

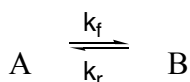
## Ch. 14 Chemical Equilibrium

## 14.1 Characteristics of a System at Chemical Equilibrium

- Not completely on product or reactant side – have some of each
  - “product favored” – mostly product
  - “reactant favored” – mostly reactant
- “Dynamic Equilibrium”
  - forward rate = reverse rate
  - reactions happen, but cancel each other out, so that no overall change
    - “steady state”
- Previous Examples
  - Vapor Pressure (closed container)      Liquid  $\rightleftharpoons$  vapor
  - Saturated solution                              solid  $\rightleftharpoons$  dissolved
  - Weak acid/weak electrolyte              HF  $\rightleftharpoons$  H<sup>+</sup> + F<sup>⊖</sup>
- In achieving an equilibrium it doesn't matter whether you start from left side or right side
  - Either way, you end up with the same balance between reactants and products once equilibrium has been established
  - Figure 14.1 Moore, Fig T-117, 118 Brown
- Catalysts Don't Change The Equilibrium Situation
  - they just help you to achieve it faster
- A Change in Temperature does change the equilibrium steady state

## 14.2 The Equilibrium “Constant”

## A. Intro

K<sub>f</sub>

At equilibrium, rate <sub>forward</sub> = rate <sub>reverse</sub>
---

Rate forward:

Rate reverse:

Equality:

Rearrange:

Notes

- K<sub>c</sub> is a constant
- Each reaction has its own unique “K” constant
  - called the “EQUILIBRIUM CONSTANT”
- Ratio of **product over reactant**
  - concentrations in “M” (Molarity) = mole/liter

Qualitative things the value of K tells us:

- At equilibrium, is the product favored or the reactant?
  - $K \gg 1 \rightarrow$  product favored ( $K \gg 1$ )
  - $K \ll 1 \rightarrow$  reactant favored ( $K \gg 1$ )

**Concept Problem:**  $A \rightleftharpoons B$ Find  $[B]/[A]$  ratios if:

- $K_c = 10^3$
  - $K_c = 10^{-3}$
- If the actual initial ratio Q does not equal the equilibrium ratio K, in which direction will reaction go to achieve equilibrium? (In other words, if you start out not at equilibrium, how will the system respond to get to equilibrium?)
    - If actual ratio "Q" < K, the product/reactant ratio will need to increase in order to reach the target "K"
      - Practical:
    - If actual ratio "Q" > K, the product/reactant ratio will need to decrease in order to reach the target "K"
      - Practical:

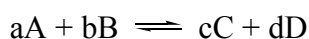
**Concept Problem:**  $A \rightleftharpoons B$  $K_c = 0.20$ How will  $[A]$  and  $[B]$  change/respond to reach equilibrium (which will go up and which will get down?) if initial conditions are:

- $[A] = 0.10 \text{ M}$  and  $[B] = 0.00 \text{ M}$
  - $[A] = 0.10 \text{ M}$  and  $[B] = 0.40 \text{ M}$
- What will specific concentrations be at equilibrium?
    - If you know K and some other initial information, you can solve for all the equilibrium concentrations

**Concept Problem:**  $A \rightleftharpoons B$  $K_c = 0.20$ 

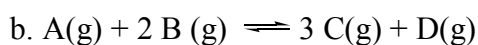
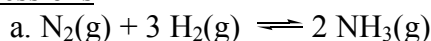
- If the system is at equilibrium and  $[A] = 0.10 \text{ M}$ , what must be  $[B]$ ?
- If the system is at equilibrium and  $[B] = 0.10 \text{ M}$ , what must be  $[A]$ ?
- If the system is **not** at equilibrium, and the initial scenario has  $[B] = 0.50 \text{ M}$ , what will be the concentrations after equilibrium is achieved?

