# Example 4:

In one more variation of "ICE", they may give you initial moles and equilibrium moles (not equilibrium concentration) in something other than a 1.0 L container.

In this case, you would have to calculate the [E] (Equilibrium concentration) from the equilibrium moles and the volume of the container using:

$$M = \frac{mol}{L}$$

Given the equilibrium: X (g) + 2Y (g) <-- --> 2 Z (g)

When 2.0 moles of X and 3.5 moles of Y are placed in a 5.0 L container at 25°C, an equilibrium is established in which there are 2.5 moles of Z. /5

Calculate [X], [Y] and [Z] at equilibrium and the Keq.

2/5 3.5/5

| | X (g) + | 2Y (g) <-- --> | 2 Z (g) |
|---|---|---|---|
| [I] | 0.40 | 0.70 | 0 |
| [C] | −0.25 | −0.50 | +0.50 |
| [E] | 0.15 | 0.20 | 0.5 |

Notice that ***moles*** of Z (not [Z]) is given at equilibrium. We can find the Equilibrium [Z] using the formula: M = moles/ L. This can then be placed in the table and the rest of the calculations can be done:

*Equilibrium [ Z ] = mol / L =*

*Therefore the change in concentration for Z must be:*

*Using mole ratios : determine the change in concentrations for all others.*

*Keq =* $\frac{[Z]^2}{[X][Y]^2}$

$$K_{eq} = \frac{[Z]^2}{[X]\cdot[Y]^2}$$

$$K_{eq} = \frac{(0.50)^2}{(0.15)(0.20)^2} = \quad K_{eq} = 42 > 1$$

favors products

## One more example - you try:

Given the equilibrium equation:
A + 2B <-- --> C

$\div 10$ $\div 10$

When 2.0 moles of A and 4.0 moles of B are added to a 10.0 L container, an equilibrium established in which 1.4 moles of C are found.

$\div 10$

Find the equilibrium concentrations of A, B and C.

| | A + | 2B <-- --> | C |
|---|---|---|---|
| [I] | 0.20 | 0.40 | 0 |
| [C] | −0.14 | −0.28 | +0.14 |
| [E] | 0.06 | 0.12 | 0.14 |

$$K_{eq} = \frac{[C]}{[A]\cdot[B]^2} = \frac{0.14}{(0.06)(0.12)^2} = 162$$

$$\boxed{K_{eq} = 160}$$

Products

Keq When Given Initial concentrations and Keq value only!!

Example:

Given the reaction :
H2 (g) + I2 (g) <--> 2HI (g) The Keq = 55.6

If the initial [H2] = 0.200 M and [I2] = 0.200M, what is the equilibrium [HI]?

| | $H_2$ + | $I_2$ <-> | $2HI$ |
|---|---|---|---|
| [I] | 0.200 M | 0.200 M | 0 |
| [C] | $-x$ | $-x$ | $+2x$ |
| [E] | $0.2-x$ | $0.2-x$ | $2x$ |

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

$$55.6 = \frac{(2x)^2}{(0.2-x)(0.2-x)}$$

$$\sqrt{55.6} = \sqrt{\frac{(2x)^2}{(0.2-x)^2}}$$

$$7.4565 = \frac{2x}{(0.2-x)}$$

$$1.4913 - \cancel{7.4565x} = 2x$$
$$\cancel{+7.4565x} \qquad +7.4565x$$

$$\frac{1.4913}{9.4565} = \frac{\cancel{9.4565}x}{\cancel{9.4565}}$$

$$0.1577 = x$$

$$[HI] = 2x$$
$$= 0.3154M$$

$$\boxed{[HI] = 0.315M}$$

Consider the reaction:
H2 (g) + CO2 (g) <--> H2O (g) + CO (g)

At a certain temp, Keq = 1.50. If the initial concentration values of the reactants is 0.500M, calculate the equilibrium concentration of all species.

