# EQUILIBRIUM

CHEMISTRY 165 // SPRING 2020





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

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2 HI  $I_2$  $\neq$ 

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	$H_2$	+
	_	
С		
Е		

2 HI  $\Rightarrow$  $I_2$ 

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 at 298 K

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nt.

- I <sub>2</sub>	 2 HI	
2.00 M	0	

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-	<b>I</b> <sub>2</sub>	⇒	2 HI	
	2.00 M		0	
	- x		+ 2x	

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F	$ _2 \rightleftharpoons$	2 HI	
2.0	0 M	0	
_	×	+ 2x◀	

How do I know what to put here? Because we start with 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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			_
<b>I</b> <sub>2</sub>	4	2 HI	
2.00 M		0	
- x		+ 2x <b>←</b>	- How do I know what to put h
2.00 – x		2x	Because we start with reactants (Q < K), the reaction shift right; hence we subtract value "x" from our reactants a some value "2x" to our product value of "x" will reflect stoichiometry in the back

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

ł	<b>I</b> <sub>2</sub>	4	2 HI	
	2.00 M		0	
	- x		+ 2x <b>←</b>	How do
	2.00 – x		2x	Because
			· · ·	reactant
				shift rig

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quilibrium.	[]	$HI]^2$	
nt.	$K_{\rm c} = \overline{[{\rm H}]}$	$_{2}][I_{2}]$	
ession.	$50.5 = \frac{1}{(1.5)}$ x = 2.3 x = 0.9	$(2x)^2$ $(2x)^2$ $(2.00)^2$	0 – x)
+ l <sub>2</sub>	<del>\_</del>	2 HI	
2.00 N	Л	0	
- X		+ 2x <b>←</b>	
2.00 –	Χ	2x	Because we start with reactants (Q < K), the reaction shift right; hence we subtract value "x" from our reactants an some value "2x" to our product value of "x" will reflect

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$$K_{\rm c} = 50.5$$
 at 298 K

	-		
nt.	$K_{\rm c} = \frac{[{\rm H}]}{[{\rm H}_2]}$	$\frac{II}{2}$	Solving the quadratic for gives two results for
ession.	$50.5 = \frac{1}{(1.0)}$ x = 2.3	$(2x)^2$ (2x)	However, $x \neq 2.32$ b that would give us needed (0 - x) values of concentration equilibrium! So $x = 0.935$
	x = 0.9	35	
· I <sub>2</sub>	$\rightleftharpoons$	2 HI	
2.00 M	1	0	
- X		+ 2x ←	— How do I know what to put he
2.00 –	Χ	2x	Because we start with reactants (Q < K), the reaction shift right; hence we subtract value "x" from our reactants a some value "2x" to our product value of "x" will reflect stoichiometry in the back

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$$[H_2]_{eq} = 1.00 \text{ M} - \text{x} = 1.00 \text{ M} - 0.93$$
$$[I_2]_{eq} = 2.00 \text{ M} - \text{x} = 2.00 \text{ M} - 0.933$$
$$[HI]_{eq} = 2\text{x} = 2(0.935 \text{ M}) = 2$$

(g) 
$$K_{\rm c} = 50.5$$
 at 298 K

uilibrium	F-			
umprum.	$K_{c} = \frac{\lfloor \mathbf{H} \rfloor}{ \lfloor \mathbf{H} \rfloor}$	$\frac{1}{1}$		Solving the quadratic form
nt.		2][I <sub>2</sub> ]		gives two results for However $x \neq 2.32$ bec
	505 =	$(2x)^2$		that would give us negative that would give the second se
scion	50.5 - (1.5)	(00 - x)(2.00)	(x - 0)	values of concentrations
:551011.	x = 2.3	2 -		equilibrium! So $x = 0.935$ .
	x = 0.9	35		
			_	
l <sub>2</sub>		2 HI	_	
2.00 N	1	0		
- x		+ 2x ◀━━	- How	do I know what to put here
2.00 -	X	2x	Beca	ause we start with
			reac shift	right: hence we subtract so
5 M - 0.07	' М		valu	e "x" from our reactants and
5 - 0.07	1•1		som	e value "2x" to our product.
M = 1.07	Μ		valu	e of "x" will reflect
.87 M			stoid	hiometry in the balar
			equa	Ation; hence $H_2$ and $I_2$ decre
			DY	IX and HI increases by "2x."





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

lodine 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<sup>-4</sup> M, what is the equilibrium constant for this reaction? - answer -



 $I_2(aq) +$ 

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

quilibrium. 
$$K_{c} = \frac{[I_{3}^{-}]}{[I_{2}][I^{-}]}$$
nt.

1	I —	<b></b>	—	
T			2	
	_		-3	



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nt.

ł	- [-	4	l <sub>3</sub> -
1	$1.000 \times 10^{-3}$		0
	- X		+ x
X	$1.000 \times 10^{-3} - 3$	×	Х



 $I_{2}(aq) +$ 

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What is "x" though? They tell us that the equilibrium concentration of  $I_2$  is 6.61 × 10<sup>-4</sup> M. This means that:  $[I_2]_{eq} = 1.000 \times 10^{-3} - x$  $6.61 \times 10^{-4} \text{ M} = 1.000 \times 10^{-3} - \text{x}$  $x = 3.39 \times 10^{-4} M$ 

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

 $1.000 \times 10^{-3}$  M - X  $1.000 \times 10^{-3} - x$ F

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

quilibrium. 
$$K_{c} = \frac{[I_{3}^{-}]}{[I_{2}][I^{-}]}$$
nt.

