

# EQUILIBRIUM

CALCULATIONS INVOLVING EQUILIBRIUM

CHEMISTRY 165 // SPRING 2020

## PRACTICE PROBLEM 1

A 1.00 L container holds 1.00 moles of  $\text{H}_2$  gas and 2.00 moles of  $\text{I}_2$  gas, which react to form HI gas.



Calculate the equilibrium concentrations of the  $\text{H}_2$ ,  $\text{I}_2$ , and HI gases.

— *answer* —

## PRACTICE PROBLEM 1

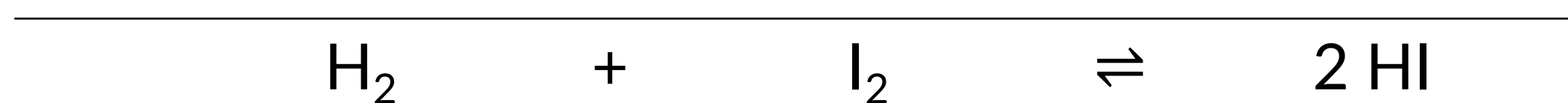
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Step 1: Write down the balanced chemical equation for the equilibrium.



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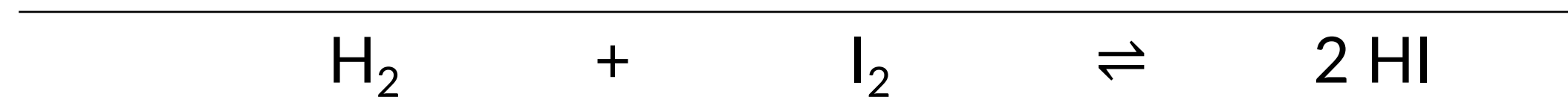
Calculate the equilibrium concentrations of the  $\text{H}_2$ ,  $\text{I}_2$ , and HI gases.

— *answer* —

Step 1: Write down the balanced chemical equation for the equilibrium.

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

Step 2: Write down the expression for the equilibrium constant.



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Step 1: Write down the balanced chemical equation for the equilibrium.

Step 2: Write down the expression for the equilibrium constant.

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

Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium)

	$\text{H}_2$	+	$\text{I}_2$	$\rightleftharpoons$	2 HI
I					
C					
E					

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A 1.00 L containers holds 1.00 moles of H<sub>2</sub> gas and 2.00 moles of I<sub>2</sub> gas, which react to form HI gas.



Calculate the equilibrium concentrations of the H<sub>2</sub>, I<sub>2</sub>, and HI gases.

— *answer* —

Step 1: Write down the balanced chemical equation for the equilibrium.

Step 2: Write down the expression for the equilibrium constant.

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

Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium)

List the initial concentrations here.



	H <sub>2</sub>	+	I <sub>2</sub>	⇌	2 HI
I	1.00 M		2.00 M		0
C					
E					

# PRACTICE PROBLEM 1

A 1.00 L container holds 1.00 moles of H<sub>2</sub> gas and 2.00 moles of I<sub>2</sub> gas, which react to form HI gas.



Calculate the equilibrium concentrations of the H<sub>2</sub>, I<sub>2</sub>, and HI gases.

— answer —

Step 1: Write down the balanced chemical equation for the equilibrium.

Step 2: Write down the expression for the equilibrium constant.

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

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		H <sub>2</sub>	+	I <sub>2</sub>	⇌	2 HI
List the initial concentrations here. →	I	1.00 M		2.00 M		0
List the stoichiometric changes here. →	C	- x		- x		+ 2x
	E					

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A 1.00 L containers holds 1.00 moles of H<sub>2</sub> gas and 2.00 moles of I<sub>2</sub> gas, which react to form HI gas.



Calculate the equilibrium concentrations of the H<sub>2</sub>, I<sub>2</sub>, and HI gases.

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Step 1: Write down the balanced chemical equation for the equilibrium.

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List the initial concentrations here. →	I	1.00 M		2.00 M		0
List the stoichiometric changes here. →	C	- x		- x		+ 2x ←
	E					

**How do I know what to put here?**

Because we start with only reactants ( $Q < K$ ), the reaction will shift right; hence we subtract some value "x" from our reactants and add some value "2x" to our product. The value of "x" will reflect the stoichiometry in the balanced equation; hence H<sub>2</sub> and I<sub>2</sub> decrease by "1x" and HI increases by "2x."



