Chemistry H

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, $H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g)$, calculate the equilibrium concentration of each species if the initial concentration of $H_2 = I_2 = 0.200\ M$ and $K_{eq} = 64$.

	$H_2(g)$	+	$I_2(g)$	$\rightleftharpoons$	2 HI(g)
Initial	0.200 *M*		0.200 *M*		0
Change	− x		− x		+ 2x
Equilibrium	0.200 − x		0.200 − x		2x

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]} \rightarrow K_{eq} = \frac{[2x]^2}{[0.200 - x]^2} \rightarrow 64 = \frac{[2x]^2}{[0.200 - x]^2}$$ *taking the square of both sides to solve* $$8 = \frac{2x}{0.200 - x}$$ *solving for x gives x = 0.16*

$[H_2] = 0.200 - 0.16 =$ **0.04 *M***, $[I_2] = 0.200 - 0.16 =$ **0.04 *M***, and $[HI] = 2(0.16) =$ **0.32 *M***

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

DIRECTIONS: Answer the following in the space provided.

1. For the reaction $CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$, the equilibrium concentrations are: [CO] = 0.300*M*, $[H_2]$ = 0.800*M* and $[CH_4]$ = 0.400*M*. Determine the equilibrium concentration of H_2O in this mixture if $K_{eq} = 5.67$. *(No ICE table req'd.)*

2. The value of K_{eq} for the equilibrium $N_2O_4\ (g) \rightleftharpoons 2NO_2\ (g)$ is 4.64×10^{-3} at 25°C. If the initial concentrations of N_2O_4 is 0.0367M and the initial concentration of NO_2 is zero, what will be the concentration of both gases at equilibrium?

3. Given the equation, $H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g)$, calculate the equilibrium concentration of each species if the initial concentration of $H_2 = I_2 = 0.100\ M$ and $K_{eq} = 64$.

Answers: 1) $[H_2O] = 0.0677\ M$ 2) $[N_2O_4] = 0.0307\ M$, $[NO_2] = 0.0119\ M$ 3) $[H_2] = [I_2] = 0.020\ M$, [HI] = 0.16 0.0313 *M* 4) [CO] = $[H_2O] = 0.307\ M$, $[CO_2] = [H_2] = 0.693\ M$ 5) $[H_2] = 0.32\ M$, $[F_2] = 0.32\ M$, [HF] = 3.37 *M* 6) $[H_2]$ and $[I_2] = 0.066\ M$; [HI] = 0.467 *M* 7) 0.80 8) [CO] = $[H_2O] = 0.234\ M$ $[H_2] = [CO_2] = 0.266\ M$

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4. For the reaction $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$, $K_{eq} = 5.10$. 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 $H_2(g) + F_2(g) \rightleftharpoons 2HF(g)$, $K_{eq} = 1.15 \times 10^2$. 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 H_2, 0.200mol of I_2, and 0.200mol of HI were placed in a 1.00L flask and allowed to come to equilibrium. The K_{eq} value of the reaction at this temperature is 49.5. Determine the equilibrium concentrations of all species.

7. Find the equilibrium constant, K_{eq}, for the following equilibrium. The initial concentrations of AB and A_2D are 0.30 M before they are mixed and when equilibrium is reached the equilibrium concentration of A_2D is 0.20 M. Be sure to show an ICE table for your calculation.

$$2AB(g) + C_2D(s) \rightleftharpoons A_2D(g) + 2CB(s)$$

8. At 650°C, the reaction below has a K_{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?

$$H_2(g) + CO_2(g) \rightleftharpoons CO(g) + H_2O(g)$$

Answers: 1) $[H_2O] = 0.0677\ M$ 2) $[N_2O_4] = 0.0307\ M$, $[NO_2] = 0.0119\ M$ 3) $[H_2] = [I_2] = 0.020\ M$, [HI] = 0.16 0.0313 *M* 4) $[CO] = [H_2O] = 0.307\ M$, $[CO_2] = [H_2] = 0.693\ M$ 5) $[H_2] = 0.32\ M$, $[F_2] = 0.32\ M$, [HF] = 3.37 *M* 6) $[H_2]$ and $[I_2] = 0.066\ M$; [HI] = 0.467 *M* 7) 0.80 8) $[CO] = [H_2O] = 0.234\ M$ $[H_2] = [CO_2] = 0.266\ M$