+	<b>–</b>	+	l <sub>3</sub> -	
	$1.000 \times 10^{-3}$		0	
	- X		+ x	
	$1.000 \times 10^{-3} - x$	7	Х	



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12  $1.000 \times 10^{-3}$  M - X  $1.000 \times 10^{-3} - x$ F **1**<sub>2</sub>  $1.000 \times 10^{-3}$  M  $-3.39 \times 10^{-4}$ С  $6.61 \times 10^{-4}$ Ε

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

quilibrium. 
$$K_{c} = \frac{[I_{3}^{-}]}{[I_{2}][I^{-}]}$$
nt.

	+	-	$\stackrel{\sim}{\leftarrow}$	l <sub>3</sub> -
		$1.000 \times 10^{-3}$		0
		- X		+ x
(		$1.000 \times 10^{-3} - x$		X
	Ŧ	—	<del>\</del>	l <sub>3</sub> -
		$1.000 \times 10^{-3}$		0
		- 3.39 × 10 <sup>-4</sup>		$+ 3.39 \times 10^{-4}$
		$6.61 \times 10^{-4}$		$3.39 \times 10^{-4}$



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

uilibrium.  

$$K_{c} = \frac{[I_{3}^{-}]}{[I_{2}][I^{-}]}$$
t.  

$$= \frac{3.39 \times 10^{-4}}{(6.61 \times 10^{-4})(6.61 \times 10^{-4})}$$
ssion  

$$K_{c} = 776$$

+	-	<del>\</del>	l <sub>3</sub> -
	$1.000 \times 10^{-3}$		0
- X			+ x
	$1.000 \times 10^{-3} - x$		X

+	-	+	l <sub>3</sub> -
	$1.000 \times 10^{-3}$		0
	- 3.39 × 10 <sup>-4</sup>		$+ 3.39 \times 10^{-4}$
	6.61 × 10 <sup>-4</sup>		3.39 × 10 <sup>-4</sup>



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 -

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



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	$2 H_2S$	 2 H <sub>2</sub>	+	
I				
С				
Е				

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$$K_{\rm p} = \frac{P_{\rm H_2}^2 P_{\rm S_2}}{P_{\rm H_2}^2 S}$$

 $S_2$ 



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	$2 H_2S$	$\rightleftharpoons$	2 H <sub>2</sub>	+	
I	0.824 atm		0		
С	- 2x		+ 2x		
Е	0.824 – 2x		2x		

$$K_{\rm p} = 2.2 \times 10^{-6}$$

nt.  

$$K_{\rm p} = \frac{P_{\rm H_2}^2 P_{\rm S_2}}{P_{\rm H_2S}^2}$$
ession.  

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

Х



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	$2 H_2S$	$\rightleftharpoons$	2 H <sub>2</sub>	+	
I	0.824 atm		0		
С	- 2x		+ 2x		
Е	0.824 – 2x		2x		

What is "x" though?

We know  $K_{\rm p}$  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:

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

$$K_{\rm p} = 2.2 \times 10^{-6}$$

nt.  

$$K_{\rm p} = \frac{P_{\rm H_2}^2 P_{\rm S_2}}{P_{\rm H_2S}^2}$$
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 $2 H_2 S (g) \rightleftharpoons 2 H_2 (g)$ 

- answer -

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	$2 H_2S$	<del>\</del>	2 H <sub>2</sub>	+	
Ι	0.824 atm		0		
С	- 2x		+ 2x		
Е	0.824 – 2x		2x		

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 $2 H_2 S (g) \rightleftharpoons 2 H_2 (g)$ 

- answer -

Step 1: Write down the balanced chemical equation for the equilibrium. Step 2: Write down the expression for the equilibrium constar Step 3: Prepare an ICE chart (Initial, Change, and Equilibrium) Step 4: Use the Equilibrium (E) values to plug into the  $K_p$  expre Step 5: Approximate x to be small, solve for x and partial press

	$2 H_2S$	<del>\</del>	2 H <sub>2</sub>	+	
Ι	0.824 atm		0		
С	- 2x		+ 2x		
Е	0.824 – 2x		2x		

What is "x" though?

We know  $K_{\rm p}$  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:

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

$$K_{\rm p} = 2.2 \times 10^{-6}$$

nt.  

$$K_{p} = \frac{P_{H_{2}}^{2} P_{S_{2}}}{P_{H_{2}S}^{2}}$$
ession.  
sure of S<sub>2</sub> gas.  

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

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

+ X

0

Х





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.

 $2 H_2 S(g) \rightleftharpoons 2 H_2(g) + S_2(g)$   $K_n = 2.2 \times 10^{-6}$ 

- answer -

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

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

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

$$x = 0.0072 \text{ atm} = P_{S_{2}}$$



