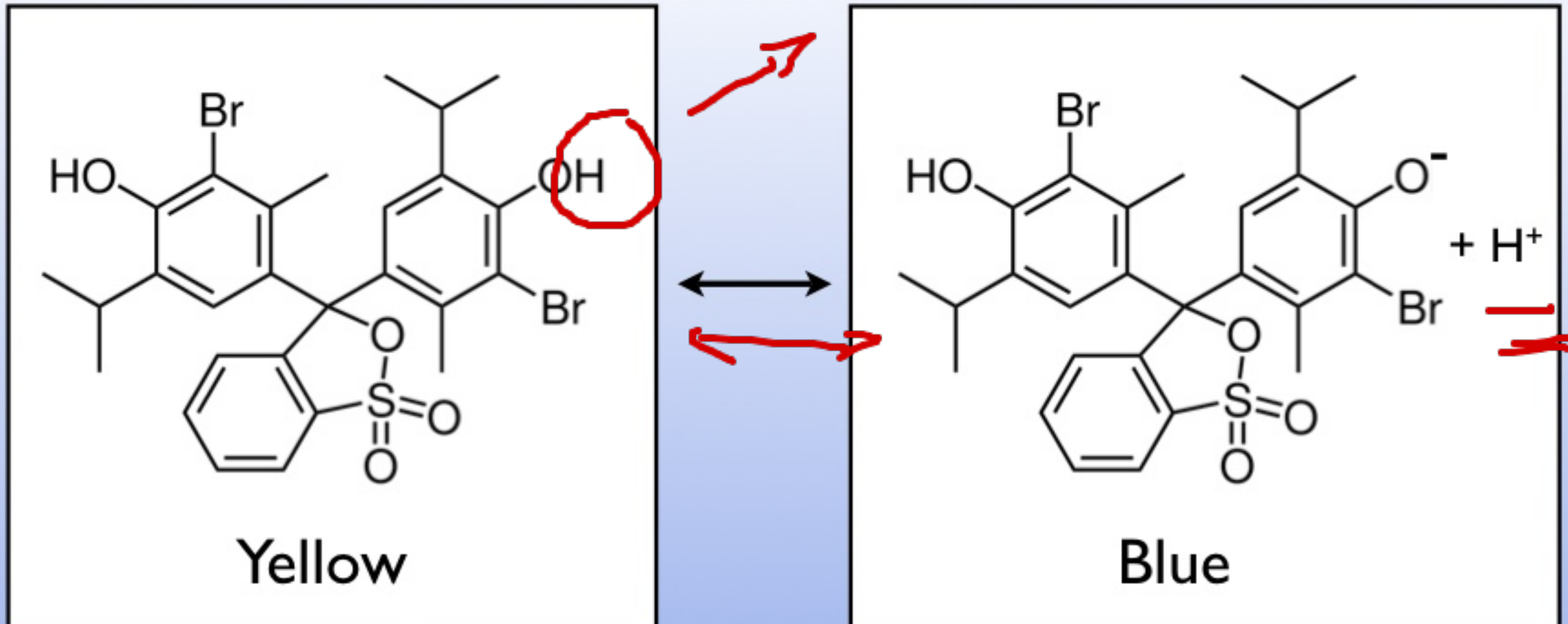
Chemical Equilibria

Why did the color stop changing?

