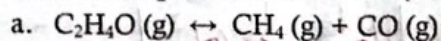


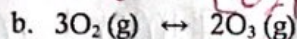
Equilibrium Constant ( $K_{eq}$ ) - Chemistry

Name: KEY Pd: \_\_\_\_\_

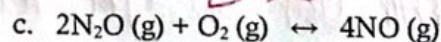
1. Write the equilibrium constant ( $K_{eq}$ ) expressions for the following homogeneous equilibria.



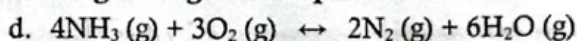
$$K_{eq} = \frac{[CH_4][CO]}{[C_2H_4O]}$$



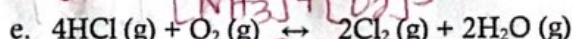
$$K_{eq} = \frac{[O_3]^2}{[O_2]^3}$$



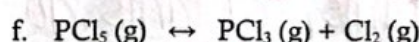
$$K_{eq} = \frac{[NO]^4}{[N_2O]^2 [O_2]}$$



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

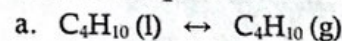


$$K_{eq} = \frac{[Cl_2]^2 [H_2O]^2}{[HCl]^4 [O_2]}$$

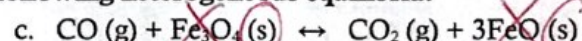


$$K_{eq} = \frac{[PCl_3][Cl_2]}{[PCl_5]}$$

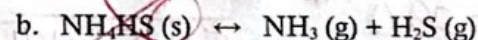
2. Write the equilibrium constant ( $K_{eq}$ ) expressions for the following heterogeneous equilibria.



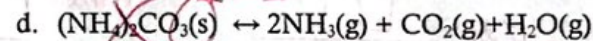
$$K_{eq} = [C_4H_{10}]$$



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



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



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

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.

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

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  $[CO] = 0.518 M$ .

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

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

$$K_{eq} = \frac{[H_2]^3}{[H_2O]^3} = \frac{[4.50]^3}{[1.00]^3} = 91.1$$

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.

$K_{eq} = \frac{[\text{NCl}_5]}{[\text{NCl}_3][\text{Cl}_2]} = 39.3$  products,  $K_{eq} > 1$  is favored by the products, products are numerator

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, reactants go on the denominator so greater denominator gives a #  $< 1$

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.

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

mix #1  $\frac{[.425]^2}{[.716][.208]^3} = 29.0$       mix #2  $\frac{[.789]^2}{[.562][.491]^3} = 9.36$

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

No, only mixture #2 b/c the calculated  $K_{eq}$  is equal to the  $K_{eq}$

$K_{eq} = \frac{[\text{C}]^2}{[\text{A}][\text{B}]^3} = \frac{[.789]^2}{[.562][.491]^3} = 9.36$

$K_{eq} = \frac{[\text{C}]^2}{[\text{A}][\text{B}]^3} = \frac{[.425]^2}{[.716][.208]^3} = 29.0$

$K_{eq} = \frac{[\text{C}]^2}{[\text{A}][\text{B}]^3} = \frac{[.789]^2}{[.562][.491]^3} = 9.36$