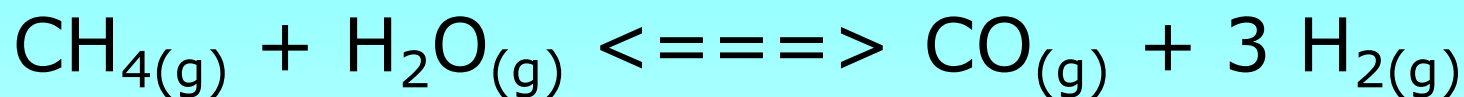


# CALCULATIONS GIVEN $K_{eq}$

# CALCULATIONS GIVEN $K_{eq}$

## Example #1



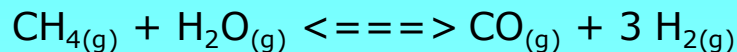
At  $1500^\circ\text{C}$ ,  $K_{eq} = 5.67$

$[\text{CO}] = 0.300 \text{ M}$ ,  $[\text{H}_2] = 0.800 \text{ M}$  and  
 $[\text{CH}_4] = 0.400 \text{ M}$

What is  $[\text{H}_2\text{O}]$  at equilibrium?

# CALCULATIONS GIVEN $K_{eq}$

Example #1



At 1500°C,  $K_{eq} = 5.67$ ,  $[\text{CO}] = 0.300 \text{ M}$ ,  $[\text{H}_2] = 0.800 \text{ M}$  and,  $[\text{CH}_4] = 0.400 \text{ M}$

What is  $[\text{H}_2\text{O}]$  at equilibrium?

$$K_{eq} = \frac{[\text{CO}][\text{H}_2]^3}{[\text{CH}_4][\text{H}_2\text{O}]}$$

$$[\text{H}_2\text{O}] = \frac{[\text{CO}][\text{H}_2]^3}{[\text{CH}_4] K_{eq}}$$

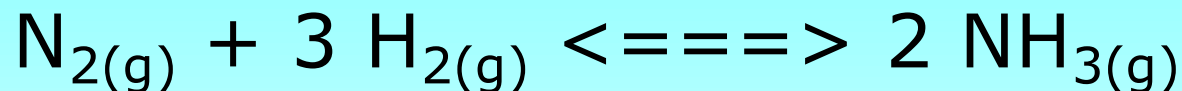
$$[\text{H}_2\text{O}] = \frac{[0.300\text{M}][0.800\text{M}]^3}{[0.400\text{M}] 5.67}$$

$$[\text{H}_2\text{O}] = 0.0677\text{M}$$

$\therefore [\text{H}_2\text{O}]$  at equilibrium was  $6.77 \times 10^{-2}\text{M}$

# CALCULATIONS GIVEN $K_{eq}$

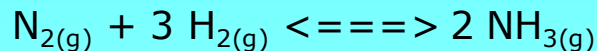
## Example #2



What is  $[\text{NH}_3]$  when  $[\text{N}_2] = 0.45 \text{ M}$ ,  
 $[\text{H}_2] = 1.10$  and  $K_{eq} = 1.7 \times 10^{-2}$  ?

# CALCULATIONS GIVEN $K_{eq}$

Example #2



What is  $[\text{NH}_3]$  when  $[\text{N}_2] = 0.45 \text{ M}$ ,  $[\text{H}_2] = 1.10 \text{ M}$  and  $K_{eq} = 1.7 \times 10^{-2}$  ?

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

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

$$0.017 [0.45\text{M}][1.10\text{M}]^3 = [\text{NH}_3]^2$$

$$0.01018215 = [\text{NH}_3]^2$$

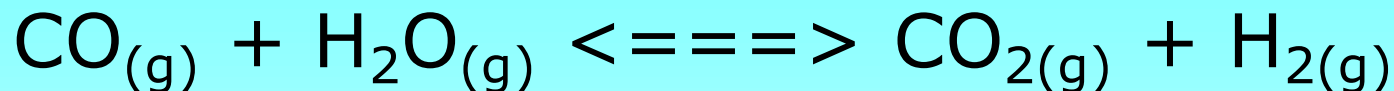
$$\sqrt{0.01018215} = \sqrt{[\text{NH}_3]^2}$$

$$0.10090664 = [\text{NH}_3]$$

$\therefore$  the  $[\text{NH}_3]$  is  $1.0 \times 10^{-1} \text{ M}$

# CALCULATIONS GIVEN $K_{eq}$

## Example #3



At equilibrium,  $K_{eq} = 4.06$ .

