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Chemical Equilibrium

Chemistry in Two Directions

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Chemical Reactions

Up until now, we have talked about reactions as though they proceed in one direction: from left to right, as written in a balanced equation:

$$A_{(aq)} + 2 B_{(aq)} \rightarrow 6 C_{(aq)}$$

In fact, this is rarely true.

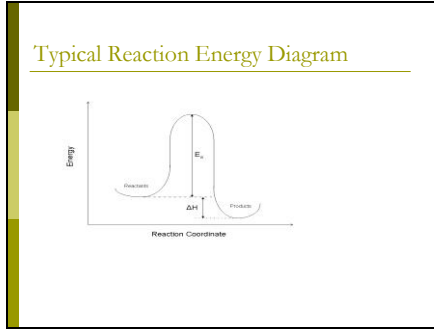
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If you can go left, you can turn around and go right

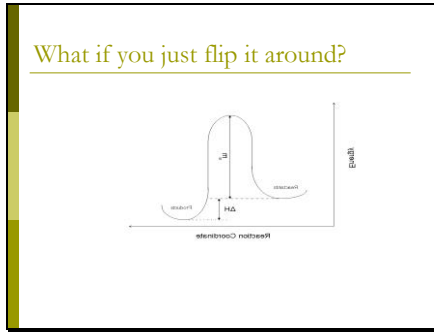
If you consider the energetics of a reaction, it is usually represented by a graph of Energy vs. Reaction Coordinate

(Reaction Coordinate is complicated, but you can think of it as the lowest energy path between chemical species)

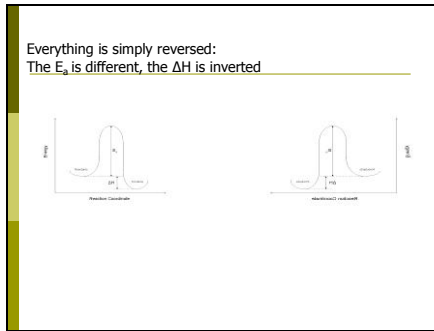
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Reactions can go both ways

- The hill is just higher going one way than the other.
- If it is exothermic going one way, it is endothermic going the other way.
- The world is reversible! (Well, except for time ☺)

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Implications for Chemical Reactions

The reversibility of chemical reactions means that rather than proceed from reactants to products, most reactions reach a state where there is no further change.

THIS DOES NOT MEAN NOTHING IS HAPPENING!

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DYNAMIC EQUILIBRIUM

Chemical Equilibrium is a **Dynamic Equilibrium**.

It is not that there is no reaction occurring, it is that the forward reaction and the reverse reaction are occurring at equal rates.

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$A_{(aq)} + 2 B_{(aq)} \leftrightarrow 6 C_{(aq)}$

$6 C_{(aq)} \rightarrow A_{(aq)} + 2 B_{(aq)}$
Rate = $-\Delta[C]/\Delta t$ (C is disappearing)
AND
 $A_{(aq)} + 2 B_{(aq)} \rightarrow 6 C_{(aq)}$
Rate = $+\Delta[C]/\Delta t$ (C is being created)

Both reactions occur simultaneously with the Rate of destruction = rate of creation

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Equilibrium is Balance

- The forward and reverse reactions are balanced
- The concentrations of all species (reactants and products) become stable
- The equilibrium position is not the same for all reactions

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The Equilibrium Constant

The balance between forward and reverse reaction is summed up in the Equilibrium Constant ("K_{eq}")

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

If K is large, most of the reactants form products.

If K is small, most of the reactants stay as reactants

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Writing Equilibrium Constant Expressions

It is very simple to write the expression for "K_{eq}"

You simply take the concentration (molarity) of all gases, and solutions and raise them to their stoichiometric coefficient

You ignore pure liquids and solids – they don't have a "concentration" and their concentration is considered constant (a unitless 1) and so it doesn't contribute

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$A_{(aq)} + 2 B_{(aq)} \leftrightarrow 6 C_{(aq)}$

$K_{eq} = \frac{[C]^6}{[B]^2[A]}$

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$2 H_2(g) + O_2(g) \leftrightarrow 2 H_2O(g)$

$K_{eq} = \frac{[H_2O]^2}{[H_2]^2[O_2]}$

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$$2 \text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \leftrightarrow 2 \text{H}_2\text{O} (\text{l})$$

$$K_{\text{eq}} = \frac{1}{[\text{H}_2]^2[\text{O}_2]}$$

Notice that the pure liquid is not included - it doesn't really have a concentration and so it becomes '1'

