

# Exercise 18.1b(H)

## Equilibrium & ICE Tables

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

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

Organizing algebraic relationships when solving equilibrium problems can be done using an “ICE” table.

### Initial Quantity

- Express initial concentrations in molarity (mol/L), or for gas equilibria express all pressures in a single pressure unit.
- If quantities of products are given, calculate the reaction quotient ( $Q$ ) to find the direction the reaction proceeds.

### Change in Quantity

- Unknown changes in quantity are represented by multiples of the variable ‘ $x$ ’ and should reflect the stoichiometric relationships between the reactant and products.
- If the reaction is moving forward, reaction changes are negative and product changes are positive. If the reaction is moving in reverse, it is the opposite.

### Equilibrium Quantity

- Find the equilibrium concentrations by adding/subtracting the change to/from the initial concentration.
- Because liquids and solids do not appear in the equilibrium expression, it is not necessary to find their equilibrium quantity.

**Example:** Given the equation,  $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2 \text{HI}(\text{g})$ , calculate the equilibrium concentration of each species if the initial concentration of  $\text{H}_2 = \text{I}_2 = 0.200 \text{ M}$  and  $K_{\text{eq}} = 64$ .

	$\text{H}_2(\text{g})$	+	$\text{I}_2(\text{g})$	$\rightleftharpoons$	$2 \text{HI}(\text{g})$
Initial	0.200 M		0.200 M		0
Change	- x		- x		+ 2x
Equilibrium	0.200 - x		0.200 - x		2x

$$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} \rightarrow K_{\text{eq}} = \frac{[2x]^2}{[0.200 - x]^2} \rightarrow 64 = \frac{[2x]^2}{[0.200 - x]^2} \quad \begin{array}{l} \text{taking the} \\ \text{square of both} \\ \text{sides to solve} \end{array} \quad 8 = \frac{2x}{0.200 - x} \quad \begin{array}{l} \text{solving for } x \text{ gives} \\ x = 0.16 \end{array}$$

$$[\text{H}_2] = 0.200 - 0.16 = \boxed{0.04 \text{ M}}, [\text{I}_2] = 0.200 - 0.16 = \boxed{0.04 \text{ M}}, \text{ and } [\text{HI}] = 2(0.16) = \boxed{0.32 \text{ M}}$$

**These concentrations may be checked by substituting them back into the equilibrium expression.**

**DIRECTIONS:** Answer the following in the space provided.

- For the reaction  $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + 3\text{H}_2(\text{g})$ , the equilibrium concentrations are:  $[\text{CO}] = 0.300\text{M}$ ,  $[\text{H}_2] = 0.800\text{M}$  and  $[\text{CH}_4] = 0.400\text{M}$ . Determine the equilibrium concentration of  $\text{H}_2\text{O}$  in this mixture if  $K_{\text{eq}} = 5.67$ . (No ICE table req'd.)
- The value of  $K_{\text{eq}}$  for the equilibrium  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$  is  $4.64 \times 10^{-3}$  at  $25^\circ\text{C}$ . If the initial concentrations of  $\text{N}_2\text{O}_4$  is  $0.0367\text{M}$  and the initial concentration of  $\text{NO}_2$  is zero, what will be the concentration of both gases at equilibrium?
- Given the equation,  $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2 \text{HI}(\text{g})$ , calculate the equilibrium concentration of each species if the initial concentration of  $\text{H}_2 = \text{I}_2 = 0.100 \text{ M}$  and  $K_{\text{eq}} = 64$ .

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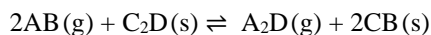
4. For the reaction  $\text{CO(g)} + \text{H}_2\text{O(g)} \rightleftharpoons \text{CO}_2\text{(g)} + \text{H}_2\text{(g)}$ ,  $K_{\text{eq}} = 5.10$ . Calculate the equilibrium concentrations of each species if 1.000 mol of each reactant is added to a 1.000L container.

5. For the reaction  $\text{H}_2\text{(g)} + \text{F}_2\text{(g)} \rightleftharpoons 2\text{HF(g)}$ ,  $K_{\text{eq}} = 1.15 \times 10^2$ . Calculate the equilibrium concentration of each species if 3.000 mol of each reactant was added to a 1.500L container.

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

6. 0.200mol of  $\text{H}_2$ , 0.200mol of  $\text{I}_2$ , and 0.200mol of  $\text{HI}$  were placed in a 1.00L flask and allowed to come to equilibrium. The  $K_{\text{eq}}$  value of the reaction at this temperature is 49.5. Determine the equilibrium concentrations of all species.

7. Find the equilibrium constant,  $K_{\text{eq}}$ , for the following equilibrium. The initial concentrations of  $\text{AB}$  and  $\text{A}_2\text{D}$  are 0.30 M before they are mixed and when equilibrium is reached the equilibrium concentration of  $\text{A}_2\text{D}$  is 0.20 M. Be sure to show an ICE table for your calculation.



8. At  $650^\circ\text{C}$ , the reaction below has a  $K_{\text{eq}}$  value of 0.771. If 2.00 mol of both hydrogen and carbon dioxide are placed in a 4.00 L container and allowed to react, what will be the equilibrium concentrations of all four gases?