If 0.100 mol of CO and 0.100 mol of  $\text{H}_2\text{O}_{(g)}$  are placed in a 1.00 L container,

- What are the concentrations of the reactants and products at equilibrium?
- What is the final mass of  $\text{CO}_{2(g)}$ ?

# CALCULATIONS GIVEN $K_{eq}$

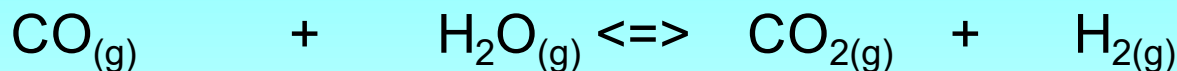
Example #3

At equilibrium,  $K_{eq} = 4.06$ .

If 0.100 mol of CO and 0.100 mol of  $H_2O_{(g)}$  are placed in a 1.00 L container,

a) What are the concentrations of the reactants and products at equilibrium?

b) What is the final mass of  $CO_{2(g)}$ ?

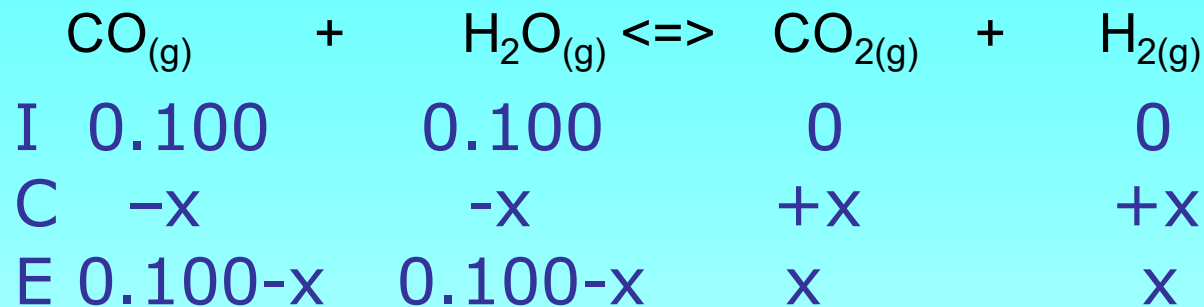


I	0.100	0.100	0	0
C	-x	-x	+x	+x
E	0.100-x	0.100-x	x	x

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

$$4.06 = \frac{[x][x]}{[0.100-x][0.100-x]}$$

# CALCULATIONS GIVEN $K_{eq}$



$$K_{eq} = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}$$
$$4.06 = \frac{[x][x]}{[0.100-x][0.100-x]}$$

$$4.06 [0.100-x][0.100-x] = x^2$$

$$(0.406-4.06x)[0.100-x] = x^2$$

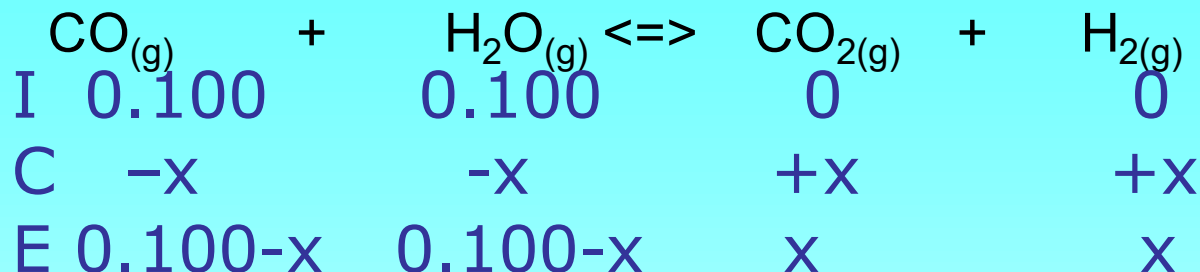
$$0.0406 - 0.406x - 0.406x + 4.06x^2 = x^2$$

$$3.06x^2 - 0.812x + 0.0406 = 0$$

Time to use the quadratic formula!



