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 $2\ SO_3(g) \rightleftharpoons O_2(g) + 2\ SO_2(g)$

A reaction container has an initial $[SO_3]$ of 0.020 M. At equilibrium, $\boxed{[SO_2] = 0.012\ M}$. Calculate the value of the equilibrium constant.

Reaction	$2\ SO_3\ (g)$ $\rightleftharpoons$	$O_2(g)$ +	$2\ SO_2(g)$
Initial concentration	0.020	∅	∅
Change	$-2x$	$+x$	$+2x$
Equilibrium concentration	$0.020-2x$	x	$2x = 0.012\ M$

$= 0.020 - 2(0.0060)$ $= 0.0060\ M$ $\Rightarrow x = 0.0060\ M$

$= 0.008\ M$] 1 s.f.

$$K = \frac{[O_2][SO_2]^2}{[SO_3]^2} = \frac{(0.0060)(0.012)^2}{(0.008)^2} = 0.0135 = \boxed{0.01}$$

1 s.f. ☺

Now you try! Consider the following reaction: $2\ CH_4(g) \rightleftharpoons C_2H_2(g) + 3\ H_2(g)$

A reaction mixture at 1700 °C initially contains $[CH_4] = 0.115$ M. At equilibrium, the mixture contains $\boxed{[C_2H_2] = 0.035\ M.}$ What is the value of the equilibrium constant?

R	$2CH_4(g) \rightleftharpoons$	$C_2H_2(g)$ +	$3H_2(g)$
I	0.115	∅	∅
C	$-2x$	$+x$	$+3x$
E	$0.115-2x$	$x = 0.035$	$3x = 3(0.035)$

$= 0.115 - 2(0.035)$ $= 0.105\ M$] 2 s.f.

$= 0.045\ M$

$$K_c = \frac{[C_2H_2][H_2]^3}{[CH_4]^2} = \frac{(0.035)(0.105)^3}{(0.045)^2} = \boxed{0.020}$$

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

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

Example: Consider the reaction: $2\ NO(g) \rightleftharpoons N_2(g) + O_2(g)$ $K_c = 1.00 \times 10^{-6}$ 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?

$[NO] = 2.250\ mol / 1.50\ L = 1.50\ M$

$[N_2] = 0.0749\ mol / 1.50\ L = 0.0500\ M$

$[O_2] = 0.0750\ mol / 1.50\ L = 0.0500\ M$

$$Q = \frac{[N_2][O_2]}{[NO]^2} = \frac{(0.0500)^2}{(1.50)^2} = 1.11 \times 10^{-3}$$

$K < Q$ ⇒ rxn will shift left to make more reactants + reach equilibrium

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

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

Reaction	$2\ NO(g)$ ⇌	$N_2(g)$ +	$O_2(g)$
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.

$$K = \frac{[N_2][O_2]}{[NO]^2} = \underbrace{\frac{(0.0500 - x)^2}{(1.50 + 2x)^2}}_{\text{to solve, take square root of both sides!}} = 1.00 \times 10^{-6}$$

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

$0.0500 - x = 1.00E{-}3(1.50 + 2x)$

$= 0.0150 + 0.00200x$

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

Finally, what are your equilibrium concentrations?

$$\boxed{\begin{aligned}[NO_2] &= 1.50 + 2x = 1.60\ M \\ [N_2] = [O_2] &= 0.0500 - x = 0.0016\ M\end{aligned}}$$

Final answer!

Check yourself!

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

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) $Q_p = \frac{(P_{ICl})^2}{(P_{I_2})(P_{Cl_2})} = \frac{(0.100)^2}{(0.100)^2} = 1.00$ $\Big\}$ $K > Q$, $81.9 > 1.00$ $\Rightarrow$ rxn will shift right to make more products

2nd)

R	$I_2(g)$	+ $Cl_2(g)$	$\rightleftharpoons 2ICl(g)$
I	0.100	0.100	0.100
C	$-x$	$-x$	$+2x$
E	$0.100-x$	$0.100-x$	$0.100+2x$

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

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

$0.905 - 9.05x = 0.100 + 2x$

$0.805 = 11.05x$

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

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

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

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

$= 0.100 - 0.0729$

$= \boxed{0.027 \text{ atm}}$

*5th) if time permits ⇒ check!

$K = \frac{0.246^2}{0.027^2} = 83$ (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

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

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

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

$4NH_3$	+ $3O_2$	$\rightleftharpoons 2N_2$	+ $6H_2O$
3	4	Ø	Ø
$-4x$	$-3x$	$+2x$	$+6x$
$3-4x$	$\underbrace{4-3x}$	$2x$	$6x = 3\text{M}$

$\Rightarrow x = 0.5$

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

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