Why is there equilibrium?

If the right handside of the reaction is lower in free energy why not all "products"?

If the left handside of the reaction is lower in free enrgy why not all "reactants"?

Entropy of mixing gives the mixture a slightly lower free energy than either extreme

some product + reactants will always be lower in G than all of one or the other

This is only true to compounds that "mix"

gases
and
solutions

As a result
solids and liquids do not appear in the equilibrium expression

For example

$$CaCO_3(s) \longrightarrow CO(s) + CO_2(g)$$

$$K = P_{CO2}$$

No $CaCO_3$ or CaO
they are solids
for equilibrium you must have some solid but the amount doesn't matter

$$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$$

$$K = [H^+][OH^-]$$

(aq) is aqueous
"dissolved in water"

No H_2O its a liquid

The simplest of all equilibria

Solubility

$$MX(s) \longrightarrow M^{+}(aq) + X^{-}(aq)$$

$$K = [M^{+}][X^{-}]$$

special name "solubility product"

$$K_{sp} = [M^{+}][X^{-}]$$

Easy to solve

$$MX_2(s) \longleftrightarrow M^+(aq) \quad + 2X^-(aq)$$

R	M^+	X^-	
I	0	0	none in solution
C	+x	+2x	none in solution
E	+x	+2x	equilibrium is easy

$$K_{sp} = [M^+][X^-]^2 = (x)(2x)^2 = 4x^3$$

For the following reaction $\Delta_R G° = -542\ \mathrm{kJ\ mol^{-1}}$ at 298K
If I start out with a contain that has a pressure of
1 atm of $H_2(g)$ and 1 atm of $F_2(g)$,
at equilibrium what will the partial pressure of HF(g) be?

$$H_2(g) + F_2(g) \longrightarrow 2HF(g)$$

A. approximately 1 atm

B. approximatley 0 atm

C. approximately 2 atm

D. approximately 4 atm

E. there is no way to know

K is really really big
~ "to completion"

For the following reaction $\Delta_R G° = +740$ kJ mol^{-1} at 298K
If I start out with a contain that has a pressure of
1 mole of Fe_2O_3,
at equilibrium how much solid Fe will I have?

$$Fe_2O_3(s) \longrightarrow 2Fe(s) + (3/2)O_2(g)$$

A. approximately 0 moles

B. approximatley 1 moles

C. approximately 2 moles

D. approximately 3/2 moles

E. there is no way to know

K is really really small
~ "no products"

For the following reaction what is the change value for H_2O?

$$2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(g)$$

R	C_2H_6	O_2	CO_2	H_2O
I	1.0	1.4	1.8	0
C	-2x	?	?	?

A. -2x

B. +2x

C. +3x

D. +6x

For the following reaction what is the equlibrium value for CO_2?

$$2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(g)$$

R	C_2H_6	O_2	CO_2	H_2O
I	1.0	1.4	1.8	0
C	-.5	?	?	?
E	?	?	?	?

A. 2.0

B. 1.4

C. 2.8

D. 1.8 + 4x

$$-2x = -0.5$$
$$x=0.25$$
$$1.8+4x = 2.8$$

For this reaction which has a higher entropy?

$$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$$

A. the products

B. the reactants

C. they are the same

For this reaction which has a lower enthalpy?

$$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$$

A. the products

B. the reactants

C. they are the same

For this reaction which has a lower free energy?

$$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$$

A. the products

B. the reactants

C. they are the same

Liquid Water
will spontaneously dissociate to a small extent

$$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$$

$$K = \frac{[H^+][OH^-]}{1}$$

$$K_w = [H^+][OH^-] = 10^{-14}$$

In water what is the concentration of $[H^+]$?