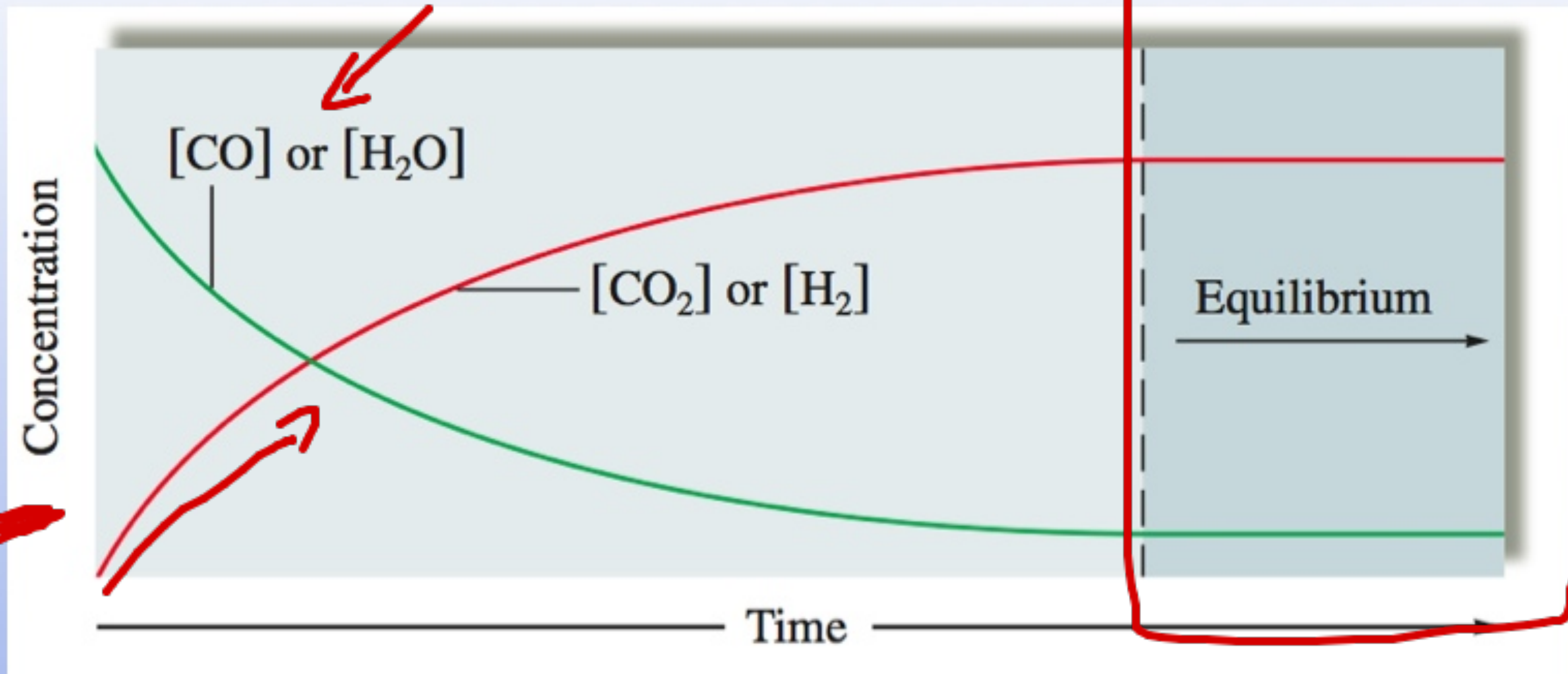
A.	the reactants were all converted to products
B.	the reaction came to equilibrium
C.	the forward and backward reaction rates are the same
D.	B & C
E.	all of the above



ACID BASE

During the reaction the ratio of yellow to blue changes

$H_2O + CO \longrightarrow H_2 + CO_2$



At equilibrium the ratio of the molecules stops changing

it is critical you remember your stiochiometery!

$$3H_2(g) + N_2(g) \longrightarrow 2NH_3(g)$$

Imagine you start out with
10 mole of H_2 and 1 moles of N_2

At equilibrium you find you have 1 mole of NH_3
How many moles of H_2 are there at equilibrium?