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$$2 \text{Na} (\text{s}) + \text{Cl}_2 (\text{g}) \leftrightarrow 2 \text{NaCl} (\text{s})$$

$$K_{\text{eq}} = \frac{1}{[\text{Cl}_2]}$$

Notice how the 2 solids do not contribute to the equilibrium constant

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$$\text{H}_2\text{S} (\text{g}) + 2 \text{O}_2 (\text{g}) \leftrightarrow \text{H}_2\text{SO}_4 (\text{g})$$

$$K_{\text{eq}} = \frac{[\text{H}_2\text{SO}_4]}{[\text{O}_2]^2[\text{H}_2\text{S}]}$$

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Different kinds of “K”

No matter what the subscript, K IS K IS K IS K IS K IS K! They are all just equilibrium constants and they all get written and used the same way.

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K_c vs K_p

- When a reaction occurs in the gas phase, you can use the **partial pressure** of the gas instead of the concentration.
- To separate the 2 different expressions, they are written differently:
 - K_c = equilibrium constant with concentrations of species
 - K_p = equilibrium constant with partial pressures of the species

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K IS K IS K IS K IS K

The subscripts – and we will see more of them – just tell you what specific type of reaction you are talking about.

Everything else about them is the same.

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Equilibrium Problems

- There are two main types of problems:
 - You know the concentrations and you are calculating the equilibrium constant
 - You know the equilibrium constant and you are calculating the concentrations
- Of course, there are nuances to these problems.

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Calculating the Equilibrium Constant

The only "trick" to calculating an equilibrium constant is that you must know the concentrations AT EQUILIBRIUM.

You cannot simply take any old concentrations at any old time. The reaction must have reached equilibrium

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SORRY!!!!!!

All the flasks on the homework should be 2.0 L.

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A simple example

Hydrogen and oxygen gas will react to form steam (gaseous water). Hydrogen and oxygen gases were mixed in a 2 L flask at 450 C. After equilibrium was established, it was determined that there were 3 moles of water, 1.2 moles of hydrogen and 0.8 moles of oxygen in the flask. What is the equilibrium constant for this reaction at 450 C?

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A simple solution

1st you need a balanced equation:
 $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{g})$

This allows us to immediately write the equilibrium constant expression:
 $K_{eq} = K_c = \frac{[\text{H}_2\text{O}]^2}{[\text{H}_2]^2[\text{O}_2]}$

Since we know the equilibrium concentrations:
hydrogen: 1.2 moles/2 L = 0.6 M
oxygen: 0.8 moles/2 L = 0.4 M
water: 3 moles/2 L = 1.5 M

$$K_c = \frac{(1.5 \text{ M})^2}{(0.6 \text{ M})^2(0.4 \text{ M})}$$

$K_c = 15.63$

Note: You could also calculate Kp by using the ideal gas law - more on that later.

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A Note about Units

Notice that I wrote the equilibrium constant without units even though the concentrations have units:

$$K_c = \frac{(1.5 \text{ M})^2}{(0.6 \text{ M})^2(0.4 \text{ M})}$$

$K_c = 15.63$

Since the units of K will depend on stoichiometry, it is generally considered a UNITLESS QUANTITY!

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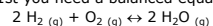
A more complicated problem

Hydrogen and oxygen gas will react to form steam (gaseous water). 4.36 g of hydrogen and 28.6 g of oxygen were mixed in a 2 L flask at 250 C. After equilibrium was established, it was determined that there was 6.6 g of water. What is the equilibrium constant for this reaction at 250 C?

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A series of simple calculations

1st you need a balanced equation:



This allows us to immediately write the equilibrium constant expression:

$$K_{\text{eq}} = K_c = \frac{[\text{H}_2\text{O}]^2}{[\text{H}_2]^2[\text{O}_2]}$$

The question is: what are the equilibrium concentrations of all of the species?

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Determining the concentrations

ICE - ICE - BABY - ICE - ICE

The easiest way to solve this problem is by using an I-C-E chart ("ice chart") where I = initial concentration, C= change in concentration, and E = the equilibrium concentration.

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An ICE Chart

$$2 \text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \leftrightarrow 2 \text{H}_2\text{O} (\text{g})$$

Initial			
Change			
Equilibrium			

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What do you know?