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A 1.00 L container holds 1.00 moles of H<sub>2</sub> gas and 2.00 moles of I<sub>2</sub> gas, which react to form HI gas.



Calculate the equilibrium concentrations of the H<sub>2</sub>, I<sub>2</sub>, and HI gases.

— answer —

Step 1: Write down the balanced chemical equation for the equilibrium.

Step 2: Write down the expression for the equilibrium constant.

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Because we start with only reactants ( $Q < K$ ), the reaction will shift right; hence we subtract some value "x" from our reactants and add some value "2x" to our product. The value of "x" will reflect the stoichiometry in the balanced equation; hence H<sub>2</sub> and I<sub>2</sub> decrease by "1x" and HI increases by "2x."

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Calculate the equilibrium concentrations of the H<sub>2</sub>, I<sub>2</sub>, and HI gases.

— answer —

Step 1: Write down the balanced chemical equation for the equilibrium.

Step 2: Write down the expression for the equilibrium constant.

Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium)

Step 4: Use the Equilibrium (E) values to plug into the K<sub>c</sub> expression.

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$
$$50.5 = \frac{(2x)^2}{(1.00 - x)(2.00 - x)}$$

		H <sub>2</sub>	+	I <sub>2</sub>	⇌	2 HI
List the initial concentrations here.	→	I	1.00 M	2.00 M		0
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Because we start with only reactants (Q < K), the reaction will shift right; hence we subtract some value "x" from our reactants and add some value "2x" to our product. The value of "x" will reflect the stoichiometry in the balanced equation; hence H<sub>2</sub> and I<sub>2</sub> decrease by "1x" and HI increases by "2x."

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Calculate the equilibrium concentrations of the H<sub>2</sub>, I<sub>2</sub>, and HI gases.

— answer —

Step 1: Write down the balanced chemical equation for the equilibrium.

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Step 5: Solve for “x” and for the equilibrium concentrations.

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$
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Calculate the equilibrium concentrations of the H<sub>2</sub>, I<sub>2</sub>, and HI gases.

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Step 1: Write down the balanced chemical equation for the equilibrium.

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$$x = 2.32$$

$$x = 0.935$$

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$x = 2.32$  ←

$x = 0.935$

Solving the quadratic formula gives two results for “x.” However,  $x \neq 2.32$  because that would give us negative values of concentrations at equilibrium! So  $x = 0.935$ .

List the initial concentrations here.



I

H<sub>2</sub> 1.00 M

+

I<sub>2</sub> 2.00 M

⇌

2 HI

0

List the stoichiometric changes here.



C

- x

- x

+ 2x



List the equilibrium values here.



E

1.00 - x

2.00 - x

2x

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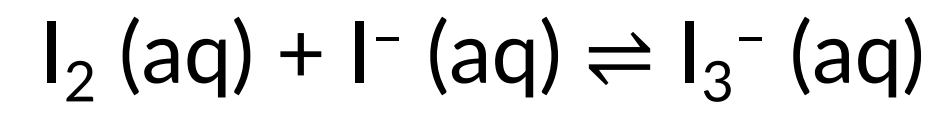
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$$[\text{H}_2]_{\text{eq}} = 1.00 \text{ M} - x = 1.00 \text{ M} - 0.935 \text{ M} = 0.07 \text{ M}$$

$$[\text{I}_2]_{\text{eq}} = 2.00 \text{ M} - x = 2.00 \text{ M} - 0.935 \text{ M} = 1.07 \text{ M}$$

$$[\text{HI}]_{\text{eq}} = 2x = 2(0.935 \text{ M}) = 1.87 \text{ M}$$

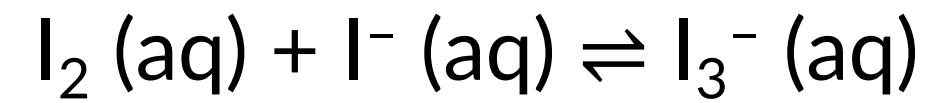
## PRACTICE PROBLEM 2



Iodine and iodide react to form triiodide ions. Assume you start with  $[\text{I}_2] = [\text{I}^-] = 1.000 \times 10^{-3} \text{ M}$ . If, at equilibrium, the concentration of  $\text{I}_2$  is  $6.61 \times 10^{-4} \text{ M}$ , what is the equilibrium constant for this reaction?

