PRACTICE EXAM #1 (Ch. 14-15)

Chem1B, General Chemistry II

1.) Peroxodisulfate ($S_2O_8^{2-}$) and iodide (I-) react according to the following **balanced** equation. The data below was collected measuring the rate of the disappearance of $S_2O_8^{2-}$ at 298 K.

$$S_2O_8^{2-}(aq) + 2\ I^-(aq) \rightarrow 2\ SO_4^{2-}(aq) + I_2(aq)$$

Experiment	$[S_2O_8^{2-}]$, initial	$[I^-]$, initial	Rate, M/s
1	1.0×10^{-4} M	0.010	1.09
2	2.1×10^{-4} M	0.010	2.18
3	2.1×10^{-4} M	0.0049	1.11

a.) Derive the complete **rate law** for the above reaction.

b.) Calculate the **average rate constant** for all three trials, including units. Express it to three significant figures.

c.) Calculate the rate constant at 315 K, given that the activation energy for this reaction is 51.8 kJ/mol.

2.) Calculate the half life ($t_{1/2}$) for a second-order reaction that has a rate constant of 3.18×10^{-4} $M^{-2}\ s^{-1}$ and an initial concentration of 0.0561 M.

3.) Consider the following equilibrium.

$CO(g) + 3\ H_2(g) \leftrightharpoons CH_4(g) + H_2O(g)$ $\Delta H = -230$ kJ/mol, $K_c = 190$ at 1000 K

a.) At 1000 K, does the reaction favor reactants or products? Explain.

b.) In which direction will this reaction shift (to reactants or to products) at 200 K? Explain.

c.) In which direction will this reaction shift (to reactants or products) if the volume of the container is *lowered*? Explain.

d.) Calculate K_p for the above equilibrium.

4.) A 1.00 mole sample of $NaHCO_3(s)$ is put in an evacuated (empty) 2.50 L flask at 373.15 K. Calculate the partial pressures of $CO_2(g)$ and $H_2O(g)$, in atm.

$2\ NaHCO_3(s) \leftrightharpoons Na_2CO_3(s) + CO_2(g) + H_2O(g)$ $K_p = 0.23$

5.) Consider the following equilibrium.

$$Ag^{+}(aq) + Fe^{2+}(aq) \leftrightharpoons Ag(s) + Fe^{3+}(aq) \qquad K_c = 2.98$$

Initially, the solutions are prepared with excess Ag(s), $[Ag^{+}] = [Fe^{2+}] = 0.200$ M, and $Fe^{3+}] = 0.350$ M. Calculate the concentrations of **each** ion at equilibrium.

6.) The following reaction has an experimentally-derived rate law as given.

$$2\ NO + 2\ H_2 \rightarrow N_2 + 2\ H_2O \qquad \text{rate} = k[NO]^2[H_2]$$

Show that this is consistent with the following three-step mechanism.

(1) $2\ NO \leftrightharpoons N_2O_2$ (forward k_1; backward k_2), *fast*
(2) $N_2O_2 + H_2 \rightarrow N_2O + H_2O$ (k_3), *slow*
(3) $N_2O + H_2 \rightarrow N_2 + H_2O$ (k_4), *fast*