4.36 g hydrogen * 1 mol H₂/2.016 g H₂ = 2.16 mol H₂ (this is the INITIAL amount)

28.6 g oxygen * 1 mol O₂/32.0 g O₂ = 0.894 mol O₂ (this is the INITIAL amount)

6.6 g H₂O * 1 mol H₂O/18.02 g H₂O = 0.366 mol H₂O (this is the EQUILIBRIUM AMOUNT)

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UNITS! UNITS! UNITS!

An ICE chart can use EITHER moles or concentration (molarity), but you must use only one of these in any single ICE chart.

K_c uses molarity, so it is usually easiest to use concentration

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An ICE Chart

$$2 \text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \rightleftharpoons 2 \text{H}_2\text{O} (\text{g})$$

Initial	2.16 mol	0.894 mol	0 mol
Change	?????	?????	???????
Equilibrium	????	?????	0.366 mol

What is the change in quantities?

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The "change" is all about stoichiometry!

$$2 \text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \leftrightarrow 2 \text{H}_2\text{O} (\text{g})$$

Initial	2.16 mol	0.894 mol	0 mol
Change	- 2 x	- x	+ 2 x
Equilibrium	????	?????	0.366 mol

Now it is easy to finish filling in the ICE chart!

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An ICE chart is really just "accounting for moles"

$$2 \text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \leftrightarrow 2 \text{H}_2\text{O} (\text{g})$$

Initial	2.16 mol	0.894 mol	0 mol
Change	- 2 x	- x	+ 2 x
Equilibrium	$2.16 - 2x$	$0.894 - x$	$2x = 0.366$ mol

It is often helpful to use an ICE chart for other types of problems, it is a great way to keep track of what is going on.

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Determining x allows me to fill in the rest of the chart

$$2 \text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \leftrightarrow 2 \text{H}_2\text{O} (\text{g})$$

Initial	2.16 mol	0.894 mol	0 mol
Change	-2x	-x	+2x
Equilibrium	2.16 - 2x	0.894 - x	2x = 0.366 mol

2x = 0.366 mol
x = 0.183 mol

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Determining x allows me to fill in the rest of the chart

$$2 \text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \leftrightarrow 2 \text{H}_2\text{O} (\text{g})$$

Initial	2.16 mol	0.894 mol	0 mol
Change	-2 (0.183 mol)	-0.183 mol	+2 (0.183 mol)
Equilibrium	2.16 - 2 (0.183) = 1.794 mol	0.894 - 0.183 = 0.711 mol	0.366 mol

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Determining x allows me to fill in the rest of the chart

$$2 \text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \leftrightarrow 2 \text{H}_2\text{O} (\text{g})$$

Initial	2.16 mol	0.894 mol	0 mol
Change	+2x	+x	-2x
Equilibrium			0.366 mol

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Now we need to calculate the concentrations and put them into K_c

$[H_2] = 1.794 \text{ mol}/2L = 0.897 \text{ M}$
 $[O_2] = 0.711 \text{ mol}/2L = 0.356 \text{ M}$
 $[H_2O] = 0.366 \text{ mol}/2L = 0.183 \text{ M}$
 $K_{eq} = K_c = \frac{[H_2O]^2}{[H_2]^2[O_2]}$
 $K_{eq} = K_c = \frac{[0.183]^2}{[0.897]^2[0.356]}$
 $K_c = 0.117$

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We could do the same calculation using concentrations directly in the ICE chart

2.16 mol H_2 /2L = 1.08 M (this is the INITIAL concentration)
0.894 mol O_2 /2L = 0.447 M (this is the INITIAL concentration)
0.366 mol H_2O /2L = 0.183 M (this is the EQUILIBRIUM concentration)

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An ICE Chart

$2 H_2 (g) + O_2 (g) \leftrightarrow 2 H_2O (g)$

Initial	1.08 M	0.447 M	0 M
Change	- 2 x	- x	+ 2 x
Equilibrium	1.08 - 2x	0.447 - x	2x = 0.183 M

$2x = 0.183 \text{ M}$
 $x = 0.0915 \text{ M}$

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An ICE Chart

$$2 \text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \leftrightarrow 2 \text{H}_2\text{O} (\text{g})$$

Initial	1.08 M	0.497 M	0 M
Change	- 2 x	-x	+ 2 x
Equilibrium	1.08 - 2(0.0915) 0.897 M	0.447 - 0.0915 0.356 M	0.183 M

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Now we simply put the concentrations into K_c

$$K_{\text{eq}} = K_c = \frac{[\text{H}_2\text{O}]^2}{[\text{H}_2]^2[\text{O}_2]}$$
$$K_{\text{eq}} = K_c = \frac{[0.183]^2}{[0.897]^2[0.356]}$$
$$K_c = 0.117$$

Same answer as before!