| | $H_2$ + | $CO_2$ ⇌ | $H_2O$ + | $CO$ |
|---|---|---|---|---|
| [I] | 0.500M | 0.500M | 0 | 0 |
| [C] | $-x$ | $-x$ | $+x$ | $+x$ |
| [E] | $0.5-x$ | $0.5-x$ | $x$ | $x$ |

$$K_{eq} = \frac{[H_2O][CO]}{[H_2][CO_2]}$$

$$1.50 = \frac{(x)(x)}{(0.5-x)(0.5-x)}$$

$$\sqrt{1.50} = \sqrt{\frac{x^2}{(0.5-x)^2}}$$

$$1.2247 = \frac{x}{0.5-x}$$

$$0.61237 - 1.224x = x$$

$$\frac{0.61237}{2.224} = \frac{\cancel{2.224}x}{\cancel{2.224}}$$

$$0.27526 = x$$

$[H_2O] = 0.275M$

$[CO] = 0.275M$

$[H_2] = 0.5 - 0.275 = 0.225M$

$[CO_2] = 0.225M$

18.2: Equilibrium Constant Equations (Harder Level)

1. At equilibrium, a 5.0L flask contains:
   0.75 moles of $PCl_5$, 0.50 mol of $H_2O$, 7.50 mol of HCl, and 5.00 mol of $POCl_3$.

   Calculate the Keq for the reaction: $PCl_5$ (s) + $H_2O$ (g) ←→ 2HCl (g) + $POCl_3$ (g)
   (A: 23)

2. Given the equilibrium reaction: $2NO_2$ (g) ←→ $N_2O_4$ (g)

   If 2.00 moles of $NO_2$ are placed in a 1.00 L flask and allowed to react. At equilibrium, 1.80 moles of $NO_2$ are present. Calculate the Keq. (A: 0.031)

3. $SO_3$ (g) + NO (g) ←→ $NO_2$ (g) + $SO_2$ (g) **Keq** = 0.800 at 100°C

   If 4.00 moles of each reactant are placed in a 2.00 L container, calculate all equilibrium concentrations at 100°C. (A: 0.944M = [$NO_2$]=[$SO_2$]; 1.06 M = [$SO_3$]=[NO])

4. Consider the following equilibrium equation: $2SO_2$ (g) + $O_2$ (g) ←→ $2SO_3$ (g)

   When a 0.600 moles of $SO_2$ and 0.600 moles of $O_2$ are placed into a 2.00 litre container and allowed to reach equilibrium, the equilibrium [$SO_3$] is to be 0.250M. Calculate the Keq value. (A: 143)

5. $H_2$ (g) + S (s) ← → $H_2S$ (g) **Keq = 14**

   0.60 moles of $H_2$ and 1.40 moles of S are placed into a 2.0 L flask and allowed to reach equilibrium. Calculate the [$H_2$] at equilibrium. (A: 0.02 M)

6. $I_2$ (g) + $Cl_2$ (g) ← → 2ICl (g) **Keq = 10.0**

   The same number of moles and $I_2$ and $Cl_2$ are placed into a 1.0 L flask and allowed to reach equilibrium. If the equilibrium concentration of ICl is 0.040 M, calculate the initial concentration and moles of $I_2$ and $Cl_2$. (A: 0.033 M; 0.033 mol)

7. Consider the equilibrium : 2ICl (g) ←→ $I_2$ (g) + $Cl_2$ (g) **Keq = 10.0**

   If *x* moles of ICl were placed into a 5.0 L container at 10°C and if an equilibrium concentration of $I_2$ was found to be 0.60 M, calculate the number of moles ICl initially present. (A: 6.9 mole)

8. A student places 2.00 moles $SO_3$ in a 1.00 L flask. At equilibrium [$O_2$] = 0.10 M at 130°C. Calculate Keq. (A: 810)

   $2SO_2$ (g) + $O_2$ (g) ←→ $2SO_3$ (g)