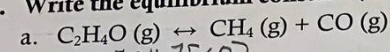


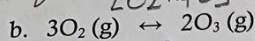
Equilibrium Constant (K_{eq}) - Chemistry

Name: Key Pd: _____

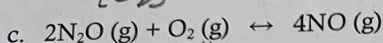
1. Write the equilibrium constant (K_{eq}) expressions for the following homogeneous equilibria.



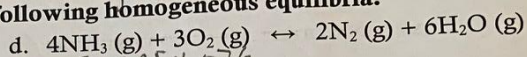
$$K_{eq} = \frac{[CH_4][CO]}{[C_2H_4O]}$$



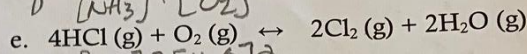
$$K_{eq} = \frac{[O_3]^2}{[O_2]^3}$$



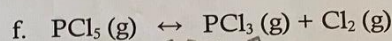
$$K_{eq} = \frac{[NO]^4}{[N_2O]^2 [O_2]}$$



$$K_{eq} = \frac{[N_2]^2 [H_2O]^6}{[NH_3]^4 [O_2]^3}$$

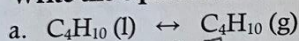


$$K_{eq} = \frac{[Cl_2]^2 [H_2O]^2}{[HCl]^4 [O_2]}$$

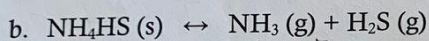


$$K_{eq} = \frac{[PCl_3][Cl_2]}{[PCl_5]}$$

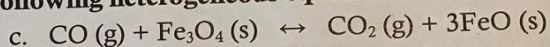
2. Write the equilibrium constant (K_{eq}) expressions for the following heterogeneous equilibria.



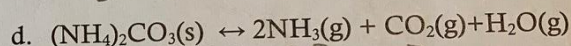
$$K_{eq} = [C_4H_{10}]$$



$$K_{eq} = [NH_3][H_2S]$$



$$K_{eq} = \frac{[CO_2]}{[CO]}$$



$$K_{eq} = [NH_3]^2 [CO_2][H_2O]$$

For the following problems, show all of your work including set-up (with K_{eq} expression) and answer with units if needed.

3. At 773 K, the reaction $2NO(g) + O_2(g) \leftrightarrow 2NO_2(g)$ produces the following concentrations: $[NO] = 3.49 \times 10^{-4} M$; $[O_2] = 0.80 M$; $[NO_2] = 0.25 M$. Calculate the equilibrium constant (K_{eq}) for this reaction.

$$K_{eq} = \frac{[NO_2]^2}{[NO]^2 [O_2]} \quad K_{eq} = \frac{(0.25)^2}{(3.49 \times 10^{-4})^2 (0.80)} = \boxed{6.41 \times 10^5}$$

products favored b/c $K_{eq} > 1$

4. The chemical equation for the decomposition of formamide is: $HCONH_2(g) \leftrightarrow NH_3(g) + CO(g)$

Calculate K_{eq} using the following equilibrium data: $[HCONH_2] = 0.0637 M$, $[NH_3] = 0.518 M$ and $[CO] = 0.518 M$.

$$K_{eq} = \frac{[NH_3][CO]}{[HCONH_2]} \quad K_{eq} = \frac{0.518 \cdot 0.518}{0.0637} = \boxed{4.21}$$

products favored b/c $K_{eq} > 1$

5. Calculate K_{eq} for the reaction for iron and water if the equilibrium concentrations are as follows: $[H_2O] = 1.00 M$ & $[H_2] = 4.50 M$. $2Fe(s) + 3H_2O(g) \leftrightarrow Fe_2O_3(s) + 3H_2(g)$

$$K_{eq} = \frac{[H_2]^3}{[H_2O]^3} \quad K_{eq} = \frac{(4.5)^3}{(1.00)^3} = \boxed{91.1}$$

products favored b/c $K_{eq} > 1$

6. At 793 K, the equilibrium constant for the reaction $\text{NCl}_3(\text{g}) + \text{Cl}_2(\text{g}) \leftrightarrow \text{NCl}_5(\text{g})$ is 39.3.
- a. Do the products or the reactants dominate in this equilibrium? Explain your answer in complete sentences.

products because the $K_{eq} = 39.3$ and that is greater than 1

- b. If the equilibrium constant for this reaction were less than 1, would the reactants or products be dominant? Explain your answer in complete sentences.

Reactants b/c K_{eq} is less than 1

7. The equilibrium constant is 9.36 for the following reaction: $\text{A}(\text{g}) + 3\text{B}(\text{g}) \leftrightarrow 2\text{C}(\text{g})$. The table below provides concentration data for two different reaction mixtures of these gases.

	A (mol/L)	B (mol/L)	C (mol/L)
Mixture 1	0.716	0.208	0.425
Mixture 2	0.562	0.491	0.789

- a. Calculate the K_{eq} for each mixture. Use the back of the sheet to show your work.

$$K_{eq} = \frac{[\text{C}]^2}{[\text{A}][\text{B}]^3}$$

$$\text{Mixture 1: } K_{eq} = \frac{(0.425)^2}{(0.716)(0.208)^3}$$

$$\text{Mixture 2: } K_{eq} = \frac{(0.789)^2}{(0.562)(0.491)^3}$$

$$K_{eq} = 28.0$$

$$K_{eq} = 9.36$$

- b. Are both reactions at equilibrium? Explain your answer in complete sentences.

No, only mixture 2 is at equilibrium because it has a K_{eq} value of 9.36.