A. 5 moles H_2

B. 7 moles H_2

C. 8.5 moles H_2

D. 9.5 moles H_2

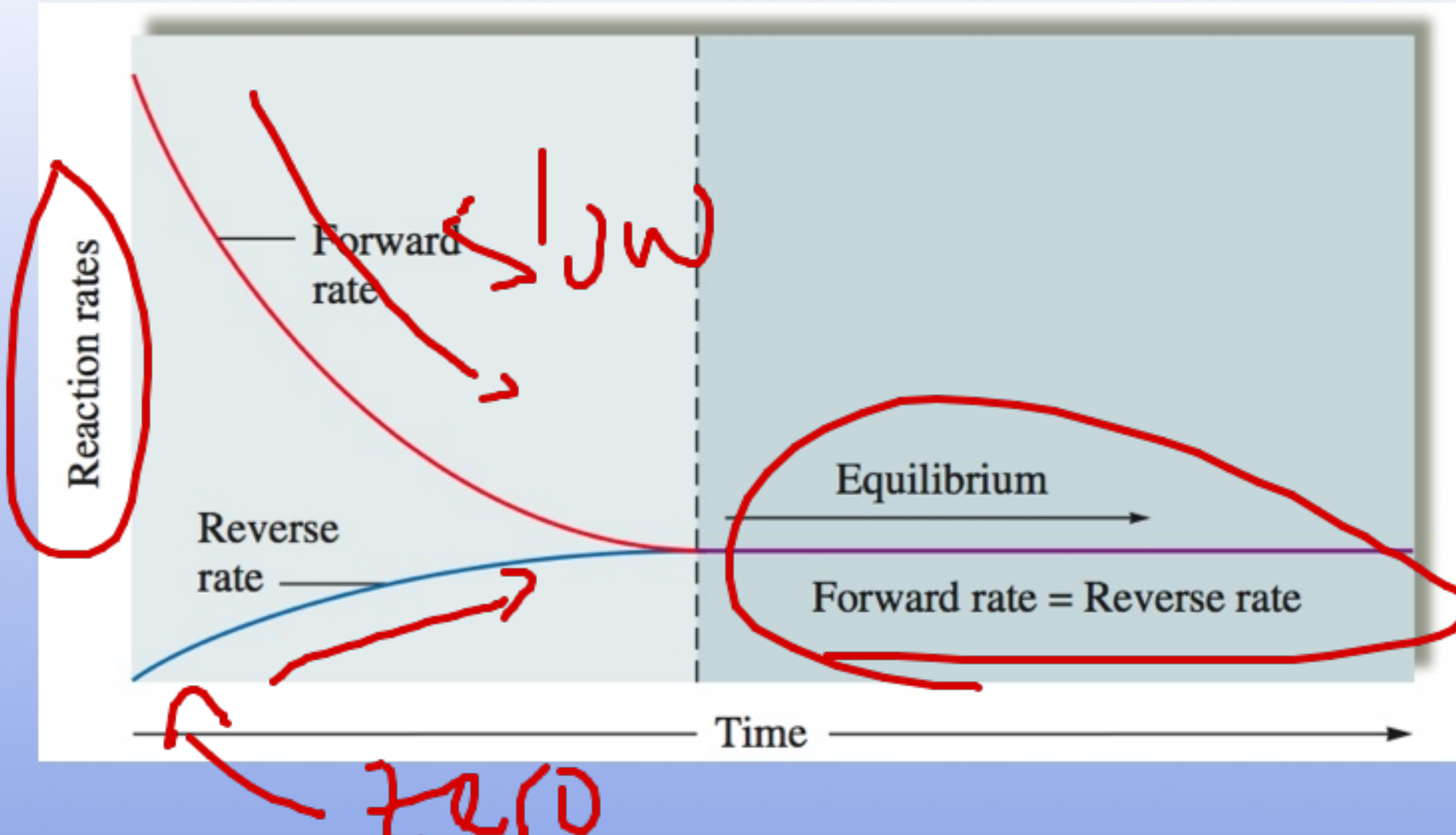
Keeping it straight (R)ICE diagram

$$3H_2(g) + N_2(g) \longrightarrow 2NH_3(g)$$

Compound	Initial	Change	Equilibrium
H_2	10	$-3x$	$10-3x = 8.5$
N_2	1	$-x$	$1-x = .5$
NH_3	0	$+2x$	$+2x = 1$