$$H_2O(l) \longleftrightarrow H^+(aq) + OH^-(aq)$$

$$K_w = [H^+][OH^-] = 10^{-14}$$

Pure Water

	H^+	OH^-
I	0	0
C	+x	+x
E	+x	+x

$$K_w = 10^{-14} = [H^+][OH^-] = (x)(x)$$

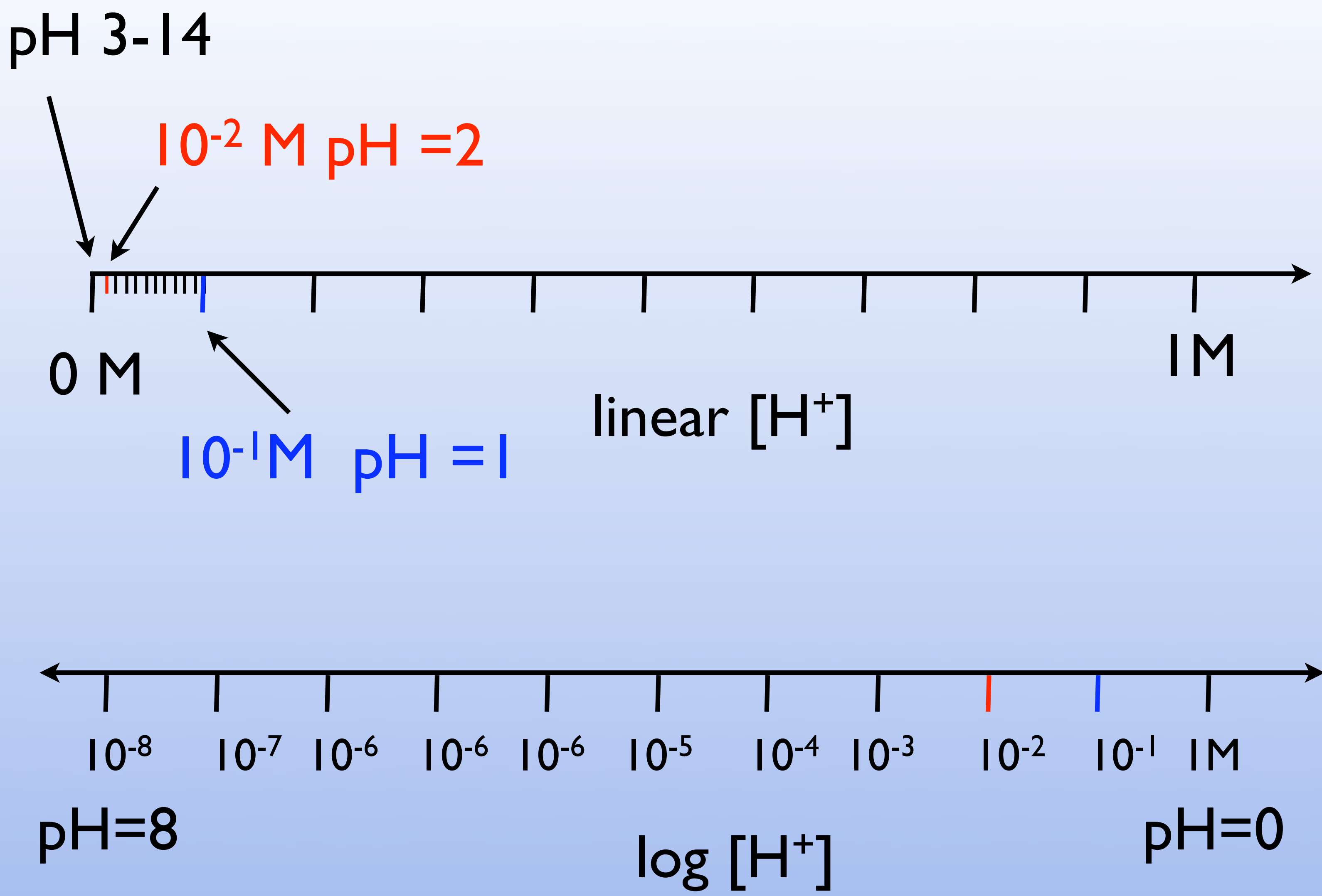
$$x = 10^{-7} \quad [H^+]=[OH^-]=10^{-7}$$

pH

Log scale.
Useful when dealing with very small
or very large number (big ranges of numbers)
every "pH" unit is 10x larger or smaller $[H^+]$

$$pH = -\log[H^+]$$

pH= 13	pH= 7	pH= 2
$[H^+] = 10^{-13}$	$[H^+] = 10^{-7}$	$[H^+] = 10^{-2}$


pH 3-14
10-2 M pH =2
0 M
1M
linear [H+]
10-1M pH =1
10-8
10-7
10-6
10-6
10-6
10-5
10-4
10-3
10-2
10-1
1M
pH=8
log [H+]
pH=0

Acids and Bases

Brønsted-Lowry
Definition

Acid is a proton (H^+) donor

Base is a proton (H^+) acceptor

For example
Hydrochloric Acid (HCl)

pH of pure water at 25°C

$x = 10^{-7}$ $[H^+]=[OH^-]=10^{-7}$

$$pH = -\log[H^+]=-\log(10^{-7}) = 7$$

Neutral	Acidic	Basic
$[H^+]=[OH^-]$	$[H^+]>[OH^-]$	$[H^+]<[OH^-]$
at 25°C	at 25°C	at 25°C
$pH = 7$ $pOH = 7$	$pH < 7$ $pOH > 7$	$pH > 7$ $pOH > 7$

$$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$$

This reaction is endothermic.
Given that information what do you think the pH is for pure water at 60°C?

A. 6.5 ←

B. 7

C. 7.5