## B. Writing K expressions: General



$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

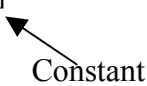
1. Products over reactants
2. To the "coefficient" power
3. Straight from the balanced equation
  - don't need mechanism, etc.
  - Not like a rate law where equation doesn't really tell you anything

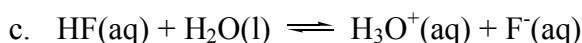
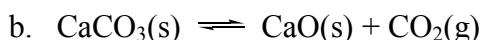
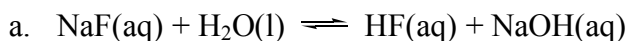
Write  $K_c$  ExpressionsC. Heterogeneous Equilibria: Do not include solids or liquids (solvents) in K expressions

- The concentrations of solids and liquids (solvents) are constant, so just wrap into K
- only gases and "aqueous things" (solutes) appear in K expressions
  - these are things whose concentrations are **not** constant
  - A(g), B(aq) Yes
  - C(l), D(s) No
- When water is involved as either a reactant or a product, it does **not** appear in the K expression (when the reaction takes place in aqueous solvent)



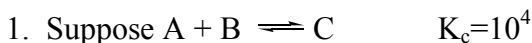
$$K_1 = \frac{[Cu^+][OH^-]}{[CuOH]} \quad \text{so} \quad K_2 = [Cu^+][OH^-]$$


  
Constant

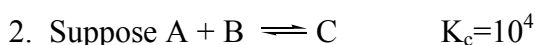
Write K Expressions

## D. K's for Related Reactions

- When a reaction is written in the reverse direction, the K value is inverted
  - $K_{\text{reverse}} = 1/K_{\text{forward}}$
- When the coefficients for a reaction are doubled or halved, the K value ends up being squared or “square rooted”



What is K for  $C \rightleftharpoons A + B$   
 (Draw both the expression and  
 the actual K value)



What is K for  $2A + 2B \rightleftharpoons 2C$   
 (Draw both the expression and  
 the actual K value)

## 14.3 Calculating K

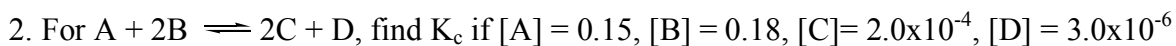
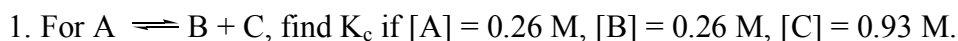
- Key: You need all equilibrium concentrations. Then you can plug into K expression and solve.

Two ways to know all the equilibrium concentrations.

1. You are simply given all of the equilibrium concentrations. (easy)
2. You are given all of the **initial** concentrations, and **at least one final concentration**, but then must use the “ICE” (**I**nitial-**C**hange-**E**quilibrium) method to figure out what they would all be at equilibrium (harder)

A. Given all of the Equilibrium concentrations, solve for K (easy)Solution steps:

- a. Based on the balanced reaction, write out the correct K expression
- b. Enter the known concentrations values and solve for K



B. Given all of the all of the **initial** concentrations, and **at least one final concentration**, solve for K (harder).

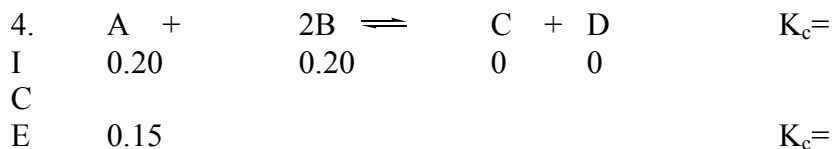
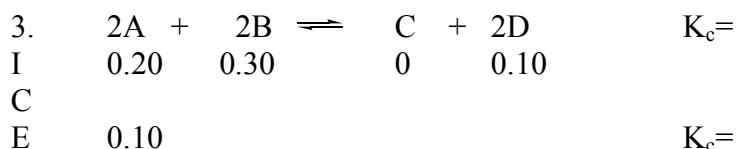
Solution steps:

- Based on the balanced reaction, write out the correct K expression
- Use the "ICE" (**I**nitial-**C**hange-**E**quilibrium) method to figure out the equilibrium concentrations
- Enter the equilibrium concentration values and solve for K