$x = 0.5$

What is happening? Reaction has not stopped



Equal reaction rates forward and backwards

The key idea

The ratios of the molecules stops changing
We discover the ratio is a constant

We'll give the ratio a name

K

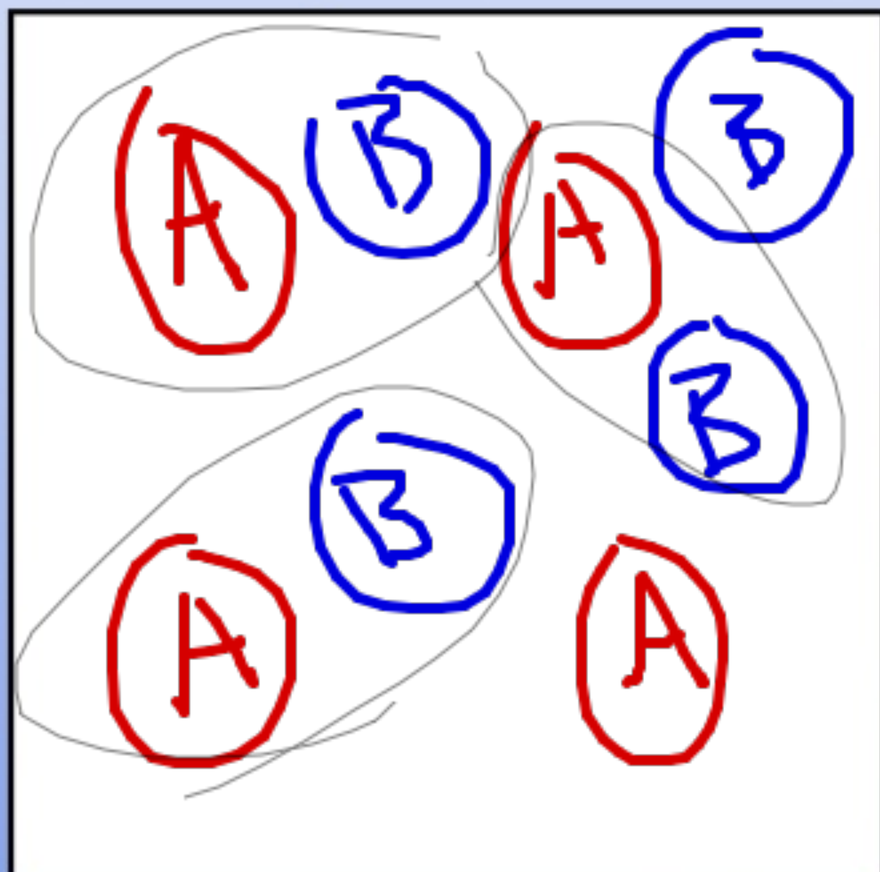
The equilibrium constant
It has to do with equilibrium
It is a constant

