

### The 3'rd Kind - Keq Calculations involving ICE tables

Remember:

*A chemical system can be thought of as being either:*

*1. At Equilibrium*

*or*

*2. Not At Equilibrium*

*A system which is **not at equilibrium** will move spontaneously to a position of **being at equilibrium**.*

In this type of problem, there will be one species which we will know the concentration of **initially and at equilibrium**. We can find the **change in the concentration** (which we abbreviate as “[C]” where the “C” stands for the words “Change in” and [ ]’s stand for Concentration) of this species and by using **mole ratios in the balanced equation**, find the changes in concentration “[C]” of the other species. From this we can calculate the **equilibrium concentration** (which we abbreviate as “[E]”) of all the species.

3rd type of Keq

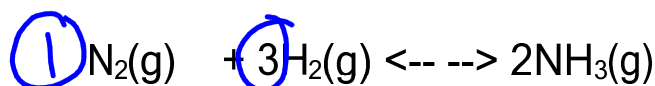
## ICE Examples # 1

Given the reaction:  $\text{N}_{2(g)} + 3\text{H}_{2(g)} \rightleftharpoons 2\text{NH}_{3(g)}$

Some  $\text{H}_2$  and  $\text{N}_2$  are added to a container so that initially, the  $[\text{N}_2] = 0.32 \text{ M}$  and  $[\text{H}_2] = 0.66 \text{ M}$ . At a certain temperature and pressure, the equilibrium  $[\text{H}_2]$  is found to be  $0.30 \text{ M}$ .

- ✓ a) Find the equilibrium  $[\text{N}_2]$  and  $[\text{NH}_3]$ .  
b) Calculate  $K_{eq}$  at this temperature and pressure.

Set up ICE table under the reaction



Initial Conc. (I)	0.32 M	0.66 M	0
Change Conc. (C)	-0.12 M	-0.36 M	+0.24 M
Equilibrium Conc (E)	0.20 M	0.30 M	0.24 M

$$K_{eq} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(0.24)^2}{(0.2)(0.3)^3} = 10.7$$

✓ Add in the info from problem

✓ Find the change from initial to equilibrium

✓ Combine (I and C) to find Equilibrium concentrations

$$K_{eq} = 11$$

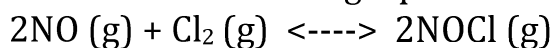
First we write the expression for  $K_{eq}$ :  $\text{N}_{2(g)} + 3 \text{H}_{2(g)} \rightleftharpoons 2\text{NH}_{3(g)}$

$$K_{eq} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

To calculate  $K_{eq}$ , we plug in the values for the equilibrium concentrations of all the species.

## Example 2

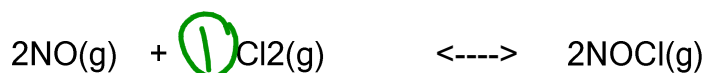
Consider the following equilibrium system:



0.80 moles of NO and 0.60 moles of Cl<sub>2</sub> are placed into a 1.0 L container and allowed to establish equilibrium. At equilibrium [NOCl] = 0.56 M.

- Calculate the equilibrium [NO]  $0.24\text{ M}$
- Calculate the equilibrium [Cl<sub>2</sub>]  $0.32\text{ M}$
- Determine the value of K<sub>eq</sub> at this temperature.

NOTE: In a 1.0 Litre container!!! so what is the [conc]??



[I]	0.80 M	0.60 M	0
[C]	-0.56 M	-0.28 M	+0.56 M
[E]	0.24 M	0.32 M	0.56 M

$$K_{eq} = \frac{[\text{NOCl}]^2}{[\text{NO}]^2 [\text{Cl}_2]} = \frac{(0.56)^2}{(0.24)^2 (0.32)} = 17$$

### Example 3:

In another variation of ICE Problems, we are sometimes given the **initial moles** when we have something *other than a 1.0 L container*. In this case, you must find **initial concentrations [I]** by using the familiar formula:

$$\text{Molarity}(M) = \frac{\text{mol}}{L}$$

Let's do an example:

Consider the equilibrium system:  $A + 3B \rightleftharpoons 2C$

0.20 moles of A and 0.60 moles of B are placed in a 2.0 L container. When equilibrium is reached, the [A] is found to be 0.08 M. Calculate the **equilibrium [B]** and the **equilibrium [C]**

$$\begin{aligned} \text{Initial [A]} &= 0.10 \text{ M} \\ \text{Initial [B]} &= 0.30 \text{ M} \end{aligned}$$

Notice that in this case the **equilibrium concentration** (not moles!) of A is given. This can go right in the table under **equilibrium concentration [E]** of A.

	A +	3B <--- -->	2C
[I]	0.10 M	0.30 M	0
[C]	-0.02 M	-0.06 M	+0.04 M
[E]	0.08 M	0.24 M	0.04 M

$$K_{eq} = \frac{[C]^2}{[A][B]^3} = 1.4$$