

1. (2 Pts) Consider the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$. If hydrogen gas is added to this system at equilibrium, will the reaction shift towards *reactants* or *products*?

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2. (2 Pts) Consider the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$. The production of ammonia is an exothermic reaction. Will heating the equilibrium system *increase* or *decrease* the amount of ammonia produced?

←

3. (2 Pts) Consider the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$. If we use a catalyst, which way will the reaction shift?

It won't shift (only reaches equilibrium faster).

4. (3 Pts) Given the following data for the reaction: $A(g) + 2B(s) \rightleftharpoons AB_2(g)$

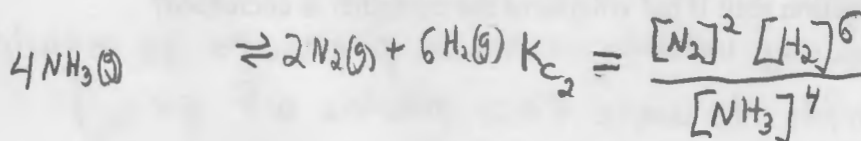
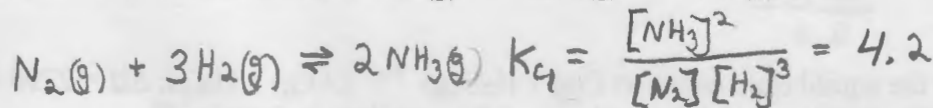
Temperature (K)	K_c
300	1.5×10^4
600	55
900	3.4×10^{-3}

$$K_c \approx \frac{\text{products}}{\text{reactants}}$$

Is the reaction *endothermic* or *exothermic* (explain your answer)?

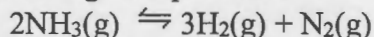
Raising the temperature gives less product, therefore it is an exothermic reaction.

5. (4 Pts) Consider the reaction, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$. $K_c = 4.2$ at 600 K. What is the value of K_c for $4NH_3(g) \rightleftharpoons 2N_2(g) + 6H_2(g)$



By inspection $K_{c2} = \frac{1}{K_{c1}^2} = \frac{1}{4.2^2} = 0.057$

6. (4 Pts) 5.00 mol of ammonia are introduced into a 5.00 L reactor vessel in which it partially dissociates at high temperatures.



At equilibrium at a particular temperature, 1.00 mole of ammonia remains. Calculate K_c for the reaction.

	$2\text{NH}_3(\text{g})$	\rightleftharpoons	$3\text{H}_2(\text{g})$	$+$	$\text{N}_2(\text{g})$	
I.	$\frac{5}{5} = 1.00 \text{ M}$		0		0	$K_c = \frac{[\text{H}_2]^3 [\text{N}_2]}{[\text{NH}_3]^2}$
C.	$-2x$		$3x$		x	
E.	$1.00 - 2x$		$3x$		x	

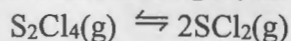
$$K_c = \frac{[1.2]^3 [0.4]}{[0.200]^2} = 17.3$$

@Eq $1.00 - 2x = \frac{1.00}{5.00} = 0.200$

$x = 0.400 \text{ M}$

So: E 0.200 1.2 0.4

7. (4 Pts) 4.21 moles of S_2Cl_4 are introduced into a 2.0 L vessel.



At equilibrium, 1.25 moles of S_2Cl_4 are found to remain in the container. Calculate K_c for this reaction.

	$\text{S}_2\text{Cl}_4(\text{g})$	\rightleftharpoons	$2\text{SCl}_2(\text{g})$	
I.	4.21 moles		0	$K_c = \frac{[\text{SCl}_2]^2}{[\text{S}_2\text{Cl}_4]}$
C.	$-x$		$+2x$	
E.	$4.21 - x = 1.25 \text{ moles}$		$2x$	

$$K_c = \frac{\left(\frac{5.92}{2.0}\right)^2}{\left(\frac{1.25}{2.0}\right)} = 14$$

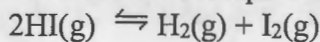
$x = 2.96$

E. $\frac{1.25}{2.0}$ $\frac{5.92}{2.0}$

8. (2 Pts) Consider the equilibrium equation $\text{C}(\text{s}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + \text{H}_2(\text{g})$, $\Delta H = 2296 \text{ J}$. Which way will the reaction shift if the volume of the container is decreased?

← (Decreasing volume increase pressure so equilibrium shifts toward less moles of gas.)

9. (2 Pts) What is the correct equilibrium constant expression for this reaction?



$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2}$$