1. (10 pts) Four identical 1.0-liter flasks contain the gases, He, CH₄, Cl₂, and NH₃, each at 1 atm, and 25°C. Circle one answer for each of the following questions.

a) For which gas do the molecules have the highest average velocity?
   \[ \text{He} \quad \text{CH₄} \quad \text{Cl₂} \quad \text{NH₃} \quad \text{all gases the same} \]

b) For which gas do the molecules have the smallest average kinetic energy?
   \[ \text{He} \quad \text{CH₄} \quad \text{Cl₂} \quad \text{NH₃} \quad \text{all gases the same} \]

c) Which gas sample has the greatest number of molecules?
   \[ \text{He} \quad \text{CH₄} \quad \text{Cl₂} \quad \text{NH₃} \quad \text{all gases the same} \]

d) Which gas sample has the greatest frequency of collisions of gas molecules with each other?
   \[ \text{He} \quad \text{CH₄} \quad \text{Cl₂} \quad \text{NH₃} \quad \text{all gases the same} \]

e) Which gas has the highest density?
   \[ \text{He} \quad \text{CH₄} \quad \text{Cl₂} \quad \text{NH₃} \quad \text{all gases the same} \]

2. (6 pts) Predict how the equilibrium will shift by the affect of compression (a decrease in volume) for each of the following systems. Circle the correct answer for each system.

a) \( \text{CaCl}_2 (s) + 2 \text{H}_2\text{O} (g) \rightleftharpoons \text{CaCl}_2\cdot2\text{H}_2\text{O} (s) \) left \[ \right \] no shift

b) \( \text{N}_2 (g) + 3 \text{Cl}_2 (g) \rightleftharpoons 2 \text{NCl}_3 (g) \) left \[ \right \] no shift

3. (10 pts) The density of a diatomic gas is 1.696 g/L at STP. Identify the gas.

\[ \frac{1.696 \text{ g/L}}{22.4 \text{ L/mol}} = 0.076 \text{ g/mol} \]

\[ \frac{0.076 \text{ g/mol}}{3.8} = 0.02 \text{ g/mol} \Rightarrow F \]

F₂ or fluorine

4. (10 pts) Calculate the temperature at which the average kinetic energy of O₂ gas is twice that of He gas at 20°C.

\[ \langle KE \rangle = \frac{3}{2} kT \]

\[ T = 20°C + 273 = 293K \]

double T

\[ T = 586K \text{ or } 313°C \]
5. (6 pts) Write the equilibrium constant expression, $K_p$, for each of the following equilibria.

a) \[ \text{CaCl}_2 (s) + 2 \text{H}_2\text{O} (g) \leftrightharpoons \text{CaCl}_2 \cdot 2\text{H}_2\text{O} (s) \]
   \[ K_p = \left( \frac{1}{P_{\text{H}_2\text{O}}} \right)^2 \]

b) \[ 2 \text{NCl}_3 (g) \leftrightharpoons \text{N}_2 (g) + 3 \text{Cl}_2 (g) \]
   \[ K_p = \frac{P_{\text{N}_2} (P_{\text{Cl}_2})^3}{(P_{\text{NCl}_3})^2} \]

6. (8 pts) The equilibrium constant $K = 1.4 \times 10^{-3}$ at 25°C, for the following reaction.
   \[ 2 \text{HI} (g) \leftrightharpoons \text{H}_2 (g) + \text{I}_2 (g) \]

The initial concentrations are: $[\text{H}_2]_0 = 0.20 \text{ M}$, $[\text{I}_2]_0 = 0.25 \text{ M}$ and $[\text{HI}]_0 = 0.50 \text{ M}$. Calculate the reaction quotient $Q$ to determine the direction the reaction proceeds in order to reach equilibrium? Show your work.

\[
Q = \frac{[\text{H}_2][I_2]}{[\text{HI}]^2} = \frac{(0.2)(0.25)}{(0.5)^2} = 0.2
\]

$Q > K$ reaction proceeds left

7. (10 pts) Consider the following equilibrium.

\[ 2 \text{NOCl} (g) \leftrightharpoons 2 \text{NO} (g) + \text{Cl}_2 (g) \]

The equilibrium constant, $K = 1.6 \times 10^{-5}$ at 400°C. NOCl and Cl$_2$ are placed in a 1-Liter container. The initial concentration of NOCl is 1.0 M and the initial concentration of Cl$_2$ is 1.0 M. Calculate the equilibrium concentration of NO.

\[ \begin{align*}
\text{Initial} & : & 1.0 \text{ M} & & 0 & & 1.0 \text{ M} \\
\text{Change} & : & -2x & & +2x & & +x \\
\text{at equilibrium} & : & 1-2x & & 2x & & 1+x
\end{align*} \]

\[
K = \frac{(2x)^2 (1+x)}{(1-2x)^2} = 1.6 \times 10^{-5}
\]

\[
i \ f \ x = \text{small} \\
i \ f \ 1 + x \approx 1 \\
i \ f \ 1 - 2x \approx 1
\]

\[
4x^2 = 1.6 \times 10^{-5}
\]

\[
x = 0.002
\]

\[
[\text{NO}] = 2x = 0.004 \text{ M}
\]
8. (10 pts) Consider the following three equilibria occurring simultaneously in solution.