**Using the "ICE" Method to Calculate K, Given Initial Concentrations and One Final Concentration**

- Write balanced equation, and expression for  $K_c$
- Make an "ICE" table, and enter the knowns
  - Initial
  - Change
  - Equilibrium
- Find the Change for the chemical whose final concentration is known
- Use stoichiometric relationship to determine the change in concentrations for the others
- From the initial concentrations and the deduced changes, determine all equilibrium concentrations
- With all equilibrium concentrations now known, plug into the  $K_c$  expression and solve for K
- Check: Does Answer Make Any Sense?
  - Note: equilibrium concentrations must be in Molarity, moles/liter. If information is given in grams or moles plus solvent volume, you will need to convert into molarity.

Use the ICE method to find the equilibrium concentrations and the K values for the following:



#### 14.4 Chemical Meaning of K

- $K_c \gg 1$  Product Favored
  - $K_c \ll 1$  Reactant Favored
  - $K_c \approx 1$  Significant quantities of both products and reactants---calculations required
- There are many cases where  $K_c \ll 1$ , but small amount of product are still very important.

14.5 Using  $K_c$ 

## A. Predicting the direction of a reaction, given information about actual situation

$K$ =equilibrium       $Q$ =actual ratios(same format)

- If  $Q=K$ , you're at equilibrium. The situation won't change
- If  $Q<K$ , you're not at equilibrium. You are short on product, so reactions will go to products to achieve equilibrium. Shift from left to right.
- If  $Q>K$ , you're not at equilibrium. You have excess product, so products will go to reactants to achieve equilibrium. Shift from right to left.

1.  $A + B \rightleftharpoons C$        $K_c = 72$   
 Suppose  $[A]$ ,  $[B]$ , and  $[C]$  all begin at 0.33 M:

- What is  $Q$ ?
- Which direction will the reaction go to achieve equilibrium?
- What will happen to the concentrations of: (will they go up or down?)

$[A]$                        $[B]$                        $[C]$

**B. Calculating Equilibrium Concentrations, Given K and Some Other Information**

1. If all but one equilibrium concentration is given: (easy, plug and chug...)

Solution steps:

- Based on the balanced reaction, write out the correct  $K$  expression
- Enter the known  $K$  and the known concentrations values and solve for the unknown concentration

Example.       $HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$        $K_c = 6.8 \times 10^{-4}$

a. Find  $[H^+]$  if  $[HF] = 0.10$  and  $[F^-] = 0.10$ .

b. Find  $[F^-]$  if  $[HF] = 0.20$  and  $[H^+] = 10^{-3}$ .

2. If only initial concentrations are given, use ICE

### Using the “ICE” Method to Calculate Equilibrium Concentrations, Given Only Initial Concentrations and $K_c$

1. Write balanced equation, and expression for  $K_c$
2. Make an “ICE” table, and enter the known initial concentrations
  - a. Initial
  - b. Change
  - c. Equilibrium
3. Use “x” to define the change of one substance.
4. Use stoichiometric relationships to determine the changes in the concentrations for the others, in terms of “x”.
5. Calculate the equilibrium concentrations of all chemicals in terms of initial concentrations and “x”, and enter them in the table.
  - Ex:  $0.30 - x$ , or  $0.30 - 2x$ , or  $0.00 + x$ , or  $0.00 + 2x$  .....
  - **If  $K$  is small** so that “x” is likely to be small, **use the simplifying assumption** that  $[A]_{\text{initial}} - “x” = [A]_{\text{initial}}$ 
    - This is often justified, and can greatly simplify the math.
    - Ex:  $0.20 - x = 0.20$  if x is smaller than 0.01
6. Solve for “x” (This is the hard part!)
7. **Once “x” is known, use it to solve for the actual equilibrium concentrations**
8. Check: Does Answer Make Any Sense?
9. Check: If you made the “simplifying assumption”, was it justified? (Was “x”  $< 5\%$  of  $[A]_{\text{initial}}$ ?)

Note: Some hard math!! Quadratics are tedious!

