Equilibrium Theory

1. When a system is at equilibrium, what can be said about:
   a) the rates of the forward and reverse processes? RATES =
   b) the amount (concentration) of the reactants and products? [ ]'s constant
   c) What's happening at the macroscopic level? NOTHING
   d) What's happening at the microscopic level? CONTINUES

2. Give at least one condition that must be met for a system to remain at equilibrium.
   Closed system, constant temp, reversible.

Equilibrium Theory

Why does the water in a glass disappear over time, while it doesn't change in a sealed container? Describe what is happening at the molecular level.

↳ Able to reach equilibrium in sealed container

↳ RATE EVAP = RATE Condense

Equilibrium Theory

1. For a system to be at equilibrium, what must be true about the:

    a) Rates of the forward and reverse processes? = , constant

    b) Concentrations of the reactants and products? Constant

2. Give 2 **conditions** that must be met in order for a system to be able to reach an equilibrium.

rev. / closed → const tem
→ nothing escapes

Calculating Keq

Calculate the value of K_{eq} for the following reaction if the concentrations at equilibrium were found to be $[N_2O]$=1 mol/L and $[H_2O]$=2 mol/L.

$$NH_4NO_{3(S)} \longleftrightarrow N_2O_{(g)} + 2H_2O_{(g)}$$

$$K_{eq} = \frac{[N_2O][H_2O]^2}{1} = (1)(2)^2 = 4 = K_{eq}$$

3 < 4

4 > 3

Which reaction is favoured at equilibrium? Prod.

Calculating Keq

Given the reaction ~~$A_{(s)}$~~ + **$2B_{(aq)}$ <----> $2C_{(aq)}$**. The equilibrium concentrations were found to be [A]=2.5mol/L, [B]=2.5mol/L and [C]=0.1mol/L. Calculate Keq AND state which side of the reaction is favoured.

$$K_{eq} = \frac{[C]^2}{[B]^2} = \frac{(0.1)^2}{(2.5)^2} = 0.0016$$

Keq < 1 (more products) So prod. favoured

ICE Problems

1. For the reaction **H_2 + I_2 <----> 2HI**. If initially there were 2.4 moles of I_2 and 2.4 moles of H_2 in a 1.0L container, and at equilibrium there was found to be 1.0mol of HI, determine the equilibrium constant for the reaction.

	H_2	+ I_2	⇌ 2HI
I	2.4 mol/L	2.4 mol/L	∅
C	−0.5	−0.5	+1.0
E	1.9	1.9	1.0 mol/L

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]} = \frac{(1.0)^2}{(1.9)(1.9)}$$

$K_{eq} = 0.277$

2. state which side of the reaction is favoured at equilibrium. Explain why.

Reacts. Keq < 1 (more reacts at EQ)

ICE Problems

For the reaction $\mathbf{H_2 + I_2}$ **<----> 2HI** Keq = 46

If initially there were 2.4 moles of I_2 and 2.4 moles of H_2 in a 1.0L container, determine the concentration of HI at equilibrium.

H_2	+ I_2	⇌ 2HI
2.4	2.4	Ø
−x	−x	+2x
2.4−x	2.4−x	2x

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

$$\sqrt{46} = \sqrt{\frac{(2x)^2}{(2.4-x)^2}}$$

$$6.8 = \frac{2x}{(2.4-x)}$$

$$6.8(2.4-x) = 2x$$

$$16.3 - 6.8x = 2x$$

$$16.3 = 8.8x$$

$$x = 1.85 \ \tfrac{mol}{L}$$

$$[HI] = 2x = 3.7 \ \tfrac{mol}{L}$$

ICE Problems

Keq= 78.0 for the reaction $\mathbf{\mathit{A_{(s)}} + 2\mathit{B_{(g)}}}$ **<----> $\mathbf{2\mathit{C_{(g)}}}$**

If 4 moles of A and 6 moles of B are placed in a 2.0L container, determine the concentration of C at equilibrium. **Hint...watch out for the states!!!**

	A	+ 2B	⇌ 2C
I	X	3	Ø
C	X	−x	+x
E	X	3−x	x

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

$$\sqrt{78} = \sqrt{\frac{(x)^2}{(3-x)^2}}$$

$$8.832 = \frac{x}{(3-x)}$$

$$8.832(3-x) = x$$

$$26.496 - 8.832x = x$$

$$26.496 = 9.832x$$

$$x = 2.694$$

$$[C]^2 = 2.7 \ \tfrac{mol}{L}$$

Reaction Quotient

Keq = 1.5 for the reaction **$H_2 + I_2$ <----> 2HI**

A system containing the above reaction was found to have the following concentrations: **$[H_2]$ = 2mol/L** **$[I_2]$ = 1mol/L** **[HI] = 2mol/L**

Is the system at equilibrium? If not, which reaction is favoured?

$$Q = \frac{[HI]^2}{[H_2][I_2]}$$

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

$Q > Keq$

NOT AT EQ. Reverse favoured

Reaction Quotient

Keq = 0.5 for the reaction **$3H_2 + N_2$ <----> $2NH_3$** (assume all gases)

A system containing the above reaction was found to have the following concentrations: **$[H_2]$ = 2mol/L** **$[N_2]$ = 1mol/L** **$[NH_3]$ = 2mol/L**

Is the system at equilibrium? If not, which reaction is favoured?

$$Q = \frac{[NH_3]^2}{[H_2]^3[N_2]} = \frac{(2)^2}{(2)^3(1)} = \frac{4}{(8)(1)} = 0.5$$

@ EQ.

When the system reaches equilibrium, which side will be favoured?