\[ \text{Ca}^{2+} (aq) + \text{C}_2\text{O}_4^{2-} (aq) \rightleftharpoons \text{CaC}_2\text{O}_4 (s) \quad (1) \]
\[ \text{H}_2\text{C}_2\text{O}_4 (aq) \rightleftharpoons \text{H}^+ (aq) + \text{HC}_2\text{O}_4^- (aq) \quad (2) \]
\[ \text{HC}_2\text{O}_4^- (aq) \rightleftharpoons \text{H}^+ (aq) + \text{C}_2\text{O}_4^{2-} (aq) \quad (3) \]

a) If NaOH is added to the solution, write the net ionic equation for the reaction that occurs.

\[ \text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} \]

b) When NaOH is added to the solution, what is the affect on equilibrium (2). Does equilibrium (2) shift to the right, to the left or is there no shift? Circle the correct answer.

Left \[ \rightleftharpoons \] Right \[ \text{No shift} \]

c) When NaOH is added to the solution, what is the affect on equilibrium (3). Does equilibrium (3) shift to the right, to the left or is there no shift? Circle the correct answer.

Left \[ \rightleftharpoons \] Right \[ \text{No shift} \]

d) If there is a shift in equilibrium (2) or (3), how does this affect equilibrium (1). Does equilibrium (1) shift to the right, to the left or is there no shift? Circle the correct answer.

Left \[ \rightleftharpoons \] Right \[ \text{No shift} \]

e) Will the amount of \text{CaC}_2\text{O}_4 (s) precipitate increase or decrease? Circle the correct answer.

Increase \[ \] Decrease

9. (10 pts) For the thermal decomposition of \text{CaCO}_3, the equilibrium constant \( K_p = 1.26 \) at 800°C.

\[ \text{CaCO}_3 (s) \rightleftharpoons \text{CaO} (s) + \text{CO}_2 (g) \]

If a 20.0-gram sample of \text{CaCO}_3 is put into a 10.0 liter container and heated to 800°C, what percent of the \text{CaCO}_3 will react to reach equilibrium.

\[ K_p = P_{\text{CO}_2} = 1.26 \text{ atm} \]
\[ V = 10.0 \text{ L} \]
\[ T = 800^\circ \text{C} + 273 = 1073 \text{ K} \]
\[ n = \frac{PV}{RT} = \frac{(1.26 \text{ atm})(10.0 \text{ L})}{(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(1073 \text{ K})} \]
\[ n = 0.143 \text{ mol} \text{ CO}_2 \]
\[ 1 \text{ mol CaCO}_3 \]

\[ 0.143 \text{ mol CaCO}_3 \left( \frac{100 \text{ g}}{\text{ mol}} \right) = 14.3 \text{ g } \text{CaCO}_3 \text{ reacts} \]

\[ \frac{14.3}{20} \times 100 = \boxed{71.5 \%} \]
10. (10 pts) A mixture of KCl and KClO₃ weighing 1.2 grams was heated. After the reaction goes to completion, the O₂ gas generated occupied 130 mL at STP. Calculate the percent of the original mixture that was KClO₃. KClO₃ decomposes according to the following reaction.

\[
2 \text{KClO}_3(s) \rightarrow 2 \text{KCl}(s) + 3 \text{O}_2(g)
\]

\[
0.13 \text{ L O}_2 \left(\frac{1 \text{ mol}}{22.4 \text{ L}}\right) = 0.0058 \text{ mol O}_2 \quad \frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2}
\]

\[
0.00357 \text{ mol KClO}_3 \quad \frac{122.559 \text{ g}}{\text{mol}} = 0.4749
\]

\[
\frac{0.4749}{1.29} \times 100 = 39.5\%
\]

11. (10 pts) The oxidation of nitric oxide to nitrogen dioxide occurs according to the following reaction.

\[
2 \text{NO(g)} + \text{O}_2(g) \rightarrow 2 \text{NO}_2(g)
\]

200.0 mL of NO at STP reacts with 300.0 mL of O₂ at STP. The reaction goes to completion. Calculate the partial pressure of NO₂ in the final mixture if the total pressure is 1 atm.

\[
0.2 \text{ L NO} \left(\frac{1 \text{ mol}}{22.4 \text{ L}}\right) = 0.00893 \text{ mol NO} \quad \frac{2 \text{ mol NO}_2}{2 \text{ mol NO}} = 0.00893 \text{ mol NO}_2
\]

\[
0.3 \text{ L O}_2 \left(\frac{1 \text{ mol}}{22.4 \text{ L}}\right) = 0.0134 \text{ mol O}_2 \quad \frac{2 \text{ mol NO}_2}{1 \text{ mol O}_2} = 0.0268 \text{ mol NO}_2
\]

\[
0.00893 \text{ mol NO} \quad \frac{1 \text{ mol O}_2}{2 \text{ mol NO}} = 0.00447 \text{ mol O}_2 \quad \text{reacts}
\]

\[
0.0134 \text{ mol O}_2 - 0.00447 = 0.00893 \text{ mol O}_2 \quad \text{remain}
\]

\[
P_{\text{NO}_2} = x_{\text{NO}_2} P_{\text{tot}} = \left(\frac{0.00893}{0.00893 + 0.00893}\right) 1\text{ atm}
\]

\[
P_{\text{NO}_2} = 0.5 \text{ atm}
\]