EQUILIBRIUM

CALCULATIONS INVOLVING EQUILIBRIUM

DR. MIOY T. HUYNH | 2020

CHEMISTRY 165 // SPRING 2020

PRACTICE PROBLEM 1

A 1.00 L containers holds 1.00 moles of H_2 gas and 2.00 moles of I_2 gas, which react to form HI gas.

$$H_2\,(g) + I_2\,(g) \rightleftharpoons 2\,HI\,(g) \qquad K_c = 50.5 \text{ at } 298\text{ K}$$

Calculate the equilibrium concentrations of the H_2, I_2, and HI gases.

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

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I					
C					
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List the initial concentrations here. → I

List the stoichiometric changes here. → C

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_2 and I_2 decrease by "1x" and HI increases by "2x."

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How do I know what to put here?

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

$$50.5 = \frac{(2x)^2}{(1.00 - x)(2.00 - x)}$$

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

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$$[H_2]_{eq} = 1.00\text{ M} - x = 1.00\text{ M} - 0.935\text{ M} = 0.07\text{ M}$$

$$[I_2]_{eq} = 2.00\text{ M} - x = 2.00\text{ M} - 0.935\text{ M} = 1.07\text{ M}$$

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

PRACTICE PROBLEM 2

$$I_2\ (aq) + I^-\ (aq) \rightleftharpoons I_3^-\ (aq)$$

Iodine and iodide react to form triiodide ions. Assume you start with $[I_2] = [I^-] = 1.000 \times 10^{-3}$ M. If, at equilibrium, the concentration of I_2 is 6.61×10^{-4} M, what is the equilibrium constant for this reaction?

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C	$-\ x$		$-\ x$		$+\ x$
E	$1.000 \times 10^{-3} - x$		$1.000 \times 10^{-3} - x$		x

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What is "x" though?

They tell us that the equilibrium concentration of I_2 is 6.61×10^{-4} M. This means that:

$$[I_2]_{eq} = 1.000\times10^{-3} - x$$
$$6.61\times10^{-4}\ M = 1.000\times10^{-3} - x$$
$$x = 3.39\times10^{-4}\ M$$

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

	I_2	+	I^-	$\rightleftharpoons$	I_3^-
I	1.000×10^{-3} M		1.000×10^{-3}		0
C	– x		– x		+ x
E	1.000×10^{-3} – x		1.000×10^{-3} – x		x

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C	– x		– x		+ x
E	1.000×10^{-3} – x		1.000×10^{-3} – x		x

	I_2	+	I^-	$\rightleftharpoons$	I_3^-
I	1.000×10^{-3} M		1.000×10^{-3}		0
C	$- 3.39 \times 10^{-4}$		$- 3.39 \times 10^{-4}$		$+ 3.39 \times 10^{-4}$
E	6.61×10^{-4}		6.61×10^{-4}		3.39×10^{-4}

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$$K_c = \frac{[I_3^-]}{[I_2][I^-]}$$

$$= \frac{3.39\times10^{-4}}{(6.61\times10^{-4})(6.61\times10^{-4})}$$

$$K_c = 776$$

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

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C	– x		– x		+ x
E	1.000×10^{-3} – x		1.000×10^{-3} – x		x

	I_2	+	I^-	$\rightleftharpoons$	I_3^-
I	1.000×10^{-3} M		1.000×10^{-3}		0
C	– 3.39×10^{-4}		– 3.39×10^{-4}		+ 3.39×10^{-4}
E	6.61×10^{-4}		6.61×10^{-4}		3.39×10^{-4}

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.

$$2\ H_2S\ (g) \rightleftharpoons 2\ H_2\ (g) + S_2\ (g) \qquad K_p = 2.2 \times 10^{-6}$$

— answer —

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— answer —

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

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

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	$2\ H_2S$	$\rightleftharpoons$	$2\ H_2$	+	S_2
I					
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Step 4: Use the Equilibrium (E) values to plug into the K_p expression.

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

$$2.2\times10^{-6} = \frac{(2x)^2(x)}{(0.824 - 2x)^2}$$

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

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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.

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I	0.824 atm		0		0
C	– 2x		+ 2x		+ x
E	0.824 – 2x		2x		x

What is "x" though?

We know K_p is very small (~10^{-6}), 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:

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

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

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

$$2.2\times10^{-6} = \frac{(2x)^2(x)}{(0.824 - 2x)^2}$$

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

What is "x" though?

We know K_p is very small (~10^{-6}), 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_2 gas in an equilibrium mixture that results from the decomposition of H_2S gas with an initial concentration of 0.824 atm.

$$2\ H_2S\ (g) \rightleftharpoons 2\ H_2\ (g) + S_2\ (g) \qquad K_p = 2.2 \times 10^{-6}$$

— 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.

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

$$2.2\times10^{-6} = \frac{(2x)^2(x)}{(0.824 - 2x)^2}$$

$$2.2\times10^{-6} \approx \frac{(2x)^2(x)}{(0.824)^2}$$

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C	– 2x		+ 2x		+ x
E	0.824 – 2x		2x		x

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

$$2.2\times10^{-6} = \frac{(2x)^2(x)}{(0.824 - 2x)^2}$$

$$2.2\times10^{-6} \approx \frac{(2x)^2(x)}{(0.824)^2}$$

$$2.2\times10^{-6} = \frac{4x^3}{(0.824)^2}$$

$$x = 0.0072\ \text{atm} = P_{S_2}$$

What is "x" though?

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