— *answer* —

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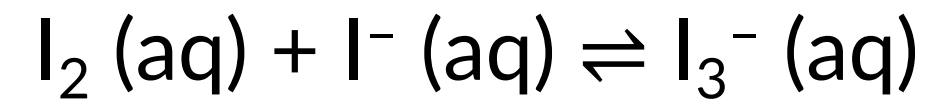
Step 2: Write down the expression for the equilibrium constant.

Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium)

	$\text{I}_2$	+	$\text{I}^-$	$\rightleftharpoons$	$\text{I}_3^-$
I					
C					
E					



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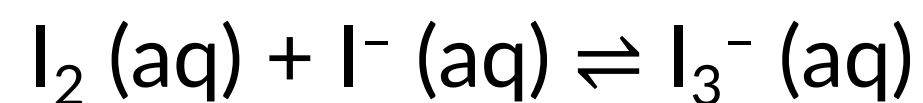
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	$\text{I}_2$	+	$\text{I}^-$	$\rightleftharpoons$	$\text{I}_3^-$
I	$1.000 \times 10^{-3} \text{ M}$		$1.000 \times 10^{-3}$		0
C	- x		- x		+ x
E	$1.000 \times 10^{-3} - x$		$1.000 \times 10^{-3} - x$		x

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Iodine and iodide react to form triiodide ions. Assume you start with  $[\text{I}_2] = [\text{I}^-] = 1.000 \times 10^{-3} \text{ M}$ . If, at equilibrium, the concentration of  $\text{I}_2$  is  $6.61 \times 10^{-4} \text{ M}$ , what is the equilibrium constant for this reaction?

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Step 2: Write down the expression for the equilibrium constant.

Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium)

### What is “x” though?

They tell us that the equilibrium concentration of  $\text{I}_2$  is  $6.61 \times 10^{-4} \text{ M}$ .

This means that:

$$[\text{I}_2]_{\text{eq}} = 1.000 \times 10^{-3} - x$$

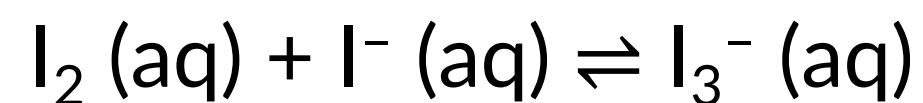
$$6.61 \times 10^{-4} \text{ M} = 1.000 \times 10^{-3} - x$$

$$x = 3.39 \times 10^{-4} \text{ M}$$

So, let's re-write the ICE chart with the value of x plugged in.

	$\text{I}_2$	+	$\text{I}^-$	$\rightleftharpoons$	$\text{I}_3^-$
I	$1.000 \times 10^{-3} \text{ M}$		$1.000 \times 10^{-3}$		0
C	- x		- x		+ x
E	$1.000 \times 10^{-3} - x$		$1.000 \times 10^{-3} - x$		x

## PRACTICE PROBLEM 2



Iodine and iodide react to form triiodide ions. Assume you start with  $[\text{I}_2] = [\text{I}^-] = 1.000 \times 10^{-3} \text{ M}$ . If, at equilibrium, the concentration of  $\text{I}_2$  is  $6.61 \times 10^{-4} \text{ M}$ , what is the equilibrium constant for this reaction?

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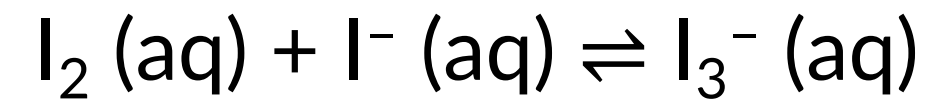
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	$\text{I}_2$	+	$\text{I}^-$	$\rightleftharpoons$	$\text{I}_3^-$
I	$1.000 \times 10^{-3} \text{ M}$		$1.000 \times 10^{-3}$		0
C	- x		- x		+ x
E	$1.000 \times 10^{-3} - x$		$1.000 \times 10^{-3} - x$		x
	$\text{I}_2$	+	$\text{I}^-$	$\rightleftharpoons$	$\text{I}_3^-$
I	$1.000 \times 10^{-3} \text{ M}$		$1.000 \times 10^{-3}$		0
C	$- 3.39 \times 10^{-4}$		$- 3.39 \times 10^{-4}$		$+ 3.39 \times 10^{-4}$
E	$6.61 \times 10^{-4}$		$6.61 \times 10^{-4}$		$3.39 \times 10^{-4}$

## PRACTICE PROBLEM 2



Iodine and iodide react to form triiodide ions. Assume you start with  $[\text{I}_2] = [\text{I}^-] = 1.000 \times 10^{-3} \text{ M}$ . If, at equilibrium, the concentration of  $\text{I}_2$  is  $6.61 \times 10^{-4} \text{ M}$ , what is the equilibrium constant for this reaction?

