

Exam 2A

Chem 1142

Spring 2019

Name: _____

MULTIPLE CHOICE. [2 pts ea.] Record the best response on the scantron sheet. [50 pts total.]

Q1. Which version of the exam do you have?

- A) 2A
- B) 2B
- C) 2C
- D) 2D

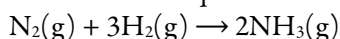
Q2. Given the chemical equation:



The reaction rate is equal to:

- A) $+\frac{1}{3} \Delta[\text{SO}_3]/\Delta t$
- B) $+\frac{1}{2} \Delta[\text{SO}_2]/\Delta t$
- C) $-\Delta[\text{O}_2]/\Delta t$
- D) $-\Delta[\text{O}]/\Delta t$

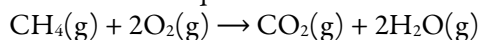
Q3. Given the chemical equation:



If the hydrogen concentration decreases by 0.033 M over a 15.0 s period, the average rate of reaction is:

- A) 0.033 M/s
- B) 0.50 M/s
- C) 2.2×10^{-3} M/s
- D) 7.3×10^{-4} M/s

Q4. Given the chemical equation:



The rate law is:

- A) impossible to determine
- B) $\text{rate} = k [\text{CH}_4][\text{O}_2]^2$
- C) $\text{rate} = k [\text{CO}_2][\text{H}_2\text{O}]^2$
- D) $\text{rate} = \frac{[\text{CO}_2][\text{H}_2\text{O}]^2}{[\text{CH}_4][\text{O}_2]^2}$

Q5. The units for the rate constant, k , for a first-order reaction are:

- A) M
- B) $\text{M}^{-1}\cdot\text{s}^{-1}$
- C) $\text{M}\cdot\text{s}^{-1}$
- D) s^{-1}

- Q6. If the rate of a reaction increases by a factor of nine when the concentration of one reactant is increased by a factor of three, the order with respect to this reactant must be:
- first
 - second
 - third
 - sixth
- Q7. A plot of $\ln [A]$ vs time is linear for the reaction: $A \rightarrow B$. This information can be used to determine:
- the reaction is zeroth order with respect to A
 - the reaction is first order with respect to A
 - the reaction is second order with respect to A
 - the slope of the graph is equal to $-E_A/R$
- Q8. A first order reaction takes 12 seconds for the concentration of A (the reactant) to decrease from 1.0 M to 0.50 M. How long will it take for the concentration to decrease from 1.0 M to 0.25 M?
- 6.0 s
 - 12 s
 - 24 s
 - 48 s
- Q9. The highest point on an energy vs. reaction progress plot is called the:
- transition state
 - reaction intermediate
 - frequency factor
 - degree of exothermicity
- Q10. In general, reactions speed up with increasing temperature. This is primarily due to the fact that increasing temperature...
- increases the reactant concentrations
 - increases the frequency of collisions
 - increases the probability of a collision with $E \geq E_A$
 - increases the probability of a collision with a successful orientation
- Q11. What type of plot can be used to determine the activation energy for a reaction?
- $\ln [A]$ vs. t
 - $1 / [A]$ vs. t
 - $1 / k$ vs. $1 / T$
 - $\ln (k)$ vs. $1 / T$
- Q12. The chemical equation, $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ is postulated to occur via the mechanism:
- $N_2(g) \rightarrow 2N(g)$ *slow*
 - $2N(g) + H_2(g) \rightarrow N_2H_2(g)$ *fast*
 - $N_2H_2(g) + 2H_2(g) \rightarrow 2NH_3(g)$ *fast*
- What should the rate law be?
- rate = $k [N_2][H_2]^3$
 - rate = $k [N_2]$
 - rate = $k [N]^2$
 - rate = $k [N_2H_2][H_2]^2$

