

RICE Tables: Delicious, Delicious Equilibrium

Many equilibrium calculations are solved using RICE tables: this is an organization method used to clearly track what happens as a reaction established equilibrium.

Tasty RICE

- **R** = Balanced reaction
- **I** = initial concentrations (or pressures) for each species in the reaction mixture
- **C** = change in concentrations (or pressures) for each species as the system moves towards equilibrium: everything changes stoichiometrically!
- **E** = equilibrium concentrations (or pressures) of each species when the system reaches equilibrium

Note: If the amounts are given in moles BE WARY – you must convert to molarity (M).

Example: Consider the reaction $2 \text{SO}_3(\text{g}) \rightleftharpoons \text{O}_2(\text{g}) + 2 \text{SO}_2(\text{g})$

A reaction container has an initial $[\text{SO}_3]$ of 0.020 M. At equilibrium, $[\text{SO}_2] = 0.012 \text{ M}$. Calculate the value of the equilibrium constant.

| Reaction | $2 \text{SO}_3(\text{g})$ | \rightleftharpoons | $\text{O}_2(\text{g})$ | + | $2 \text{SO}_2(\text{g})$ |
|---------------------------|---------------------------|----------------------|------------------------|---|---------------------------|
| Initial concentration | 0.020 | | \emptyset | | \emptyset |
| Change | $-2x$ | | $+x$ | | $+2x$ |
| Equilibrium concentration | $0.020 - 2x$ | | x | | $2x = 0.012 \text{ M}$ |

$$= 0.020 - 2(0.0060) = 0.008 \text{ M} \quad \Rightarrow x = 0.0060 \text{ M}$$

$$K = \frac{[\text{O}_2][\text{SO}_2]^2}{[\text{SO}_3]^2} = \frac{(0.0060)(0.012)^2}{(0.008)^2} = 0.0135 = \boxed{0.01} \quad \text{1 s.f.}$$

Now you try! Consider the following reaction: $2 \text{CH}_4(\text{g}) \rightleftharpoons \text{C}_2\text{H}_2(\text{g}) + 3 \text{H}_2(\text{g})$

A reaction mixture at 1700 °C initially contains $[\text{CH}_4] = 0.115 \text{ M}$. At equilibrium, the mixture contains $[\text{C}_2\text{H}_2] = 0.035 \text{ M}$.

What is the value of the equilibrium constant?

$$\begin{array}{l}
 \text{R } 2\text{CH}_4(\text{g}) \rightleftharpoons \text{C}_2\text{H}_2(\text{g}) + 3\text{H}_2(\text{g}) \\
 \text{I } 0.115 \quad \quad \quad \emptyset \quad \quad \quad \emptyset \\
 \text{C } -2x \quad \quad \quad +x \quad \quad \quad +3x \\
 \text{E } 0.115 - 2x \quad x = 0.035 \quad 3x = 3(0.035) \\
 \quad = 0.115 - 2(0.035) \quad \quad \quad = 0.105 \text{ M} \quad] \text{ 2 s.f.} \\
 \quad = 0.045 \text{ M}
 \end{array}$$

$$K_c = \frac{[\text{C}_2\text{H}_2][\text{H}_2]^3}{[\text{CH}_4]^2} = \frac{(0.035)(0.105)^3}{(0.045)^2} = \boxed{0.020}$$

Questionable Q: Predicting the Direction of a Reaction with Yummy RICE!

If you're given initial concentrations of both reactants and products, you need to calculate Q to predict the direction of shift before using a RICE table to calculate equilibrium concentration.

Example: Consider the reaction: $2 \text{NO}(g) \rightleftharpoons \text{N}_2(g) + \text{O}_2(g)$ $K_c = 1.00 \times 10^{-6}$ At 298 K, 2.25 moles of NO, 0.0749 moles of N_2 , and 0.0750 moles of O_2 are placed into a 1.50 L flask and allowed to reach equilibrium. Calculate the equilibrium concentrations for each substance listed below at 298 K.

First, calculate Q. Which direction will the reaction shift?

$$\begin{aligned}
 [\text{NO}] &= 2.250 \text{ mol} / 1.50 \text{ L} = 1.50 \text{ M} \\
 [\text{N}_2] &= 0.0749 \text{ mol} / 1.50 \text{ L} = 0.0500 \text{ M} \\
 [\text{O}_2] &= 0.0750 \text{ mol} / 1.50 \text{ L} = 0.0500 \text{ M}
 \end{aligned}$$

$$Q = \frac{[\text{N}_2][\text{O}_2]}{[\text{NO}]^2} = \frac{(0.0500)^2}{(1.50)^2} = 1.11 \times 10^{-3}$$

$$K < Q \Rightarrow \text{rxn will shift left to make more reactants + reach equilibrium.}$$

$$1.00\text{E-}6 < 1.11\text{E-}3$$

Now you can make a RICE table! (Knowing which direction reaction shifts allows you to determine correct signs for x.)

| Reaction | $2 \text{NO}(g)$ | \rightleftharpoons | $\text{N}_2(g)$ | + | $\text{O}_2(g)$ |
|---------------------------|------------------|----------------------|-----------------|---|-----------------|
| Initial concentration | 1.50 | | 0.0500 | | 0.0500 |
| Change | +2x | | -x | | -x |
| Equilibrium concentration | $1.50 + 2x$ | | $0.0500 - x$ | | $0.0500 - x$ |

