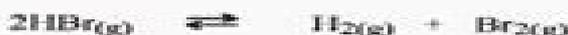


7. At a high temperature, 0.50 mol of HBr was placed in a 1.0 L container and allowed to decompose according to the reaction:



At equilibrium the  $[\text{Br}_2]$  was measured to be 0.13 M. What is  $K_{eq}$  for this reaction at this temperature?

$$2\text{HBr} \rightleftharpoons \text{H}_2 + \text{Br}_2$$

[I]	0.50	0	0
[C]	-0.26	+0.13	+0.13
[E]	0.24	0.13	0.13

$$K_{eq} = \frac{[\text{H}_2][\text{Br}_2]}{[\text{HBr}]^2} = \frac{(0.13)^2}{(0.24)^2} = 0.29$$

Answer  $K_{eq} = 0.29$

8. When 1.0 mol of  $\text{NH}_3(g)$  and 0.40 mol of  $\text{N}_2(g)$  are placed in a 5.0 L vessel and allowed to reach equilibrium at a certain temperature, it is found that 0.78 mol of  $\text{NH}_3$  is present. The reaction is:



initial  $[\text{NH}_3] = \frac{1.0 \text{ mol}}{5.0 \text{ L}} = 0.20 \text{ M}$       initial  $[\text{N}_2] = \frac{0.40 \text{ mol}}{5.0 \text{ L}} = 0.080 \text{ M}$   
 Equil<sup>m</sup>  $[\text{NH}_3] = \frac{0.78 \text{ mol}}{5.0 \text{ L}} = 0.156 \text{ M}$

$$2\text{NH}_3 \rightleftharpoons 3\text{H}_2 + \text{N}_2$$

[I]	0.20	0	0.080
[C]	-0.044	+0.066	+0.022
[E]	0.156	0.066	0.102

- a) Calculate the equilibrium concentrations of all three species.

$[\text{NH}_3] = 0.16 \text{ M}$      $[\text{H}_2] = 0.066 \text{ M}$      $[\text{N}_2] = 0.10 \text{ M}$

- b) Calculate the value of the equilibrium constant at this temperature.

*(use unrounded concs, then round to 2 sig)*

$$K_{eq} = \frac{[\text{H}_2]^3 [\text{N}_2]}{[\text{NH}_3]^2} = \frac{(0.066)^3 (0.102)}{(0.156)^2} = 0.0012$$

Answer  $1.2 \times 10^{-3}$

- c) How many moles of  $\text{H}_2$  are present at equilibrium?

$$0.066 \text{ M} \times 5.0 \text{ L} = 0.33 \text{ mol}$$

Answer  $0.33 \text{ mol}$

- d) How many moles of  $\text{N}_2$  are present at equilibrium?

$$0.102 \text{ M} \times 5.0 \text{ L} = 0.51 \text{ mol}$$

Answer  $0.51 \text{ mol}$

# Modeling Chemistry Unit 3 1 Answers

**Rachel S Tattersall**



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