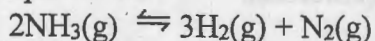
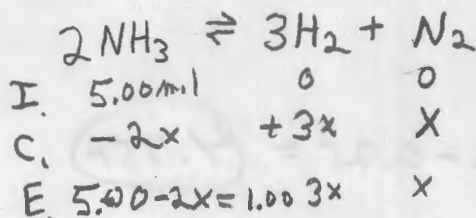


Show all work to receive credit.

1. (5 Pts) 5.00 mol of ammonia are introduced into a 5.00 L reactor vessel in which it partially dissociates at high temperatures.



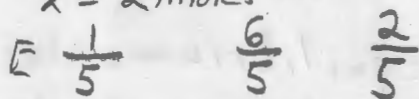
At equilibrium at a particular temperature, 1.00 mole of ammonia remains. Calculate  $K_c$  for the reaction.



$$K_c = \frac{[\text{H}_2]^3 [\text{N}_2]}{[\text{NH}_3]^2} = \frac{\left(\frac{6}{5}\right)^3 \left(\frac{1}{5}\right)}{\left(\frac{1}{5}\right)^2}$$

$$K_c = 17.3$$

$$x = 2 \text{ moles}$$



2. (3 Pts) Given the following data for the reaction:  $\text{A}(\text{g}) + 2\text{B}(\text{s}) \rightleftharpoons \text{AB}_2(\text{g}) + \text{heat}$

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

Is the reaction *endothermic* or *exothermic*?

Explain your reasoning:

**Exothermic**

As temperature increased, the value of  $K_c$  decreased, meaning less products were formed.

3. (4 Pts) Consider the reaction,  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ .  $K_c = 8.1 \times 10^{-3}$  at 900 K.

What is the value of  $K_c$  for  $\text{NH}_3(\text{g}) \rightleftharpoons \frac{1}{2}\text{N}_2(\text{g}) + \frac{3}{2}\text{H}_2(\text{g})$

$$K_{c1} = \frac{[\text{NH}_3]^2}{[\text{N}_2] [\text{H}_2]^3}$$

$$K_{c2} = \frac{[\text{N}_2]^{1/2} [\text{H}_2]^{3/2}}{[\text{NH}_3]} = \sqrt{\frac{1}{K_{c1}}} = \sqrt{\frac{1}{8.1 \times 10^{-3}}} = 11$$

4. (2 Pts) Consider the reaction  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ . If nitrogen is removed from the system at equilibrium, what will happen to the hydrogen ( $\text{H}_2$ ) concentration?

$[\text{H}_2]$  will increase.

5. (2 Pts) The data below refer to the following reaction:

$$2\text{NO}(\text{g}) + \text{Br}_2(\text{g}) \rightleftharpoons 2\text{NOBr}(\text{g})$$

Concentration (M)	[NO]	[Br <sub>2</sub> ]	[NOBr]
Initial	2.5	5.0	1.0
Equilibrium	2.0		

Find the concentration of Br<sub>2</sub> when the system reaches equilibrium.

$$2\text{NO} + \text{Br}_2 \rightleftharpoons 2\text{NOBr}$$

I	2.5	5.0	1.0
C	-2x	-x	+2x
E	2.0	5.0 - x	1.0 + 2x

So  $2x = 0.5$  and  $x = 0.25$  So For Br<sub>2</sub>  $5.0 - 0.25 = 4.75\text{M}$

6. (2 Pts) Consider the reaction  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ . If we use a catalyst, which way will the reaction shift?

A catalyst does not affect equilibrium values.  
It only moves the system to equilibrium faster.

7. (5 Pts) The data below refer to the following reaction: **Refer to question 5.**

$$2\text{NO}(\text{g}) + \text{Br}_2(\text{g}) \rightleftharpoons 2\text{NOBr}(\text{g})$$

Concentration (M)	[NO]	[Br <sub>2</sub> ]	[NOBr]
Initial	2.5	5.0	1.0
Equilibrium	2.0	4.75	1 + 2x = 1.5

Calculate K<sub>c</sub>.

$$K_c = \frac{[\text{NOBr}]^2}{[\text{NO}]^2 [\text{Br}_2]} = \frac{[1.5]^2}{[2.0]^2 [4.75]} = 0.12$$

8. (2 Pts) Consider the reaction  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) + \text{heat}$ . The production of ammonia is an exothermic reaction. Will heating the equilibrium system *increase* or *decrease* the amount of ammonia produced?

←  
decrease since heat is a product