}\text{s}^{-1})t + 1.980\text{M}^{-1}$$

$$0.242\text{M}^{-1} / (0.80824\text{M}^{-1}\text{s}^{-1}) = t$$

$$t = 2.99\text{s}$$

8. The following experimental data was collected for this reaction. $2\text{NO} + 2\text{H}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O}$. Write the rate law for this reaction with the correct exponents on your answer sheet.

Trial	[NO] (M)	[H ₂] (M)	Rate (M/s)
1	0.60	0.15	0.076
2	0.60	0.30	0.15
3	0.60	0.60	0.30
4	1.20	0.60	1.21

When NO doubles in trials 3 to 4, the rate quadrupled so that is second order with respect to NO.
 When H₂ doubles in trials 2 to 3, the rate doubles so that is first order with respect to H₂.
 So the rate law is $\text{rate} = k[\text{NO}]^2[\text{H}_2]$

9. Calculate the rate constant for the reaction above (number 9). Write the value of the rate constant and units on your answer sheet.

Using data from trial one: $0.076\text{M/s} = k(0.60\text{M})^2(0.15\text{M})$

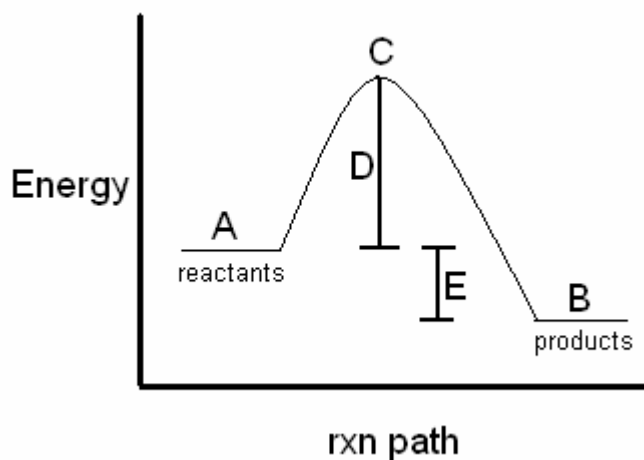
$$k = (0.076\text{M/s}) / 0.54\text{M}^3$$

$$k = 1.4 \text{ M}^{-2}\text{s}^{-1}$$

(Only two sig dig)

10. Which of the following factors will NOT increase the number of successful collisions between reactant molecules?
- Increasing the reactant concentration
 - Increasing the temperature
 - Aim the reactants at 180° angles
 - Increasing the surface area of the reactant (exposing more reactant)
 - Increasing the product concentration**

11. In the following potential energy diagram, which letter represents the transition state? **C**



12. In the above potential energy diagram, which letter represents the activation energy? **D**

13. Is the reaction diagramed above endothermic or **exothermic**?

14. The rate constant for a reaction at 35.0°C is 0.0660 s⁻¹. At 75.0°C it is 0.923 s⁻¹. Find the activation energy in kJ/mol.

Remember to put Celsius in Kelvin.

$$\ln(0.0660\text{s}^{-1}/0.923\text{s}^{-1}) = (E_a/R)(1/348\text{K} - 1/308\text{K})$$

$$(8.314\text{J/molK})(-2.63797) = E_a (-3.731 \times 10^{-4}\text{K}^{-1})$$

$$E_a = 58.8 \text{ kJ/mol}$$

15. What percentage of a sample would remain after 8 half lives?

Divide 100% by 2 eight times = 0.39%

(You can have more sig dig since it is technically exactly 100.000000 infinity % we start with)

16. Which of the following statements is FALSE regarding catalysis?

- a. A catalyst lowers the activation energy of a reaction.
- b. A catalyst is either homogeneous or heterogeneous.
- c. A catalyst increases the reactant's energy.**
- d. A catalyst provides an alternate reaction pathway with lower energy.
- e. A catalyst is not consumed in the reaction.

17. A bimolecular elementary step in a reaction mechanism follows

- a. zero order kinetics.
- b. first order kinetics.
- c. second order kinetics.**
- d. third order kinetics.
- e. need more information.

18. The following reaction was experimentally found to be third order with a rate law of rate = k[NO]²[O₂] :
2NO(g) + O₂(g) → 2NO₂(g). Which of the following steps in the reaction mechanism must be the rate-determining step?

- a. NO + NO → N₂O₂ step one **rate = k[NO][NO] which doesn't match so it must be step two**
- b. N₂O₂ + O₂ → 2NO₂ step two

19. For the reaction above, which species is an intermediate?

- a. NO b. O₂ **c. N₂O₂** c. NO₂

20. In a graph of lnk versus 1/T, the slope is found to be -765 K. What is the activation energy?

$$\text{slope} = -E_a/R$$

$$\text{slope}(R) = -E_a$$

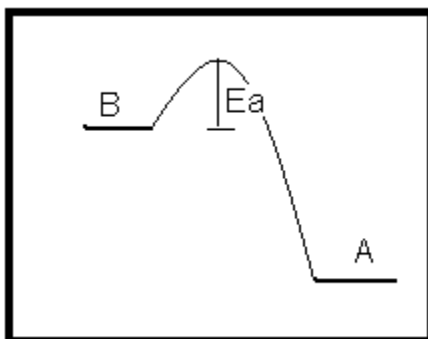
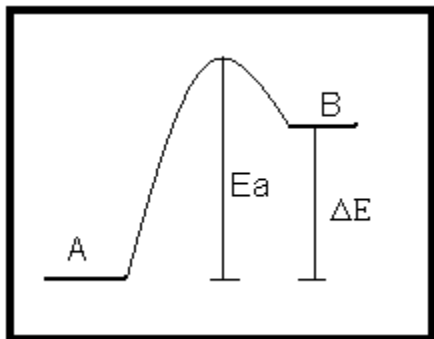
$$(-765\text{K})(8.314\text{J/molK}) = -E_a$$

$$E_a = 6360 \text{ J/mol}$$

Bonus 1: Second order kinetic data can be plotted much like first order data. To obtain the rate constant for a second order reaction graphically as the slope, what would you plot?

Look at the second order integrated rate law. It is already in the form $y = mx + b$. So plot $1/[A]_t$ versus time and you get k as your slope.

Bonus 2: The reaction $2\text{NOCl} \rightarrow 2\text{NO} + \text{Cl}_2$ has an E_a of 102 kJ/mol and a ΔE_{rxn} of 75.5 kJ. What is the activation energy for the reverse reaction? (*Hint: Draw the potential energy diagram, then draw it backwards and think about what the activation energy would be*)



In the first diagram the reaction is forwards and is endothermic because ΔE_{rxn} is positive.

In the second diagram the reaction is backwards and is exothermic now. E_a is much smaller and is $102 - 75.5 = 26.5$ kJ