- If possible, being able to avoid needing to solve quadratics saves a lot of time (and mistakes)

### \*Great simplifying assumption (avoids a lot of quadratics):

1. changes “x” are often small
  - For something with nonzero original concentration, the change “x” in concentration is often so small that it doesn’t significantly change the original concentration
  - $[\text{original concentration}] - “x” (\text{change in concentration}) = [\text{original concentration}]$
2. Useful strategy:
  - a. Make the simplifying assumption
  - b. Check at end whether “x”  $> 5\%$  of the original concentration
    - If not, the simplifying assumption is pretty good
    - If yes, the simplifying assumption is not valid and leads to some nontrivial error. In which case you need to go back and solve completely (often involving quadratic or perhaps even worse).

**Concept Example: Why is the Simplifying Assumption Often Justifiable?**  $0.20-x = ???$ 

<u>If "x" =</u>	<u>Then <math>0.20-x =</math></u>	<u>Is simplifying assumption valid?</u>
$2 \times 10^{-7}$	$0.20 - 0.0000002 =$	
$2 \times 10^{-6}$	$0.20 - 0.000002 =$	
$2 \times 10^{-5}$	$0.20 - 0.00002 =$	
$2 \times 10^{-4}$	$0.20 - 0.0002 =$	
$2 \times 10^{-3}$	$0.20 - 0.002 =$	
$2 \times 10^{-2}$	<b><math>0.20 - 0.02 =</math></b>	

**-Check at end. If  $x > 5\%$  of number, the assumption causes trouble.**

Problems

1. If 0.10 mol of  $N_2O_4$  is added to 1L flask, what will be the concentrations at equilibrium?

**Initial****Change****Equilibrium****Equilibrium with  
Simplifying Assumption**

Equilibrium:

$[N_2O_4]$

$[N_2O_4]$

2. Find final concentrations:

**Initial Concentrations**

0.20      0      0

**Change****Equilibrium****Equilibrium with  
Simplifying Assumption**

Equilibrium:

3. Find final concentrations given:



**Initial** 0.30

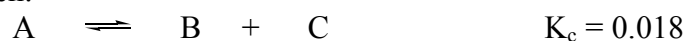
**Change**

**Equilibrium**

**Equilibrium with  
Simplifying Assumption**

Equilibrium:

4. Find final concentrations given:



**Initial** 0.30

**Change**

**Equilibrium**

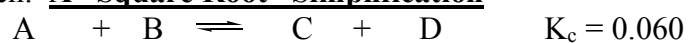
**Equilibrium with  
Simplifying Assumption**

Equilibrium:

[N<sub>2</sub>O<sub>4</sub>]

[N<sub>2</sub>O<sub>4</sub>]

5. Find final concentrations given: **A "Square Root" Simplification**



**Initial** 0.20

**Change**

**Equilibrium**

**Equilibrium with  
Simplifying Assumption**

Equilibrium:

6. Find final concentrations given: Solution with and Without Simplifying Assumption  
 $A \rightleftharpoons 2B + C$   $K_c = 2.0 \times 10^{-8}$

**Initial** 0.20

**Change**

**Equilibrium**

**Equilibrium with**

**Simplifying Assumption**

With Simplifying Assumption

Without Simplifying Assumption

Solution With and Without Simplifying Assumption:  $A \rightleftharpoons 2B + C$   $K_c = 2.0 \times 10^{-8}$

$A \rightleftharpoons 2B + C$				$A \rightleftharpoons 2B + C$ $K_c = 2.0 \times 10^{-8}$	
With I	0.20	0	0	Without	(using .20 - x for [A] at eq)
C	-x	2x	x		
E	.20-x	2x	x		
E*	0.20	x	x		
So	$K = 2.0 \times 10^{-8} = \frac{(2x)^2 x}{0.20} = \frac{4x^3}{.20}$				$K = 2.0 \times 10^{-8} = \frac{(2x)^2 x}{0.20 - x}$
	$0.4 \times 10^{-8} = 4x^3$				So $2.0 \times 10^{-8} (0.20 - x) = 4x^3$
	$x^3 = 1.0 \times 10^{-9}$				So $4.0 \times 10^{-9} - 2.0 \times 10^{-8} x = 4x^3$
	<b>Easy</b> $x = 1.0 \times 10^{-3}$				So $4x^3 + 2.0 \times 10^{-8} x - 4.0 \times 10^{-9} = 0$
					<b>Not Easy!!!</b>

## 14.6 How Systems Respond When an Equilibrium is Disturbed. LeChatelier's Principle.

1. At Equilibrium, there is a steady state of concentrations.
2. How does a system respond when an equilibrium is disturbed?