changes w/ T

Let's Look an example

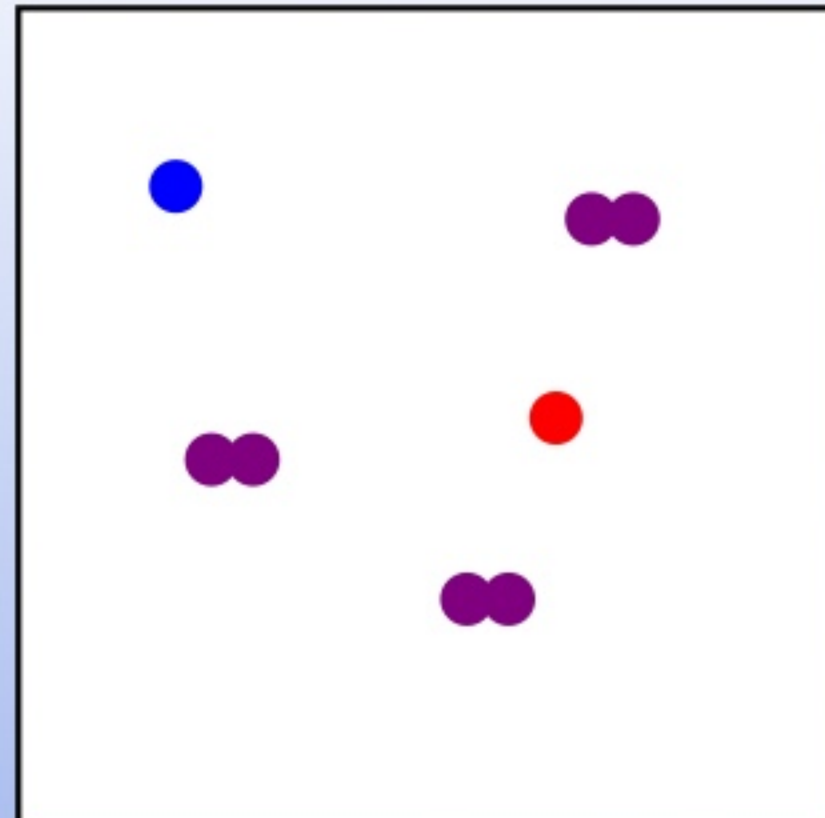
$$A + B \longleftrightarrow AB$$

Initially 4A and 4B



Equilibrium

What is the ratio at equilibrium?

