CHEM 220 - Exam #1 - Spring 2018
Answer Section

MULTIPLE CHOICE

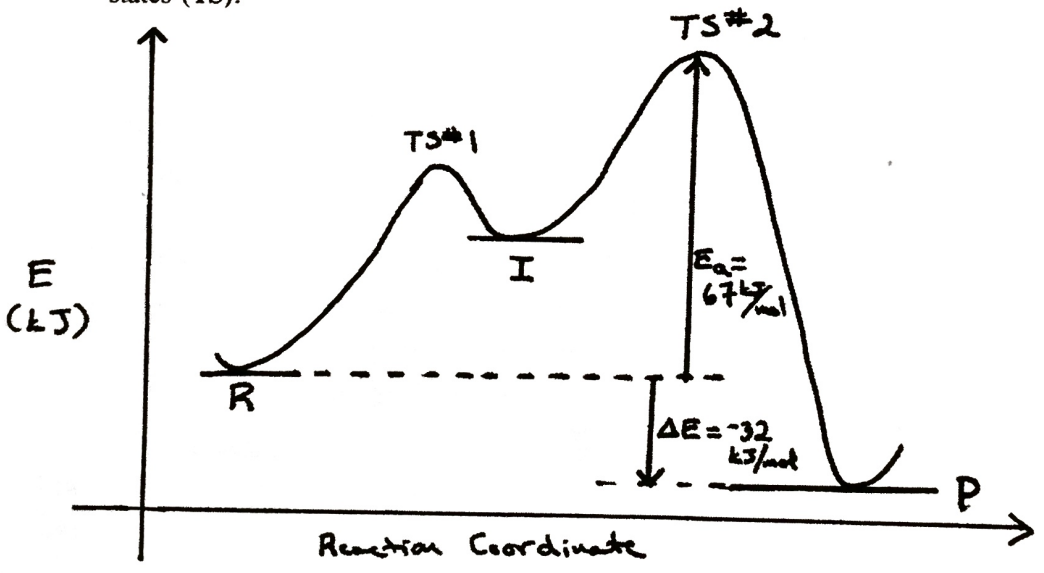
1.	ANS: E	PTS: 1
2.	ANS: D	PTS: 1
3.	ANS: E	PTS: 1
4.	ANS: C	PTS: 1
5.	ANS: D	PTS: 1
6.	ANS: A	PTS: 1
7.	ANS: B	PTS: 1

Free Response (12 points)

8. Consider the general reaction: R → P

A) Sketch a reaction-energy diagram for the reaction, reflecting the following characteristics:

- The reaction is exothermic with **ΔE = –32 kJ/mol.**
- The reaction occurs by a two-step mechanism. The second step is the slow step.
- Indicate the activation energy (**E_a**) for the forward reaction and **ΔE** on the plot.
- Label the relative energy position of reactants (**R**), products (**P**), intermediates (**I**), and transition states (**TS**).



B) If the activation energy for the reaction is $E_{a,forward}$ = 67 kJ/mol, what is $E_{a,reverse}$?

$E_{a,rev}$ = 99 kJ/mol

FREE RESPONSE: *Please answer the following questions in the spaces provided.* ***You must show all work and units! Consider significant figures. Use proper scientific notation when using exponential notation***

9. (10 points) Consider the gas-phase reaction: $2\ NO_{2(g)} + O_{2(g)} \rightleftharpoons 2\ NO_{3(g)}$ with $K_c = 240$

The following concentrations were found at equilibrium: $[O_2] = 0.12\ M$ and $[NO_3] = 8.0\ M$. What will be the $[NO_2]$ in this set of equilibrium conditions?

$$240 = \frac{(8)^2}{[NO_2]^2(0.12)}$$

$$[NO_2]^2 = 2.\bar{2}2\ \underline{M}^2$$

$$[NO_2] = 1.\bar{4}9\ \underline{M} = 1.5\ \underline{M}$$

10. (10 points) The reaction below was studied at 350 K. A mixture of gases initially contained the reactants at the following concentrations: $[H_2] = [CO] = 0.800\ M$ with no products present. At equilibrium, $[CH_4]$ was found to be $0.150\ M$. What is the value of the equilibrium constant, K_C, for this reaction?