**Le Chatelier's Principle:** when a system is disturbed from equilibrium, it responds so as to reduce or counteract the effect of the change.

1. A revised equilibrium gets reestablished
  - a. Usually the same K value must still apply
    - unless the temperature is changed
  - b. But different actual concentrations result
    - whether you added or subtracted a reactant, a product, some solvent, or the volume
2. The reaction can move in the forward direction (more reactants go to product side) or the reverse direction (more products go back to starting materials) as needed to offset a change

Qual: If something extra is added, react from that side toward the other to remove some of it  
If something is removed, react from the other side to replace it

LeChateliers's Principle: Summary

A change in any of the factors that determine the equilibrium conditions of a system will cause the system to change in such a manner as to reduce or counteract the effect of the change.

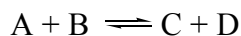
Factor that Changes	Direction of the Change	Direction to Restore Equilibrium
1. Concentration	↑ Reactant	Forward
	↑ Product	Reverse
2. Temperature	↑ T, when $\Delta H > 0$	Forward – K increases
	↑ T, when $\Delta H < 0$	Reverse – K decreases
3. Volume	↑ V, when $\Delta n_{\text{gas}} > 0$	Forward
	↑ V, when $\Delta n_{\text{gas}} = 0$	No effect
A change in volume impacts gas pressure. An increase in volume reduces pressure.	↑ V, when $\Delta n_{\text{gas}} < 0$	Reverse

If you reverse any of the Changes above, the **Direction to reach Equilibrium** reverses

3 Factors That can Be Changed 1) Concentrations 2) Temperature 3) Volume (which impacts gas pressure)	Goals 1) restore $Q = K$ 2) offset change
--	---

## A. Concentration Change: Addition or Removal of Either a Reactant or a Product

1. If you add (or remove) a chem, at that moment  $Q \neq K$
2. The system (and concentrations) will shift forward or backward to adjust  $Q$  so  $Q = K$ 
  - **NOTE: K value does not change**
  - many chemical ways to remove a chem (precipitations, acid-base, etc.)

Example

<u>Disruption</u>	<u>Immediate Impact on Q</u>	<u>Direction to Restore Eq</u>	Response of:			
			[A]	[B]	[C]	[D]
Add A	$Q < K$	Forward	↓	↓	↑	↑
Add B						
Add C						
Add D						
Remove A						
Remove B						
Remove C						
Remove D						

## B. Volume Change: Only Relevant When Gases are Involved

1. When the volume changes, it causes a change in concentrations
  - at the instant of change  $Q \neq K$
  - note: the actual  $K$  value is unchanged

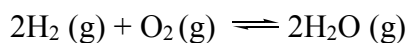
## 2. Qualitative sense:

- a. **When the volume expands, you want to generate more gas molecules to fill it**
- b. **When the volume shrinks, you want to get rid of gas molecules so it isn't too full**

3. Volume changes will only matter if one side of the equation has more gas on it than the other
  - if moles of gas are equal on both sides, volume changes have no impact

<u>Disruption</u>	<u>Response</u>
Volume Increase	Shift toward side with more gas
Volume Decrease	Shift toward side with less gas

4. Volume and Pressure: A different qualitative way to understand why a system will respond
  - When the volume expands, pressure increases
    - a. When the volume expands and the pressure drops, you want to generate more gas molecules to fill the container and restore pressure
    - b. When the volume shrinks, you want to get rid of gas molecules to unclutter the container and to restore the original pressure