# CALCULATIONS GIVEN $K_{eq}$



$$3.06x^2 - 0.812x + 0.0406 = 0$$

Quadratic equation:  $ax^2 + bx + c = 0$

Quadratic formula:  $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$$x = \frac{-(-0.812) \pm \sqrt{(-0.812)^2 - 4(3.06)(0.0406)}}{2(3.06)}$$

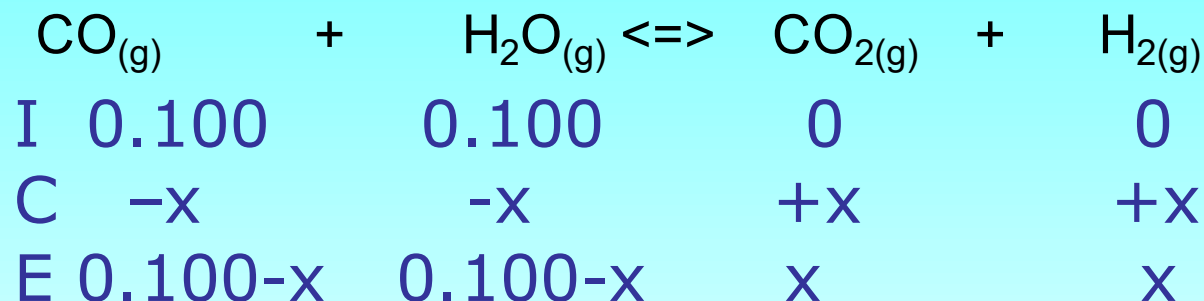
$$x = 0.19853 \text{ or } 0.066832$$

This is the correct answer. Otherwise you may end up with negative reactant concentrations

# CALCULATIONS GIVEN $K_{eq}$

Example #3

- a) What are the concentrations of the reactants and products at equilibrium?  
b) What is the final mass of  $\text{CO}_{2(g)}$ ?



$$x = 0.066832$$



a)  $\therefore [\text{CO}_{(g)}] = 3.32 \times 10^{-2} \text{M}$ ,  $[\text{H}_2\text{O}_{(g)}] = 3.32 \times 10^{-2} \text{M}$ ,  
 $[\text{CO}_{2(g)}] = 6.68 \times 10^{-2} \text{M}$ ,  $[\text{H}_{2(g)}] = 6.68 \times 10^{-2} \text{M}$

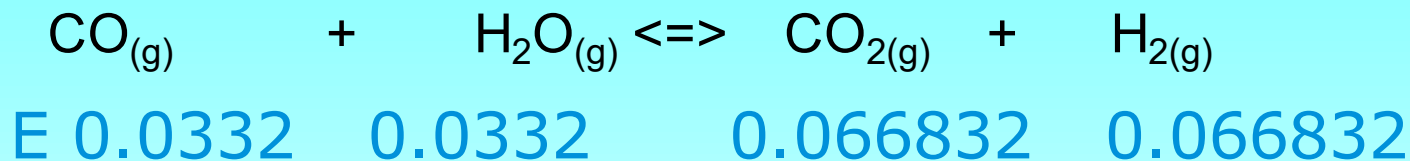
# CALCULATIONS GIVEN $K_{eq}$

Example #3

If 0.100 mol of CO and 0.100 mol of  $H_2O_{(g)}$  are placed in a 1.00 L container,

a) What are the concentrations of the reactants and products at equilibrium?

b) What is the final mass of  $CO_{2(g)}$ ?



$$[CO_{2(g)}] = 6.68 \times 10^{-2} M$$

$$C = \frac{n}{V}$$

$$n = CV$$

$$n = 0.066832 \text{ mol/L} \times 1.00 \text{ L}$$

$$n = 0.066832 \text{ mol}$$

$$n = \frac{m}{M}$$

$$m = nM$$

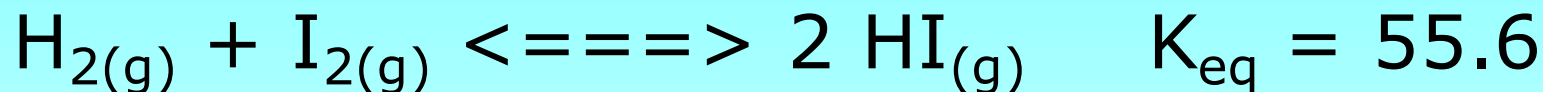
$$m = 0.066832 \text{ mol} \times 44.01 \text{ g/mol}$$

$$m = 2.94 \text{ g}$$

$\therefore 2.94 \text{ g of } CO_{2(g)} \text{ were produced}$

# CALCULATIONS GIVEN $K_{eq}$

## Example #4



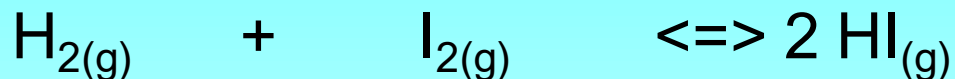
If initial  $[\text{H}_2] = 0.200 \text{ M}$  and initial  $[\text{I}_2] = 0.200 \text{ M}$ . What is  $[\text{HI}]$  at equilibrium?