- Q13. What is the molecularity of the elementary step: $\text{N}_2(\text{g}) \rightarrow 2\text{N}(\text{g})$?
- unimolecular
 - bimolecular
 - termolecular
 - quadmolecular
- Q14. Catalysts work by:
- lowering $\Delta H^\circ_{\text{rxn}}$
 - lowering the collision frequency
 - lowering the concentration of intermediates
 - lowering the activation energy
- Q15. The chemical equilibrium, $\text{MgCO}_3(\text{s}) \rightleftharpoons \text{MgO}(\text{s}) + \text{CO}_2(\text{g})$, has an equilibrium constant of 3.2×10^{-10} at room temperature. At this temperature, an equilibrium mixture should consist of:
- mainly CO_2 and MgO , with small amounts of MgCO_3
 - mainly MgCO_3 , with small amounts of MgO and CO_2
 - an equal mixture of MgCO_3 , MgO , and CO_2
 - An equal mixture of MgO and MgCO_3 . with a small amount of CO_2
- Q16. The reaction, $\text{A}(\text{g}) \rightleftharpoons 2\text{B}(\text{g})$ has an equilibrium constant of 100. Predict the equilibrium constant for the reaction, $\text{B}(\text{g}) \rightleftharpoons \frac{1}{2} \text{A}(\text{g})$
- 10
 - 50
 - 0.1
 - 25
- Q17. For which equilibrium reaction will K_p always equal K_c ?
- $\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons \text{H}_2\text{O}_2(\text{l})$
 - $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{l})$
 - $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightleftharpoons \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$
 - $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$
- Q18. For the equilibrium,
- $$2\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{CO}(\text{g})$$
- the partial pressures of gases above pure carbon are 0.010 atm (O_2) and 4.3 atm (CO) at 23 °C. What is K_p equal to?
- 0.043
 - 4300
 - 8.61
 - 1800
- Q19. For the equilibrium,
- $$2\text{A}(\text{aq}) \rightleftharpoons \text{B}(\text{aq})$$
- if we start out with $[\text{A}] = 1.00 \text{ M}$ and $[\text{B}] = 0.00 \text{ M}$, and end up with $[\text{A}] = 0.80 \text{ M}$ at equilibrium, what must the equilibrium constant, K_c , be?
- 0.16
 - 0.25
 - 0.31
 - 0.80

- Q20. Given the equilibrium,

$$\text{NO}(\text{g}) + \text{H}_2(\text{g}) \rightleftharpoons \text{H}_2\text{NO}(\text{g})$$
 which has an equilibrium constant, $K_p = 0.40$ at room temperature, predict the direction the reaction will proceed to reach dynamic equilibrium if $p_{\text{NO}} = 0.10$ atm, $p_{\text{H}_2} = 0.20$ atm, and $p_{\text{H}_2\text{NO}} = 0.10$ atm.
- A) will shift to the left
 B) will not shift—already at equilibrium
 C) will shift to the right
 D) not possible to determine the shift
- Q21. Given the equilibrium,

$$\text{NO}(\text{g}) + \text{H}_2(\text{g}) \rightleftharpoons \text{H}_2\text{NO}(\text{g}) \quad ; \Delta H^\circ_{\text{rxn}} = -34.0 \text{ kJ}$$
 Predict the direction of shift if the concentration of $\text{H}_2(\text{g})$ is decreased
- A) will shift to the left
 B) will neither shift left nor right
 C) will shift to the right
 D) not possible to determine the shift (must use an ICE-chart)
- Q22. Given the equilibrium,

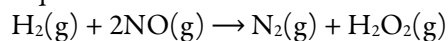
$$\text{NO}(\text{g}) + \text{H}_2(\text{g}) \rightleftharpoons \text{H}_2\text{NO}(\text{g}) \quad ; \Delta H^\circ_{\text{rxn}} = -34.0 \text{ kJ}$$
 Predict the direction of shift if the total volume is decreased
- A) will shift to the left
 B) will neither shift left nor right
 C) will shift to the right
 D) not possible to determine the shift (must use an ICE-chart)
- Q23. Given the equilibrium,

$$\text{NO}(\text{g}) + \text{H}_2(\text{g}) \rightleftharpoons \text{H}_2\text{NO}(\text{g}) \quad ; \Delta H^\circ_{\text{rxn}} = -34.0 \text{ kJ}$$
 Predict the direction of shift if the reaction temperature is decreased
- A) will shift to the left
 B) will neither shift left nor right
 C) will shift to the right
 D) not possible to determine the shift (must use an ICE-chart)
- Q24. When forming the expression for an equilibrium constant, what substances have an effective concentration/pressure of one (1)?
- A) pure solids
 B) aqueous solutes
 C) gases
 D) elements
- Q25. In the Arrhenius equation, the pre-exponential factor, A , is often written as a product of two other factors: $A = p \times Z$. The factor “ Z ” represents the collision frequency. What does the factor “ p ” represent?
- A) The probability of a collision with $E \geq E_A$
 B) The probability of a collision with proper orientation
 C) The probability of a collision with a catalyst
 D) The probability of a collision with $E \geq \Delta H^\circ_{\text{rxn}}$

Short Response.