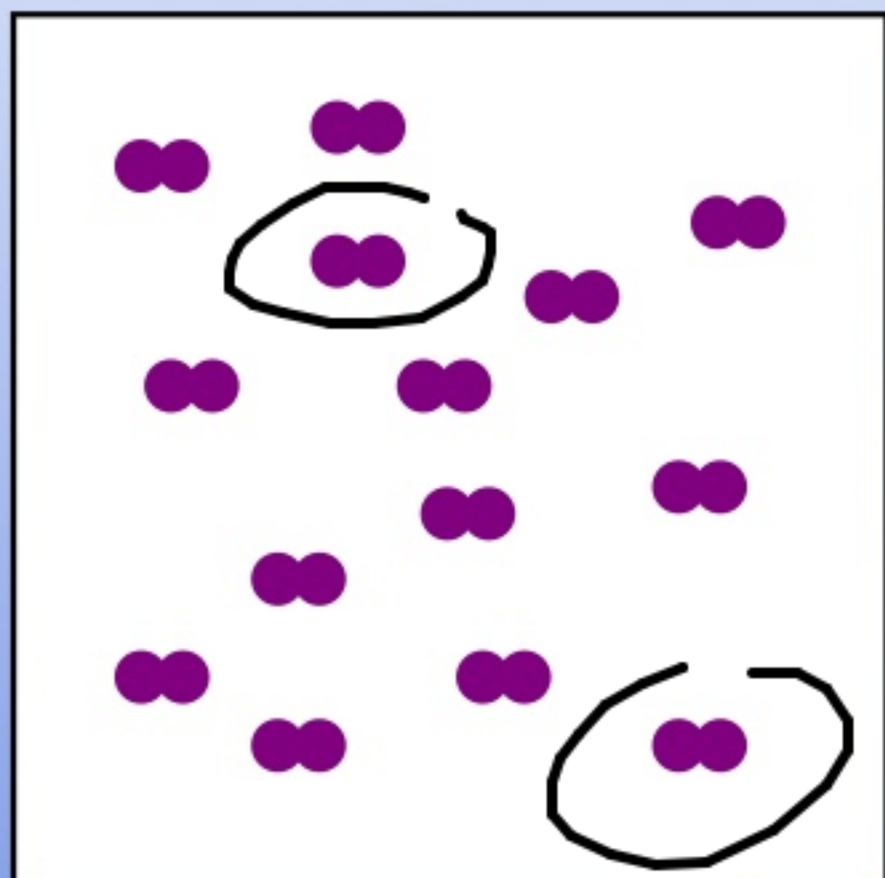


$$K = \frac{\#AB}{\#A \times \#B} = \frac{(3)}{(1)(1)} = 3$$

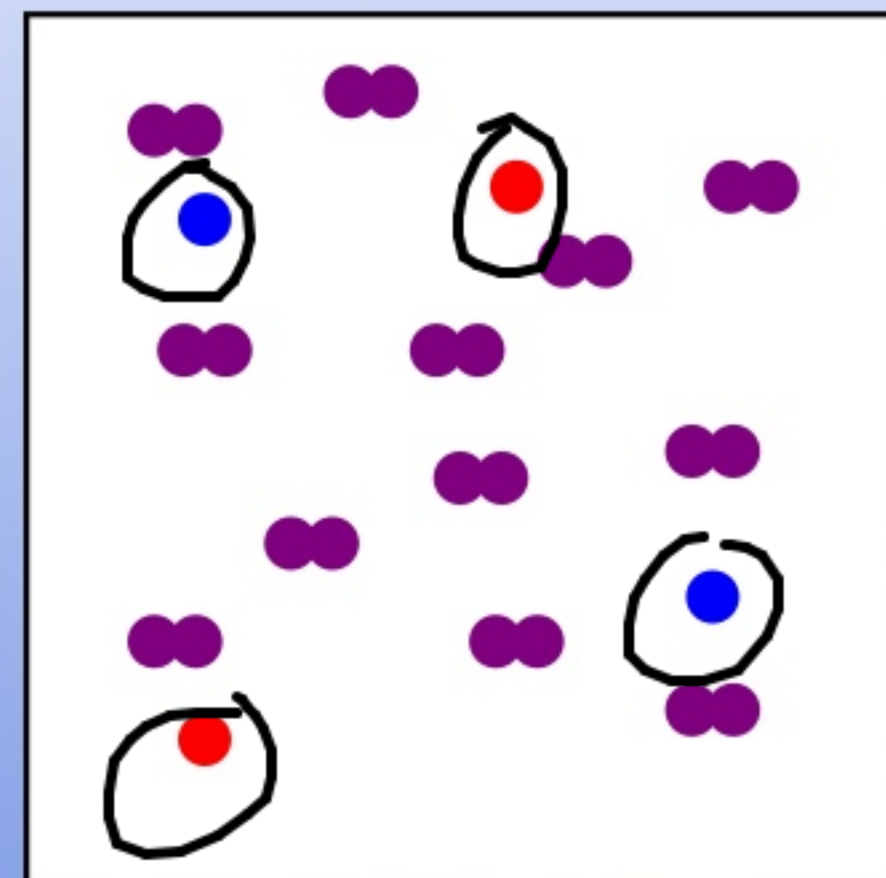
Let's Look an example

$A + B \longleftrightarrow AB$

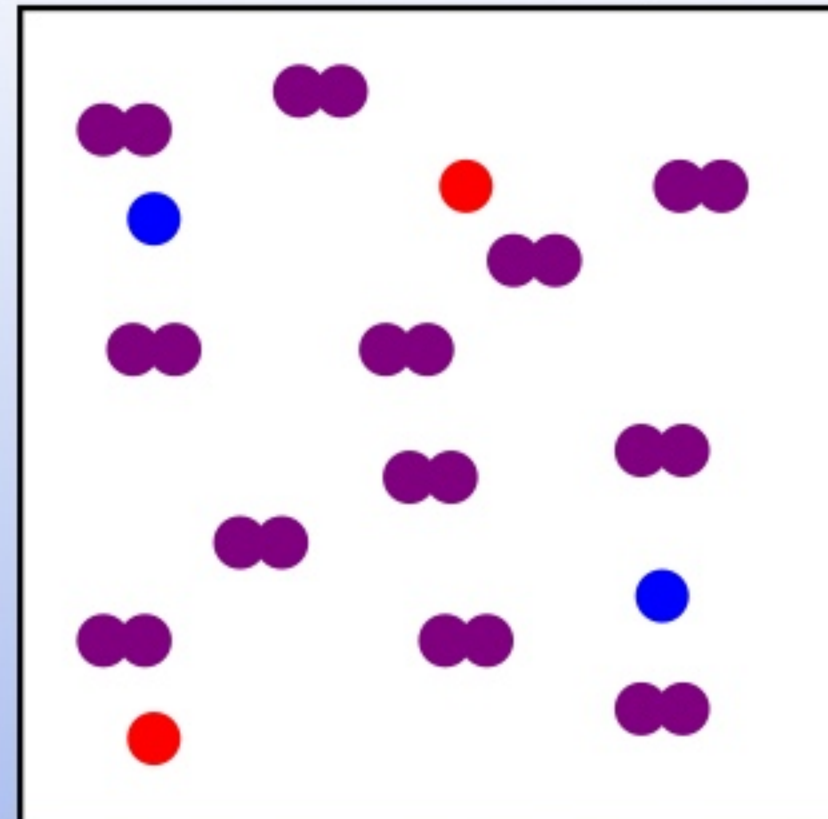
Initially 14AB



Equilibrium



What is the ratio at equilibrium?