$3\ H_{2(g)}$	+	$CO_{(g)}$	⇔	$CH_{4(g)}$	+	$H_2O_{(g)}$	
0.8		0.8		∅		∅	i
−0.450		−0.150		+0.150		+0.150	c
0.350		0.650		0.150		0.150	e

$$K_c = \frac{(0.150)^2}{(0.350)^3(0.650)}$$

$$K_c = 0.80\bar{7}4 = 0.807$$

(M^{-2} if units given)

11. (20 points) Consider the gas-phase reaction: **2 HOF → 2 HF + O_2**

Time (min)	[HOF] (*M*)	ln [HOF]	1/[HOF] (M^{-1})
0	0.850	– 0.163	1.18
2.0	0.810	– 0.211	1.23
5.0	0.754	– 0.282	1.33
20.0	0.526	– 0.642	1.90
50.0	0.243	– 1.415	4.12

rate = k[HOF]

ln[HOF] = −kt + ln[HOF]$_0$

The data to the left were collected for this reaction at 25°C.

IMPORTANT information: A plot of the ln [HOF] vs. time produced a straight line.

→ 1st order

A) Calculate k for the reaction (including units)

$-1.415 = -k(50.0\,min) - 0.163$

4

$k = 0.02504\,min^{-1} = 0.0250\,min^{-1}$

B) What is the half-life for the reaction?

3

$t_{1/2} = \frac{\ln 2}{0.02504\,min^{-1}} = 27.68\,min = 27.7\,min$

C) For the trial described by the data, what is the total time required for 90.0% of the [HOF] to react?

90% reacted = 10% remains

$\ln(0.100) = -(0.02504\,min^{-1})t$

5

$t = 91.96\,min = 92.0\,min$

D) For the trial described by the data, what will be the rate of the reaction (-Δ[HOF]/Δt) at 15.0 min?

8

$\ln[HOF] = -(0.2504\,min^{-1})(15.0\,min) + \ln(0.850)$

$[HOF] = e^{(-0.53817)} = 0.5838\,M$

$rate = (0.02504\,min^{-1})(0.5838\,M)$

$rate = 0.01462\,M/min$

$= 0.0146\,M \cdot min^{-1}$

12. (20 points) Consider the ***proposed*** reaction mechanism for phosphorus monoxide and hydrogen gas:

Proposed mechanism:

$$2\,PO \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} P_2O_2 \qquad \textbf{rxn 1: fast, equilibrium}$$

$$P_2O_2 + H_2 \xrightarrow{k_2} P_2O + H_2O \qquad \textbf{rxn 2: slow}$$

$$P_2O + H_2 \xrightarrow{k_3} P_2 + H_2O \qquad \textbf{rxn 3: fast}$$

A) Write the overall reaction:

2 $2\,PO + 2\,H_2 \longrightarrow P_2 + 2\,H_2O$

2 B) What is the role of **PO** in the reaction? reactant

2 C) What is the role of $\mathbf{P_2O_2}$ in the reaction mechanism? intermediate

D) Write the differential rate law expression for the overall reaction that is consistent with the mechanism, excluding any intermediates.

$\text{rate} = k_2[H_2][P_2O_2]$ (from slow step)

$k_1[PO]^2 = k_{-1}[P_2O_2]$ (from step 1 - eq)

6 $[P_2O_2] = \frac{k_1}{k_{-1}}[PO]^2$

$\text{rate} = k_{overall}[PO]^2[H_2]$

E) The following initial rates data was collected for the reaction.
Based on these data, what is the experimental differential rate law equation for the reaction?
You are NOT required to calculate a value for k.

Expt.	[PO] (mol/L) x	[H_2] (mol/L) y	rate (mol/L•s)
1	0.25	0.10	3.0×10^{-4}
2	0.75 (×3)	0.10	2.7×10^{-3} (×9)
3	0.25	0.20 (×2)	6.0×10^{-4} (×2)

$3^x = 9$

$x = 2$

$2^y = 2$

$y = 1$

$\text{rate} = k[PO]^2[H_2]$

6

F) Is the mechanism (as it is written) consistent with the observed rate law? ***Explain briefly.***

Yes, the rate laws match.

2