Show ALL work to receive credit.

Q26. [12 pts.] A chemist is trying to determine the rate law for a reaction between $\text{H}_2(\text{g})$ and $\text{NO}(\text{g})$. The chemical equation is:



The chemist decides to use the method of initial rates, whereby different experiments are carried out with different initial concentrations, and the rate at the beginning of the experiment is measured.

The collected data looks like:

Experiment #	$[\text{H}_2] / \text{M}$	$[\text{NO}] / \text{M}$	initial rate / $\text{M}\cdot\text{s}^{-1}$
1	0.10	0.20	3.2×10^{-5}
2	0.20	0.20	1.28×10^{-4}
3	0.10	0.35	3.2×10^{-5}

From this data, determine the rate law, showing all work.

What is the overall order of this reaction? _____

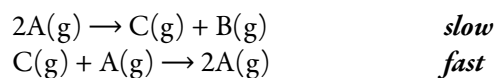
What is the value of the rate constant, including units? _____

Q27. [13 pts.] The chemical equation: $A(g) \rightarrow B(g)$ is found to have a rate law, $rate = k [A]^2$.

i) If the initial concentration of A is 0.25 M, and the rate constant is $0.011 \text{ M}^{-1}\text{s}^{-1}$, calculate how long it will take for the concentration to reach 0.0025 M.

ii) What happens to the half-life for this reaction over time? Explain.

iii) A possible mechanism for the reaction is given below:



Is this mechanism consistent with the rate law given at the beginning of this problem? What can you conclude (if anything) about this mechanism?

Q28. [12 pts.] **Using complete and grammatically correct sentences, answer the following questions:**

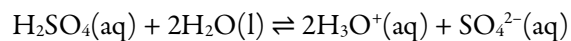
i) What is meant by the term: heterogenous catalyst?

ii) What is a reaction intermediate?

iii) What is an "Arrhenius plot?"

Give mathematical expressions for the following:

iv) Write the expression for K_c for the equilibrium:



v) Write the expression for K_p for the equilibrium



Q29. [13 pts.] The chemical reaction,
$$\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightleftharpoons 2\text{HBr}(\text{g})$$
has an equilibrium constant, $K_p = 64.0$ at $23\text{ }^\circ\text{C}$.

If the initial pressure of all three gases (H_2 , Br_2 , and HBr) is 0.250 atm , calculate the final equilibrium pressures of each gas.

BONUS Question:

What is the name given to the slowest step in a reaction mechanism?

Periodic Table of the Elements

IA 1	IIA 2											IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	VIIIA 18	
1 H 1.01																		2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18	
11 Na 22.99	12 Mg 24.31	3	4	5	6	7	8	9	10	11	12	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95	
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92160	34 Se 78.96	35 Br 79.90	36 Kr 83.80	
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc [98]	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29	
55 Cs 132.91	56 Ba* 137.33	71 Lu 174.97	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.20	83 Bi 208.98	84 Po [210]	85 At [210]	86 Rn [222]	
87 Fr [223]	88 Ra** [226]	103 Lr [262]	104 Rf [261]	105 Db [262]	106 Sg [266]	107 Bh [264]	108 Hs [265]	109 Mt [268]	110	111	112	113	114	115	116	117	118 [293]	
		* 57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [145]	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04			
		** 89 Ac [227]	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np [237]	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]			

$$R = 8.3145 \frac{\text{J}}{\text{mol} \cdot \text{K}} = 0.08206 \frac{\text{atm} \cdot \text{L}}{\text{mol} \cdot \text{K}}$$

$$k = A e^{-\frac{E_A}{RT}}$$

$$\ln k = -\frac{E_A}{R} \cdot \frac{1}{T} + \ln A$$

$$\ln \left(\frac{k_2}{k_1} \right) = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

• 0-order: $[A]_t = -kt + [A]_0$

$$t_{1/2} = \frac{[A]_0}{2k}$$

• 1-order: $\ln[A]_t = -kt + \ln[A]_0$ $\ln \left(\frac{[A]_t}{[A]_0} \right) = -kt$

$$t_{1/2} = \frac{0.693}{k}$$

• 2-order: $\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$

$$t_{1/2} = \frac{1}{[A]_0 k}$$

$$K_p = K_c (RT)^{\Delta n_g}$$

Given: $ax^2 + bx + c = 0$, then $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$