$$\frac{\#AB}{\#A \times \#B} = \frac{12}{(2)(2)} = 3$$

How do we write K for a reaction?
First concentrations

$$jA + kB \longleftrightarrow lC + mD$$

$$K = \frac{[C]^l [D]^m}{[A]^j [B]^k}$$

How do we write K for a reaction?

Pressures

$jA + kB \longleftrightarrow lC + mD$

$$K_P = \frac{P_C^{\,l}\, P_D^{\,m}}{P_A^{\,j}\, P_B^{\,k}}$$

Why are there sometime "standard pressures"

$$\left(\frac{P_c}{P^\circ}\right) = P_c$$

K_c UNITS M

if $P^\circ = 1$ atm (bar)

You can only leave it out if the pressure has the same units as the standard pressure

THESE UNITS

What is the expression for the equilibrium constant for this reaction?

$$3H_2(g) + N_2(g) \longleftrightarrow 2NH_3(g)$$

A. $(P_{NH3})/(P_{N2})(P_{H2})$

B. $(P_{N2})(P_{H2})/(P_{NH3})$

C. $(P_{NH3})^2/(P_{N2})(P_{H2})^3$

D. $(P_{N2})3(P_{H2})/2(P_{NH3})$

K depends on $\Delta_r G^\circ$

STANDARD

$$\Delta_R G^\circ = -RT \ln K$$

You need to be able to use a table to find $\Delta_r G^\circ$
from $\Delta_f G^\circ$ or
from $\Delta_f H^\circ$ to find $\Delta_r H^\circ$ and S° to find $\Delta_r S^\circ$

Interpreting K and $\Delta_r G°$

Pure **Products** (in standard state)
are Lower in Free Energy

$\Delta_R G° < 0 \quad K > 1$ Favors P

Pure **Reactants** (in standard state)
Lower in Free Energy

$\Delta_R G° > 0 \quad K < 1$ Favors R

$-T\Delta_R S^\circ$

For a particular reaction

$\Delta_r H^\circ = 10$ kJ mol^{-1} and $\Delta_r S^\circ = 20$ J K^{-1} mol^{-1} $= -8000$

Assuming $\Delta_r H^\circ$ and $\Delta_r S^\circ$ don't change with temperature does this reaction favor the products or the reactants at 400K?

A. Products

B. Reactants

C. There is no way to know without a balance equation

$\Delta_R G^\circ > 0 \quad K < 1$

$\Delta_R G^\circ = \Delta_R H^\circ - T\Delta_R S^\circ = +2000 \text{ J mol}^{-1}$

$10{,}000 \text{ J mol}^{-1} - (400 \text{ K})(20 \text{ J K}^{-1} \text{ mol}^{-1})$

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

What is K for this reaction at 298K

A. extremely small

B. extremely large

C. approximately one

DOESN'T HAPPEN

(-10 kJ mol⁻¹ ~ all Products
) +10 kJ
all R

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

What is K for this reaction at 298K given that $\Delta_r G° = +113.4$ kJ mol^{-1}

