

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:

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

Given the equilibrium: $X(g) + 2Y(g) \rightleftharpoons 2Z(g)$

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.

Calculate [X], [Y] and [Z] at equilibrium and the K_{eq} .

	$\frac{2.0}{5}$	$\frac{3.5}{5}$	
	$X(g) + 2Y(g) \rightleftharpoons 2Z(g)$		
[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: $M = \text{moles} / L$. 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.

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

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

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

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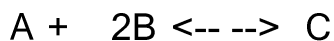
One more example - you try:

Given the equilibrium equation:



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.

Find the equilibrium concentrations of A, B and C.



[I]	0.20	0.40	0
[C]	-0.14	-0.28	+0.14
[E]	0.06	0.12	0.14

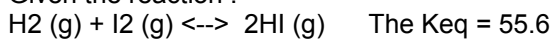
$$K_{eq} = \frac{[C]}{[A] \cdot [B]^2} = \frac{0.14}{(0.06)(0.12)^2} = 162$$

$K_{eq} = 160$
 Products

Keq When Given Initial concentrations and Keq value only!!

Example:

Given the reaction :



If the initial $[\text{H}_2] = 0.200 \text{ M}$ and $[\text{I}_2] = 0.200 \text{ M}$, what is the equilibrium $[\text{HI}]$?



[I]	0.200 M	0.200 M	0
[C]	-x	-x	+2x
[E]	0.2-x	0.2-x	2x

$$K_{eq} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

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

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

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

$$1.4913 - 7.4565x = 2x$$

$$1.4913 = 9.4565x$$

$$0.1577 = x$$

$$[\text{HI}] = 2x = 0.3154 \text{ M}$$

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

Consider the reaction:



At a certain temp, $K_{eq} = 1.50$. 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

$$K_{eq} = \frac{[\text{H}_2\text{O}][\text{CO}]}{[\text{H}_2][\text{CO}_2]}$$

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

$$\sqrt{1.50} = \frac{x}{0.5-x}$$

$$0.61237 - 1.224x = x$$

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

$$0.27526 = x$$

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

$$[\text{H}_2\text{O}] = 0.275\text{M}$$

$$[\text{CO}] = 0.275\text{M}$$

$$[\text{H}_2] = 0.5 - 0.275 = 0.225\text{M}$$

$$[\text{CO}_2] = 0.225\text{M}$$

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 K_{eq} for the reaction: $\text{PCl}_5(\text{s}) + \text{H}_2\text{O}(\text{g}) \leftrightarrow 2\text{HCl}(\text{g}) + \text{POCl}_3(\text{g})$

(A: 23)

2. Given the equilibrium reaction:
- $2\text{NO}_2(\text{g}) \leftrightarrow \text{N}_2\text{O}_4(\text{g})$

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 K_{eq} . (A: 0.031)

- 3.
- $\text{SO}_3(\text{g}) + \text{NO}(\text{g}) \leftrightarrow \text{NO}_2(\text{g}) + \text{SO}_2(\text{g})$
- $K_{\text{eq}} = 0.800$
- 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: $0.944\text{M} = [\text{NO}_2] = [\text{SO}_2]$; $1.06\text{M} = [\text{SO}_3] = [\text{NO}]$)

4. Consider the following equilibrium equation:
- $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \leftrightarrow 2\text{SO}_3(\text{g})$

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 $[\text{SO}_3]$ is to be 0.250M. Calculate the K_{eq} value. (A: 143)

- 5.
- $\text{H}_2(\text{g}) + \text{S}(\text{s}) \leftrightarrow \text{H}_2\text{S}(\text{g})$
- $K_{\text{eq}} = 14$

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 $[\text{H}_2]$ at equilibrium. (A: 0.02 M)

- 6.
- $\text{I}_2(\text{g}) + \text{Cl}_2(\text{g}) \leftrightarrow 2\text{ICl}(\text{g})$
- $K_{\text{eq}} = 10.0$

The same number of moles of 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:
- $2\text{ICl}(\text{g}) \leftrightarrow \text{I}_2(\text{g}) + \text{Cl}_2(\text{g})$
- $K_{\text{eq}} = 10.0$

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
- $[\text{O}_2] = 0.10\text{M}$
- at
- 130°C
- . Calculate
- K_{eq}
- . (A: 810)