Now you can write your equilibrium expression and plug in with the values from above.

$$K = \frac{[\text{N}_2][\text{O}_2]}{[\text{NO}]^2} = \frac{(0.0500 - x)^2}{(1.50 + 2x)^2} = 1.00 \times 10^{-6}$$

to solve, take square root of both sides!

$$\frac{0.0500 - x}{1.50 + 2x} = 1.00\text{E-}3 \Rightarrow 0.0500 - x = 1.00\text{E-}3(1.50 + 2x)$$

$$= 0.00150 + 0.00200x$$

$$0.0485 = 1.00200x \Rightarrow x = \frac{0.0485}{1.00200} = 0.0484$$

Finally, what are your equilibrium concentrations?

$$[\text{NO}_2] = 1.50 + 2x = 1.60 \text{ M}$$

$$[\text{N}_2] = [\text{O}_2] = 0.0500 - x = 0.0016 \text{ M}$$

Final answer! ^_^

Check yourself!

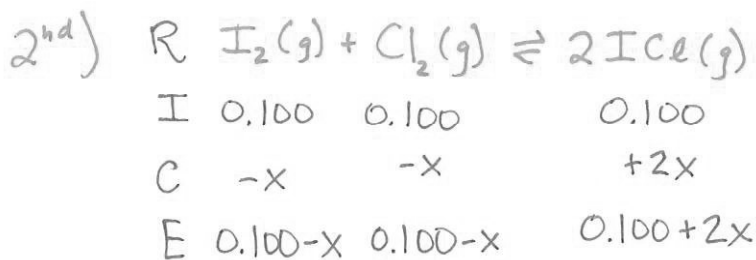
$$K = \frac{(0.0016)^2}{1.60^2} = 1 \times 10^{-6} \checkmark$$

Now you try! Consider the reaction: $I_2(g) + Cl_2(g) \rightleftharpoons 2 ICl(g)$, $K_p = 81.9$ (at 25°C).

A reaction mixture at 25°C initially contains $P_{I_2} = 0.100$ atm, $P_{Cl_2} = 0.100$ atm, and $P_{ICl} = 0.100$ atm. Find the equilibrium partial pressures of I_2 , Cl_2 , and ICl at this temperature.

1st)
$$Q_p = \frac{(P_{ICl})^2}{(P_{I_2})(P_{Cl_2})} = \frac{(0.100)^2}{(0.100)^2} = 1.00$$

$$\left. \begin{array}{l} K > Q \\ 81.9 > 1.00 \end{array} \right\} \Rightarrow \text{rxn will shift right to make more products}$$



4th)
$$P_{ICl} = 0.100 + 2x$$

$$= 0.100 + 2(0.0729) = \boxed{0.246 \text{ atm}}$$

$$P_{I_2} = P_{Cl_2} = 0.100 - x$$

$$= 0.100 - 0.0729$$

$$= \boxed{0.0271 \text{ atm}}$$

3rd)
$$K_p = \frac{(P_{ICl})^2}{(P_{I_2})(P_{Cl_2})} = \frac{(0.100 + 2x)^2}{(0.100 - x)^2} = 81.9$$

$$\sqrt{81.9} = 9.05 = \frac{0.100 + 2x}{0.100 - x}$$

$$0.905 - 9.05x = 0.100 + 2x$$

$$0.805 = 11.05x$$

$$x = \frac{0.805}{11.05} = 0.0729$$

* 5th) if time permits \Rightarrow check!

$$K = \frac{0.246^2}{0.027^2} = 83 \text{ (ugh, rounding - but close enough!)}$$

Multiple Choice Practice!

1. Ammonia and oxygen react according to the following equilibrium: $4 NH_3(g) + 3 O_2(g) \rightleftharpoons 2 N_2(g) + 6 H_2O(g)$.

A 2.0 liter flask is initially filled with 8.0 mol of oxygen and 6.0 mol of ammonia, and 6.0 mol of H_2O are present when the system reaches equilibrium. How much oxygen is present at equilibrium?

a. 2.0 mol O_2

b. 3.0 mol O_2

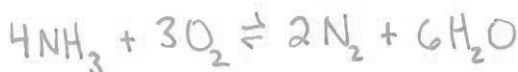
c. 5.0 mol O_2

d. 6.0 mol O_2

$$[NH_3]_i = \frac{6 \text{ mol}}{2 \text{ L}} = 3 \text{ M}$$

$$[O_2]_i = \frac{8 \text{ mol}}{2 \text{ L}} = 4 \text{ M}$$

$$[H_2O]_{eq} = \frac{6 \text{ mol}}{2 \text{ L}} = 3 \text{ M}$$



| | | | |
|-----|-----|-------------|-------------|
| 3 | 4 | \emptyset | \emptyset |
| -4x | -3x | +2x | +6x |

| | | | |
|------|------|----|---------|
| 3-4x | 4-3x | 2x | 6x = 3M |
|------|------|----|---------|

$$\Rightarrow x = 0.5$$

$$[O_2]_{eq} = 4 - 3(0.5)$$

$$= 4 - 1.5 = 2.5 \text{ M} \times 2 \text{ L} = 5 \text{ mol } O_2$$