$$K = \exp\left[-\Delta_R G^0 / RT\right]$$

J J

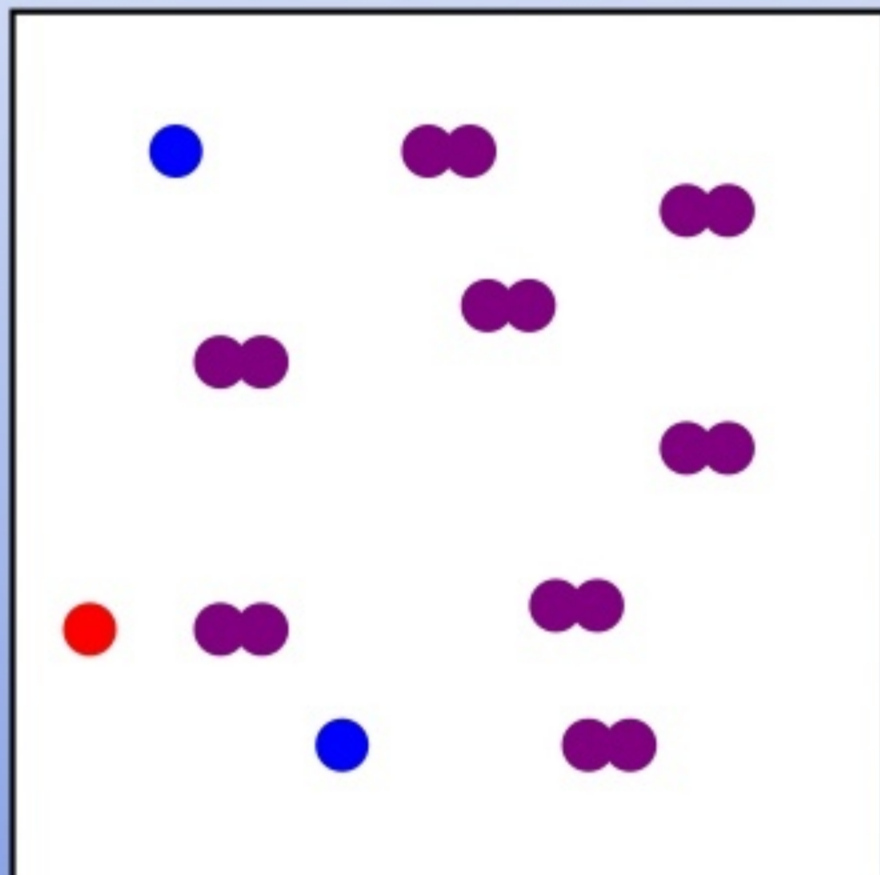
$$= e^{-(113{,}400 \text{ J mol}^{-1} / 8.314 \cdot 298)}$$

$$K = e^{-45}$$

Back to our simple reaction

$$A + B \longleftrightarrow AB$$

Equilibrium?



From before we had
K = 3

Is this system at equilibrium?

$$\text{NOW} = \frac{8}{(1)(2)} = 4$$

$$4 \neq 3$$

This the reaction quotient Q
Q is just like K but
the concentrations or pressures in the expression
are what you have right now

Q > K shift to R

Q < K shift to P

Q = K Equilibrium

K << 1

At 313 K, $\Delta_r G^\circ$ = +41 kJ mol^{-1} for this reaction

$$2H_2S(g) \longleftrightarrow H_2(g) + S_2(g)$$

You find the following partial pressures at 313K
H_2 is 1 atm, S_s is 1 atm, H_2S = 2 atm

How will this reaction proceed?

A. move toward the products

B. move towards the reactants

C. the reaction is at equilibrium

$$Q = \frac{(1)(1)}{(2)^2} = .25$$

K very small K < Q

At 313 K, $\Delta_r G° = +41$ kJ mol^{-1} for this reaction

$$2H_2S(g) \longleftrightarrow H_2(g) + S_2(g)$$

You find the following partial pressures at 313K

H_2 is 1 atm, S_s is 1 atm, H_2S = 2 atm

K = 2.2 x 10^{-3} for this reaction (at some T)

$$2HI(g) \longleftrightarrow H_2(g) + I_2(g)$$

You start with a partial pressure of 1atm of HI
what are the partial pressures at a constant P of 1atm and constant T

	HI	H_2	I_2	Total GAS
I	1	0	0	1
C	$-2x$	$+x$	$+x$	0
E	$1-2x$	$+x$	$+x$	1

$$K = \frac{P_{H_2} P_{I_2}}{P_{HI}^2} = \frac{(x)(x)}{(1-2x)^2}$$

GIVEN

$x = .043$

Do Algebra on your own