Volume Example

<u>Volume Change</u>	<u>Direction to Restore Eq</u>	Response of moles of: $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{H}_2\text{O}$
Increase		
Decrease		

## C. Temperature Change

- View heat as either a reactant or a product
  - $\Delta H > 0$  Endothermic Heat is required, on Reactant Side
  - $\Delta H < 0$  Exothermic Heat is produced, on Product Side
- When the temperature rises, the system responds so as to remove extra heat
- When the temperature drops, the system responds so as to replace missing heat
- When the system shifts to the right or to the left in response to a temperature change, this means that chemical concentrations change, too

## 5. A Temperature Changes Causes a Changes in the K Value

- Because the Product/Reactant Ratio Changes
  - $\Delta H > 0$  Rise in temp  $\rightarrow$  shift to right  $\rightarrow$  K rises
  - $\Delta H > 0$  Drop in temp  $\rightarrow$  shift to left  $\rightarrow$  K drops
  - $\Delta H < 0$  Rise in temp  $\rightarrow$  shift to left  $\rightarrow$  K drops
  - $\Delta H < 0$  Drop in temp  $\rightarrow$  shift to right  $\rightarrow$  K rises

- Write in "heat" as reactant or product, and predict the responses when temperature changes:



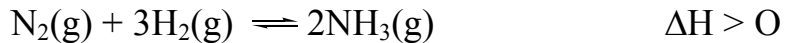
<u>Temperature Change</u>	<u>Direction to Restore Eq</u>	Response of moles of: $\text{A} \rightleftharpoons \text{B}$	Response of K Value
Increase			
Decrease			

- Write in "heat" as reactant or product, and predict the responses when temperature changes:



<u>Temperature Change</u>	<u>Direction to Restore Eq</u>	Response of moles of: $\text{A} \rightleftharpoons \text{B}$	Response of K Value
Increase			
Decrease			

3. Write in “heat” as reactant or product, and predict the responses to the changes:



<u>Disruption</u>	<u>Direction to Restore Eq</u>	Response of moles of: $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$	Response of K Value
1. Temperature Increase			
2. Volume Increase			
3. Added $\text{N}_2$			
4. Removal of $\text{NH}_3$			
5. Reduced temperature			
6. Added $\text{NH}_3$			
7. Removal of $\text{N}_2$			
8. Volume reduction			

D. Addition Catalyst: Equilibrium is reached faster, but the actual K value or equilibrium product/reactant ratios are unchanged.

14.7 LeChatelier's Principle on the Nanoscale. Skim. Equilibria favor more stable chemicals

**Some Math Tips: CHEMICAL EQUILIBRIA** Chem 210-Jasperse Ch. 14

**1.  $(a + b)^2 = a^2 + 2ab + b^2$**

example:  $(0.24 - x)^2 = 0.0576 - 0.48x + x^2 = x^2 - 0.48x + 0.0576$

example:  $(3.2e-5 - x)^2 = 1.02e-9 - 6.4e-5 x + x^2 = x^2 - 6.4e-5 x + 1.02e-9$

**2. Quadratic Equation:  $ax^2 + bx + c = 0$**

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

example:  $x^2 - 1.48x + 0.0576 = 0$        $x = 1.44$  or  $0.04$

example:  $x^2 - .60x + 0.025 = 0$        $x = 0.43$  or  $0.17$

Note 1: Quadratics are a bother to solve. Avoid them if possible. Often the “simplifying assumption” (see #4) can enable you to avoid quadratic solutions.

Note 2: Quadratics can often give 2 answers, but one will frequently be nonsense.

**3. “SIMPLIFYING ASSUMPTION”: If  $m \gg x$ , then  $m - x = m$  (approximately)**

examples:  $0.14 - 3.6e-5 = 0.14$

$1.23e-3 - 1.0e-6 = 1.23e-3$

Use: for many equations in which K is small, “x” will also be relatively small. In these circumstances, the degree to which the starting materials react (“x”) will be relatively insignificant from their concentrations perspective, and we can assume the equilibrium concentration of the reactants will not differ significantly from their initial concentration.