— answer —

Step 1: Write down the balanced chemical equation for the equilibrium.

Step 2: Write down the expression for the equilibrium constant.

Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium)

Step 4: Use the Equilibrium (E) values to plug into the  $K_c$  expression.

$$\begin{aligned} K_c &= \frac{[\text{I}_3^-]}{[\text{I}_2][\text{I}^-]} \\ &= \frac{3.39 \times 10^{-4}}{(6.61 \times 10^{-4})(6.61 \times 10^{-4})} \\ K_c &= 776 \end{aligned}$$

### What is “x” though?

They tell us that the equilibrium concentration of  $\text{I}_2$  is  $6.61 \times 10^{-4} \text{ M}$ . This means that:

$$\begin{aligned} [\text{I}_2]_{\text{eq}} &= 1.000 \times 10^{-3} - x \\ 6.61 \times 10^{-4} \text{ M} &= 1.000 \times 10^{-3} - x \\ x &= 3.39 \times 10^{-4} \text{ M} \end{aligned}$$

So, let’s re-write the ICE chart with the value of x plugged in.

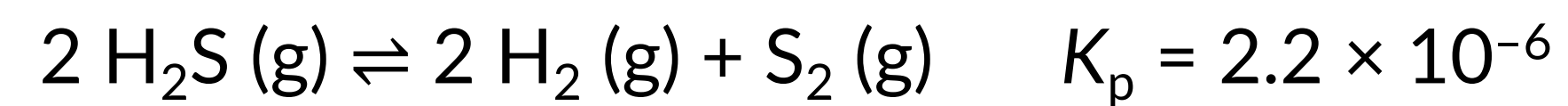
	$\text{I}_2$	+	$\text{I}^-$	$\rightleftharpoons$	$\text{I}_3^-$
I	$1.000 \times 10^{-3} \text{ M}$		$1.000 \times 10^{-3}$		0
C	- x		- x		+ x
E	$1.000 \times 10^{-3} - x$		$1.000 \times 10^{-3} - x$		x

	$\text{I}_2$	+	$\text{I}^-$	$\rightleftharpoons$	$\text{I}_3^-$
I	$1.000 \times 10^{-3} \text{ M}$		$1.000 \times 10^{-3}$		0
C	$- 3.39 \times 10^{-4}$		$- 3.39 \times 10^{-4}$		$+ 3.39 \times 10^{-4}$
E	$6.61 \times 10^{-4}$		$6.61 \times 10^{-4}$		$3.39 \times 10^{-4}$

## PRACTICE PROBLEM 3

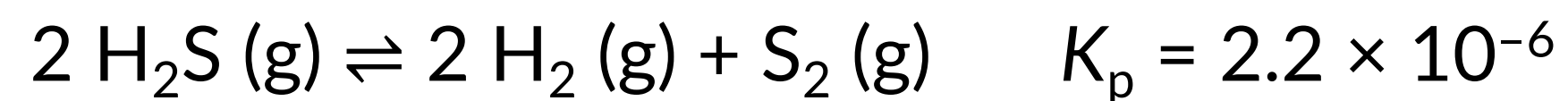
Calculate the equilibrium pressure of S<sub>2</sub> gas in an equilibrium mixture that results from the decomposition of H<sub>2</sub>S gas with an initial concentration of 0.824 atm.



— *answer* —

## PRACTICE PROBLEM 3

Calculate the equilibrium pressure of  $S_2$  gas in an equilibrium mixture that results from the decomposition of  $H_2S$  gas with an initial concentration of 0.824 atm.



— answer —

Step 1: Write down the balanced chemical equation for the equilibrium.

Step 2: Write down the expression for the equilibrium constant.