# CALCULATIONS GIVEN $K_{eq}$

Example #4

If initial  $[H_2] = 0.200\text{ M}$  and initial  $[I_2] = 0.200\text{ M}$ .  $K_{eq} = 55.6$

What is  $[HI]$  at equilibrium?



I	0.200	0.200	0
C	-x	-x	+2x
E	0.200-x	0.200-x	2x

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

$$55.6 = \frac{[HI]^2}{[H_2][I_2]}$$

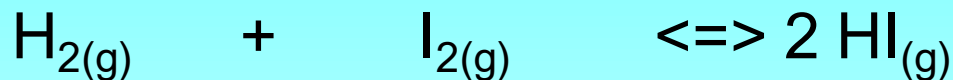
$$55.6 = \frac{[2x]^2}{[0.200-x][0.200-x]}$$

# CALCULATIONS GIVEN $K_{eq}$

Example #4

If initial  $[H_2] = 0.200\text{ M}$  and initial  $[I_2] = 0.200\text{ M}$ .  $K_{eq} = 55.6$

What is  $[HI]$  at equilibrium?



$$55.6 = \frac{[2x]^2}{[0.200-x][0.200-x]}$$

$$\sqrt{55.6} = \sqrt{\frac{4x^2}{[0.200-x]^2}}$$

$$7.456 = \frac{2x}{0.200-x}$$

$$1.4913 - 7.4565x = 2x$$

$$1.4913 = 9.4565x$$

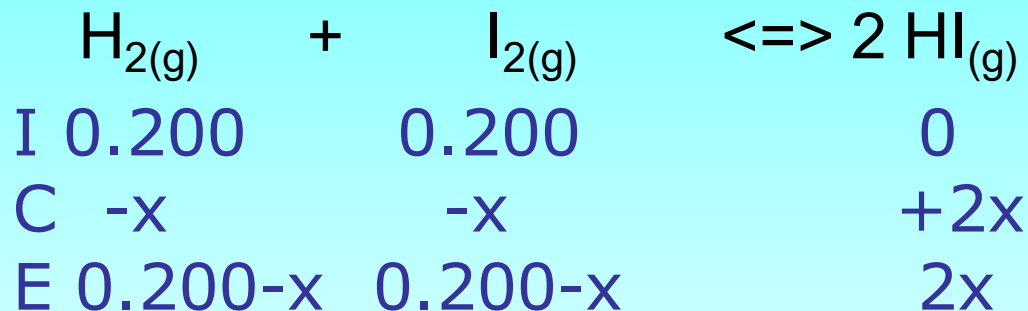
$$0.1577 = x$$

# CALCULATIONS GIVEN $K_{eq}$

Example #4

If initial  $[H_2] = 0.200\text{ M}$  and initial  $[I_2] = 0.200\text{ M}$ .  $K_{eq} = 55.6$

What is  $[HI]$  at equilibrium?



$$x = 0.1577$$

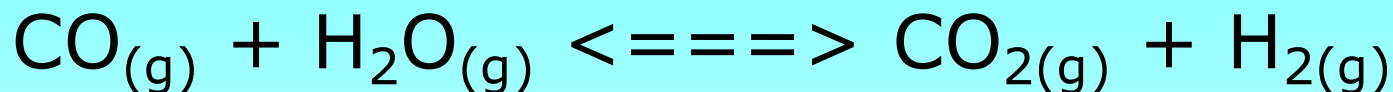
$$2x = 2 (0.1577)$$

$$2x = 0.3154\text{M}$$

$\therefore$  the equilibrium  $[HI_{(aq)}] = 3.15 \times 10^{-1}\text{M}$

# CALCULATIONS GIVEN $K_{eq}$

## Example #5



At equilibrium,  $K_{eq} = 10.0$

A reaction vessel is found to contain  
0.80 M CO, 0.050 M H<sub>2</sub>O, 0.50 M CO<sub>2</sub>  
and 0.40 M H<sub>2</sub>.