- THE SIMPLIFYING ASSUMPTION WILL OFTEN MAKE THE USE OF THE QUADRATIC EQUATION UNNEEDED
- Guide: If “x” is >5% of “m”, then the simplifying assumption is not appropriate.

**4. The “Square Root” Simplification: If both numerator and denominator have “squares”, it is convenient to take the square root of both sides to simplify**

examples:  $0.10 = x^2 / (0.050 - x)^2$  Take square root of both sides:  $0.316 = x / (0.050 - x)$   
 Rearrange:  $0.0158 - 0.316x = x$   
 Solve for “x”:  $x = 0.012$

**5.  $m = x^n$**

To solve for x, when “m” and “n” are known:

- enter “m”
- hit  $x^{1/y}$  or  $\sqrt[n]{\quad}$  button (depending on your calculator)
- enter “n”

example:

$25 = x^3$        $x = 2.92$

$1.48e-6 = x^3$        $x = 0.0114$

$2.14e-13 = x^4$        $x = 6.80e-4$

## Using the “ICE” Method to Calculate K, Given Initial Concentrations and One Final Concentration

1. Write balanced equation, and expression for  $K_c$
2. Make an “ICE” table, and enter the knowns
  - a. Initial
  - b. Change
  - c. Equilibrium
3. Find the Change for the chemical whose final concentration is known
4. Use stoichiometric relationship to determine the change in concentrations for the others
5. From the initial concentrations and the deduced changes, determine all equilibrium concentrations
6. With all equilibrium concentrations now known, plug into the  $K_c$  expression and solve for K
7. Check: Does Answer Make Any Sense?
  - Note: equilibrium concentrations must be in Molarity, moles/liter. If information is given in grams or moles plus solvent volume, you will need to convert into molarity.

## Using the “ICE” Method to Calculate Equilibrium Concentrations, Given Only Initial Concentrations and $K_c$

1. Write balanced equation, and expression for  $K_c$
2. Make an “ICE” table, and enter the known initial concentrations
  - a. Initial
  - b. Change
  - c. Equilibrium
3. Use “x” to define the change of one substance.
4. Use stoichiometric relationships to determine the changes in the concentrations for the others, in terms of “x”.
5. Calculate the equilibrium concentrations of all chemicals in terms of initial concentrations and “x”, and enter them in the table.
  - Ex:  $0.30 - x$ , or  $0.30 - 2x$ , or  $0.00 + x$ , or  $0.00 + 2x$  .....
  - **If K is small** so that “x” is likely to be small, **use the simplifying assumption** that  $[A]_{\text{initial}} - “x” = [A]_{\text{initial}}$ 
    - This is often justified, and can greatly simplify the math.
    - Ex:  $0.20 - x = 0.20$  if x is smaller than 0.01
6. Solve for “x” (This is the hard part!)
7. **Once “x” is known, use it to solve for the actual equilibrium concentrations**
8. Check: Does Answer Make Any Sense?
9. Check: If you made the “simplifying assumption”, was it justified? (Was “x”  $< 5\%$  of  $[A]_{\text{initial}}$ ?)

LeChateliers's Principle: Summary

A change in any of the factors that determine the equilibrium conditions of a system will cause the system to change in such a manner as to reduce or counteract the effect of the change.

Factor that Changes	Direction of the Change	Direction to Restore Equilibrium
1. Concentration	↑ Reactant	Forward
	↑ Product	Reverse
2. Temperature	↑ T, when $\Delta H > 0$	Forward – K increases
	↑ T, when $\Delta H < 0$	Reverse – K decreases
3. Volume	↑ V, when $\Delta n_{\text{gas}} > 0$	Forward
	↑ V, when $\Delta n_{\text{gas}} = 0$	No effect
A change in volume impacts gas pressure. An increase in volume reduces pressure.	↑ V, when $\Delta n_{\text{gas}} < 0$	Reverse

If you reverse any of the Changes above, the **Direction to reach Equilibrium** reverses