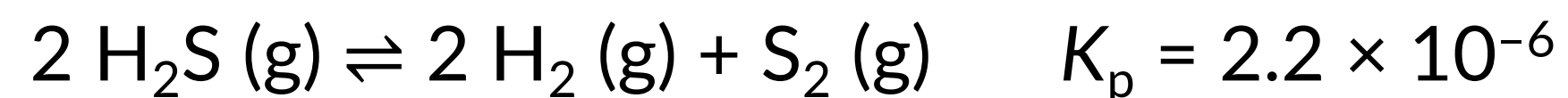
$$K_p = \frac{P_{H_2}^2 P_{S_2}}{P_{H_2S}^2}$$

Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium)

	$2 H_2S$	$\rightleftharpoons$	$2 H_2$	+	$S_2$
I					
C					
E					

## PRACTICE PROBLEM 3

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Step 2: Write down the expression for the equilibrium constant.

Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium)

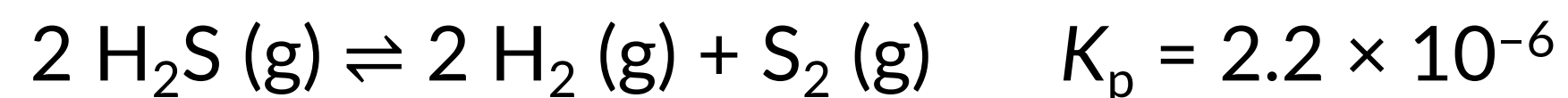
Step 4: Use the Equilibrium (E) values to plug into the  $K_p$  expression.

$$K_p = \frac{P_{H_2}^2 P_{S_2}}{P_{H_2S}^2}$$
$$2.2 \times 10^{-6} = \frac{(2x)^2 (x)}{(0.824 - 2x)^2}$$

	$2 H_2S$	$\rightleftharpoons$	$2 H_2$	+	$S_2$
I	0.824 atm		0		0
C	- 2x		+ 2x		+ x
E	0.824 - 2x		2x		x

## PRACTICE PROBLEM 3

Calculate the equilibrium pressure of S<sub>2</sub> gas in an equilibrium mixture that results from the decomposition of H<sub>2</sub>S gas with an initial concentration of 0.824 atm.



— answer —

Step 1: Write down the balanced chemical equation for the equilibrium.

Step 2: Write down the expression for the equilibrium constant.

Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium)

Step 4: Use the Equilibrium (E) values to plug into the K<sub>p</sub> expression.

$$K_p = \frac{P_{\text{H}_2}^2 P_{\text{S}_2}}{P_{\text{H}_2\text{S}}^2}$$
$$2.2 \times 10^{-6} = \frac{(2x)^2 (x)}{(0.824 - 2x)^2}$$

	2 H <sub>2</sub> S	⇌	2 H <sub>2</sub>	+	S <sub>2</sub>
I	0.824 atm		0		0
C	- 2x		+ 2x		+ x
E	0.824 - 2x		2x		x

What is “x” though?

We know K<sub>p</sub> is very small (~10<sup>-6</sup>), meaning that our equilibrium lies very far to the left (i.e. very little decomposition).

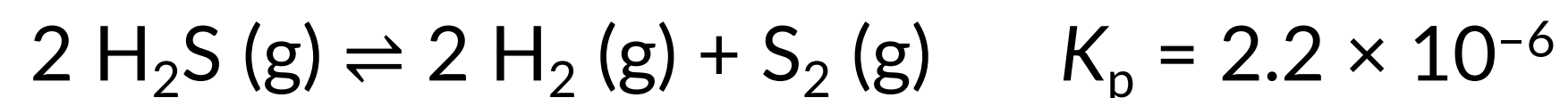
What this means for us is that our system will have to undergo very little change (the “x” value) in order to reach equilibrium. As such, we say x is small and can invoke the approximation that:

$$0.824 - 2x \approx 0.824 - 2(0) \approx 0.824$$



## PRACTICE PROBLEM 3

Calculate the equilibrium pressure of S<sub>2</sub> gas in an equilibrium mixture that results from the decomposition of H<sub>2</sub>S gas with an initial concentration of 0.824 atm.



— answer —

Step 1: Write down the balanced chemical equation for the equilibrium.

Step 2: Write down the expression for the equilibrium constant.

Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium)

Step 4: Use the Equilibrium (E) values to plug into the K<sub>p</sub> expression.

