

Name: \_\_\_\_\_

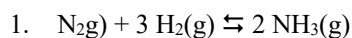
Date: \_\_\_\_\_

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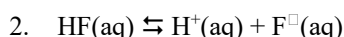
Seat #: \_\_\_\_\_

For the following three reactions,

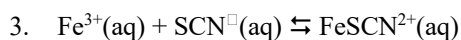
- write the  $K_{eq}$  expression in terms of concentration,  $K_c$ .
- given the equilibrium concentrations, state whether each equilibrium is product-favored, reactant-favored, or fairly even ( $[products] \approx [reactants]$ ).
- calculate the value of  $K_c$ .



At equilibrium:  $[N_2] = 1.50 \text{ M}$   
 $[H_2] = 2.00 \text{ M}$   
 $[NH_3] = 0.01 \text{ M}$



At equilibrium:  $[HF] = 0.55 \text{ M}$   
 $[H^+] = 0.001 \text{ M}$   
 $[F^-] = 0.001 \text{ M}$



At equilibrium:  $[Fe^{3+}] = 0.55 \text{ M}$   
 $[SCN^-] = 0.001 \text{ M}$   
 $[FeSCN^{2+}] = 0.001 \text{ M}$

**Summarize:**

Fill in the blanks with product-favored, reactant-favored, and approximately equal

$K_c$	state of equilibrium
$K_c \gg 1$	
$K_c \ll 1$	
$K_c \approx 1$	

4. Knowing that pure water has a density of 1g/1mL calculate the mass of 1.00 Liter of water.

Calculate the number of moles in 1.00 L of  $H_2O$ .

What is the concentration (M) of water in water?

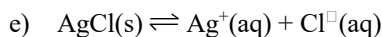
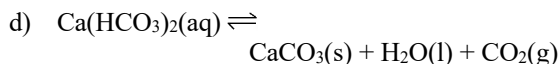
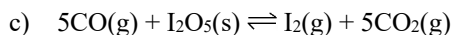
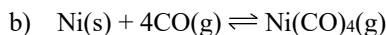
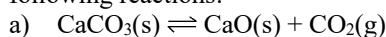
At this temperature, can you get more moles of water into this Liter of water?

The  $[H_2O]$  \_\_\_\_\_ (is / is not) constant.

**Important Note:**

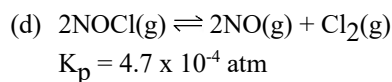
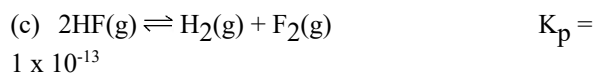
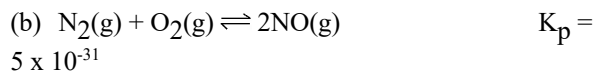
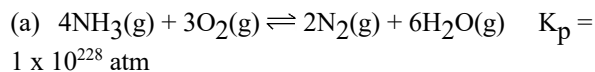
Since the concentrations of solids and liquids are constant, they are incorporated into the equilibrium constant,  $K_{eq}$ . That means, just leave them out of the  $K_c$  or  $K_p$  expression. Only include (g) and (aq)!

5. Write equilibrium expressions for each of the following reactions:



6. Write the equilibrium expression in terms of partial pressures ( $K_p$ ) for each of the following reactions. Rate the reactions in order of their increasing tendency to proceed toward completion:

— — — —



**A Question That You Should Be Able To Answer:**

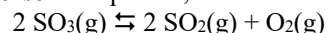
Why don't the  $K_p$ 's in (b) and (c) have units?

7. (a) Write the  $K_c$  expression for  
 $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{SO}_3(\text{g})$

Calculate the value of  $K_c$ :

At equilibrium:  $[\text{SO}_2] = 1.50 \text{ M}$   
 $[\text{O}_2] = 1.25 \text{ M}$   
 $[\text{SO}_3] = 3.50 \text{ M}$

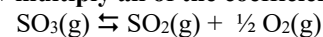
- (b) If we **reverse** the equation, it is:



Write the  $K_c$  expression for this equation and calculate the new value of  $K_c$ :

How does the expression and the value of  $K_c$  in 7(b) compare with those in 7(a)?

- (c) If we now **multiply all of the coefficients by 1/2**:



Write the  $K_c$  expression for this equation and calculate the new value of  $K_c$ :

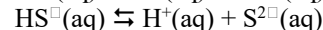
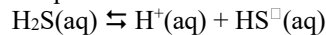
How do they compare with 7(b)?

- (d) What would happen to the  $K_c$  expression and its value if we **doubled** the coefficients?

**Summarize:**

Equation	$K_c$ expression & Value
doubled	
reversed	
halved	

8. Consider an equilibrium that occurs in two steps:



- (a) Write the overall reaction.  
 (b) How do the  $K_c$ 's for the two steps ( $K_{c1}$  &  $K_{c2}$ ) relate to the  $K_c$  of the overall reaction ( $K_c$ )?