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How to use K_c

What if you already know the equilibrium constant? How do you use it?

If you know the equilibrium constant, you can use it to determine the equilibrium concentrations of all of the reacting species!

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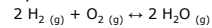
Another Simple Problem

The K_c value for the formation of water from hydrogen and oxygen at 850 C is 4×10^{-6} . If I mix 5.0 grams of hydrogen and 5.0 grams of oxygen in a 3 L flask at 850 C, what is the equilibrium concentration of the water?

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Another simple solution

1st you need a balanced equation:

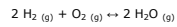


This allows us to immediately write the equilibrium constant expression:

$$K_c = \frac{[\text{H}_2\text{O}]^2}{[\text{H}_2]^2[\text{O}_2]} = 4 \times 10^{-6}$$

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Again, the Power of ICE



Initial			
Change	-2x	-x	+ 2 x
Equilibrium			

The "Change" line is always just stoichiometry

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We already know a couple of things

5.0 g hydrogen * 1 mol H₂/2.016 g H₂
= 2.48 mol H₂/3 L = 0.827 M

5.0 g oxygen * 1 mol O₂/32.0 g O₂
= 0.156 mol O₂/3 L = 0.0533 M

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Again, the Power of ICE

$2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \leftrightarrow 2 \text{H}_2\text{O}(\text{g})$

Initial	0.827 M	0.0533 M	0 M
Change	-2x	-x	+ 2 x
Equilibrium	0.827 - 2 x	0.0533 - x	2 x

Now, we know everything – well, sort of.

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We have all of the equilibrium concentrations in terms of x...

...we can use K_c to solve for x

$$K_c = \frac{[\text{H}_2\text{O}]^2}{[\text{H}_2]^2[\text{O}_2]} = 4 \times 10^{-6}$$
$$0.000004 = \frac{(2x)^2}{(0.827-2x)^2(0.0533-x)}$$

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It looks like a mess...

...and it sort of is (although your calculator can probably solve it)

BUT you can simplify it with a helpful assumption:

ASSUME $x \ll 0.0533$

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If we assume x is small

$K_c = \frac{[H_2O]^2}{[H_2]^2[O_2]} = 4 \times 10^{-6}$

$4 \times 10^{-6} = \frac{(2x)^2}{(0.827-2x)^2(0.0533-x)}$

$.0533-x \approx 0.0533$
 $0.827 - 2x \approx 0.827$

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A very simple problem remains

$4 \times 10^{-6} = \frac{(2x)^2}{(0.827-2x)^2(0.0533-x)}$
 $4 \times 10^{-6} = \frac{(2x)^2}{(0.827)^2(0.0533)}$
 $4 \times 10^{-6} = \frac{4x^2}{0.03645}$
 $1.458 \times 10^{-7} = x^2$
 $3.645 \times 10^{-8} = x^2$
 $X = 1.91 \times 10^{-4}$

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Was the assumption good?

We started by assuming $x \ll 0.0533$

We now "know" that, with this assumption, x is 1.91×10^{-4}

Is $1.91 \times 10^{-4} \ll 0.0533$?

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Critical Judgment

How small is small depends on how accurate an answer you need: If you need 1 sig fig, then any number that is a factor of 10-20 smaller is insignificant. If you need 2 sig figs, then it must be about 100 times smaller. If you need 3 sig figs it must be about 1000 times smaller.

A good general rule for our purposes is that if a number is $< 5\%$ of another number, it is significantly smaller. This maintains 2 sig figs in all the concentrations – usually enough.

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We put x back into the ICE chart

$$2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \leftrightarrow 2 \text{H}_2\text{O}(\text{g})$$

Initial	0.827 M	0.0533 M	0 M
Change	$-2(1.91 \times 10^{-4})$	-1.91×10^{-4}	$+2(1.91 \times 10^{-4})$
Equilibrium	$0.827 - 2(1.91 \times 10^{-4})$ ≈ 0.827	$0.0533 - 1.91 \times 10^{-4}$ ≈ 0.053	3.8×10^{-4}

And we have our answer.