Step 5: Approximate x to be small, solve for x and partial pressure of S<sub>2</sub> gas.

$$K_p = \frac{P_{\text{H}_2}^2 P_{\text{S}_2}}{P_{\text{H}_2\text{S}}^2}$$
$$2.2 \times 10^{-6} = \frac{(2x)^2 (x)}{(0.824 - 2x)^2}$$

	2 H <sub>2</sub> S	⇌	2 H <sub>2</sub>	+	S <sub>2</sub>
I	0.824 atm		0		0
C	- 2x		+ 2x		+ x
E	0.824 - 2x		2x		x

What is “x” though?

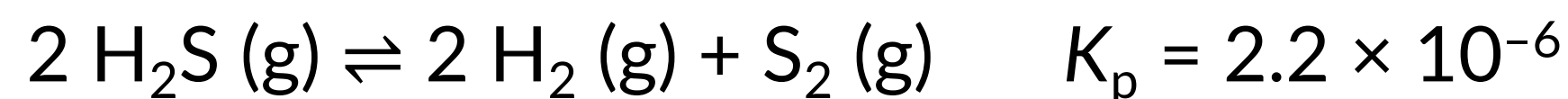
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## PRACTICE PROBLEM 3

Calculate the equilibrium pressure of  $S_2$  gas in an equilibrium mixture that results from the decomposition of  $H_2S$  gas with an initial concentration of 0.824 atm.



— answer —

Step 1: Write down the balanced chemical equation for the equilibrium.

Step 2: Write down the expression for the equilibrium constant.

Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium)

Step 4: Use the Equilibrium (E) values to plug into the  $K_p$  expression.

Step 5: Approximate  $x$  to be small, solve for  $x$  and partial pressure of  $S_2$  gas.

	$2 H_2S$	$\rightleftharpoons$	$2 H_2$	+	$S_2$
I	0.824 atm		0		0
C	- 2x		+ 2x		+ x
E	0.824 - 2x		2x		x

$$K_p = \frac{P_{H_2}^2 P_{S_2}}{P_{H_2S}^2}$$

$$2.2 \times 10^{-6} = \frac{(2x)^2 (x)}{(0.824 - 2x)^2}$$

$$2.2 \times 10^{-6} \approx \frac{(2x)^2 (x)}{(0.824)^2}$$

What is “x” though?

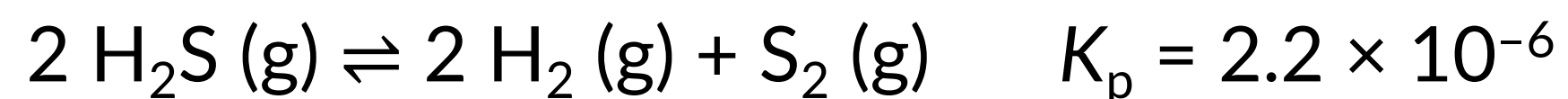
We know  $K_p$  is very small ( $\sim 10^{-6}$ ), meaning that our equilibrium lies very far to the left (i.e. very little decomposition).

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## PRACTICE PROBLEM 3

Calculate the equilibrium pressure of S<sub>2</sub> gas in an equilibrium mixture that results from the decomposition of H<sub>2</sub>S gas with an initial concentration of 0.824 atm.



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Step 1: Write down the balanced chemical equation for the equilibrium.

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Step 5: Approximate x to be small, solve for x and partial pressure of S<sub>2</sub> gas.

	2 H <sub>2</sub> S	⇌	2 H <sub>2</sub>	+	S <sub>2</sub>
I	0.824 atm		0		0
C	- 2x		+ 2x		+ x
E	0.824 - 2x		2x		x

$$K_p = \frac{P_{\text{H}_2}^2 P_{\text{S}_2}}{P_{\text{H}_2\text{S}}^2}$$

$$2.2 \times 10^{-6} = \frac{(2x)^2 (x)}{(0.824 - 2x)^2}$$

$$2.2 \times 10^{-6} \approx \frac{(2x)^2 (x)}{(0.824)^2}$$

$$2.2 \times 10^{-6} = \frac{4x^3}{(0.824)^2}$$

$$x = 0.0072 \text{ atm} = P_{\text{S}_2}$$

What is “x” though?

We know K<sub>p</sub> is very small (~10<sup>-6</sup>), meaning that our equilibrium lies very far to the left (i.e. very little decomposition).

What this means for us is that our system will have to undergo very little change (the “x” value) in order to reach equilibrium. As such, we say x is small and can invoke the approximation that:

$$0.824 - 2x \approx 0.824 - 2(0) \approx 0.824$$