Determine if the reaction is at  
equilibrium.



# CALCULATIONS GIVEN $K_{eq}$

Q

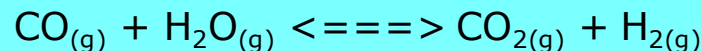
When testing if conditions are at equilibrium Q is the symbol used, rather than  $K_{eq}$ .

Q is the “test  $K_{eq}$ ” variable.

$$Q = \frac{[\text{products}]}{[\text{reactants}]}$$

# CALCULATIONS GIVEN $K_{eq}$

Example #5



At equilibrium,  $K_{eq} = 10.0$ .

A reaction vessel is found to contain

0.80 M CO, 0.050 M H<sub>2</sub>O, 0.50 M CO<sub>2</sub> and 0.40 M H<sub>2</sub>.

Determine if the reaction is at equilibrium.

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

$$Q = \frac{[0.50][0.40]}{[0.80][0.050]}$$

$$Q = 5.0$$

Since  $k_{eq} \neq Q$ , then the reaction is not at equilibrium

∴ the reaction is not at equilibrium

# CALCULATIONS GIVEN $K_{eq}$

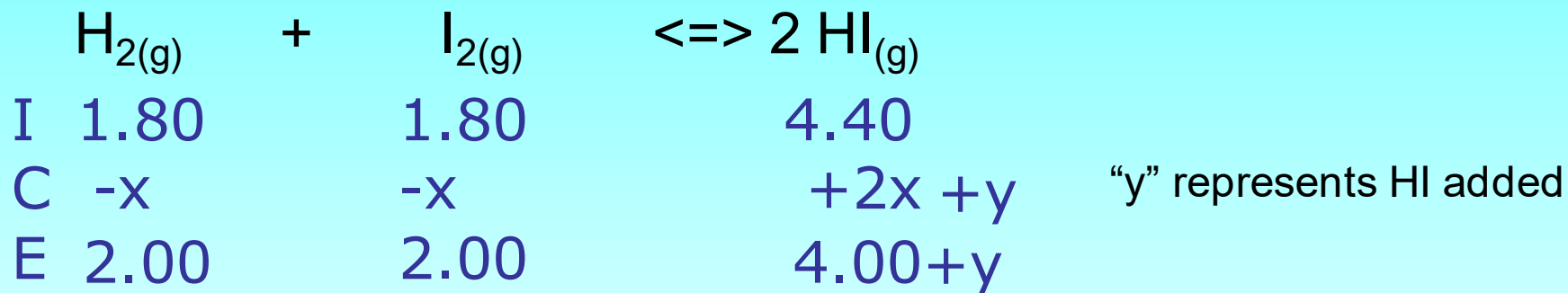
## Example #6

An equilibrium reaction of  $H_2$ ,  $I_2$ , and  $HI$  begins with initial concentrations of 1.80 M, 1.80 M and 4.40 respectively. How much  $HI$  must be added to the equilibrium concentrations to increase the concentration of  $I_2$  to 2.00 M?  $K_{eq}$  is equal to 6.0

# CALCULATIONS GIVEN $K_{eq}$

## Example #6

An equilibrium reaction of  $H_2$ ,  $I_2$ , and  $HI$  begins with initial concentrations of 1.80 M, 1.80 M and 4.40 respectively. How much  $HI$  must be added to the equilibrium concentrations to increase the concentration of  $I_2$  to 2.00 M?  $K_{eq}$  is equal to 6.0. The volume is 1.00 L.



$$x = -0.20$$

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

$$6.0 = \frac{(4.00+y)^2}{(2.00)^2}$$

$$y = \cancel{-8.899} \text{ or } 0.899$$

$\therefore$  0.899 moles of  $HI(g)$  must be added

# CALCULATIONS GIVEN $K_{eq}$

## Summarizing ICE Tables

1. Write out the balanced equation.
2. All values in the table must have mol/L units.
3. Initial [product] = 0, unless otherwise stated.
4. Changes in concentration always occur in the same stoichiometric ratio.
5. Reactants and products will change in opposite directions from each other.