Reacts (Keq < 1)

LeChatelier's Principle

1. Given the reaction **$2H_{2(g)}$ + $O_{2(g)}$ <----> $2H_2O_{(g)}$ + 200 kJ** Determine the shift:

   a) Increase concentration of Hydrogen gas R

   b) Decrease concentration of oxygen gas L

   c) Increase pressure R

   d) Remove energy R

   e) Add a catalyst. No shift

LeChatelier's Principle

Given the reaction

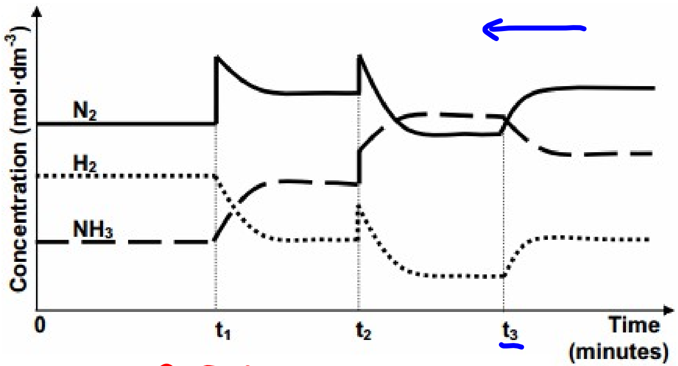
NRG + ***$CH_4(g)$ + $H_2O(g)$ <----> $CO(g)$ + $3H_2(g)$ ΔH = +206 kJ***

State two ways to shift the equilibrium to the "right" and produce more CO.

- Add reactants
- remove products
- add energy
- decrease pressure

LeChatelier's Principle

Use the following graph to describe the possible stresses at t1 and t3 for the following reaction: **$2N_2 + 3H_2 \leftrightarrow 2NH_3 + \text{Heat}$**



[N2] ↑

↑ heat
↓ pressure

LeChatelier's Principle

Use the following graph to describe the possible stresses at each disturbance:

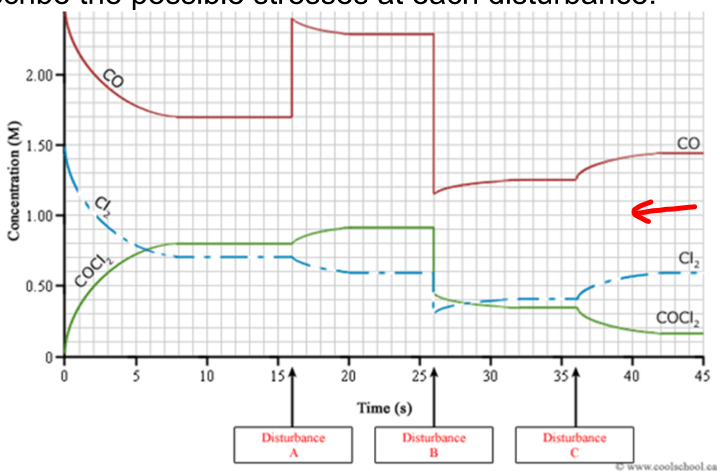
Energy + CO + Cl_2 ⟷ $COCl_2$

A: [CO] ↑

d[Cl2]

B: [COCl2] + [CO] ↓

C: Temp ↓ or press ↓



ICE Problems

Find Keq for the reaction 2A + B <----> 2C if initially 2moles of A and 4 moles of B are placed in a 2.0L container. At Equilibrium 1.0 moles of C are found.

$2A + B \rightleftharpoons 2C$

2A	B	2C
1.0	2.0	ø
-0.5	-0.25	+0.5
0.5	1.75	0.5

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

$= \frac{(0.5)^2}{(0.5)^2(1.75)}$

$K_{eq} = 0.57$

LeChatelier's Principle

Given the following equilibrium system, state the shifts for each of the stresses below

$H_2(g) + Cl_2(g) \leftrightarrow 2HCl(g)$ +heat $\Delta H = - 49.7\ kJ$

a) Increase volume of container NO SHIFT

b) Decrease temperature R

c) Remove HCl R

d) add a catalyst NO SHIFT

Ksp

Calculate the molar solubility of strontium fluoride, SrF_2, if the Ksp of SrF_2 = 2.8 x 10^{-9}

$SrF_{2(s)} \rightleftharpoons Sr^{2+}_{(aq)} + 2F^-_{(aq)}$

	$SrF_{2(s)}$	Sr^{2+}	$2F^-$
	x	∅	∅
	−x	+x	+2x
	∅	x	2x

$K_{sp} = [Sr^{2+}][F^-]^2$

$2.8\times10^{-9} = (x)(2x)^2$

$2.8\times10^{-9} = 4x^3$

$x = 8.9\times10^{-4}\ \frac{mol}{L}$

Now Determine the solubility in grams/250mL

$8.9\times10^{-4}\ \frac{mol}{L} \times \frac{125.6\ g}{1\ mol} = \frac{0.112\ g}{1000\ mL} = \frac{x}{250\ mL}$

$x = 0.028\ g/250mL$

Ksp

Calculate the concentration of Aluminum ions in a solution of Aluminum sulphate, $Al_2(SO_4)_3$, if the Ksp is 8.0 x 10^{-11}

$Al_2(SO_4)_{3(s)} \rightleftharpoons 2Al^{3+}_{(aq)} + 3SO_4^{2-}{}_{(aq)}$

	$Al_2(SO_4)_{3(s)}$	$2Al^{3+}$	$3SO_4^{2-}$
I	x	∅	∅
C	−x	+2x	+3x
E	∅	2x	3x

$K_{sp} = [Al^{3+}]^2[SO_4^{2-}]^3$

$8\times10^{-11} = (2x)^2(3x)^3$

$8\times10^{-11} = (4x^2)(27x^3)$

$8\times10^{-11} = 108x^5$

$\frac{8\times10^{-11}}{108} = x^5$

$7.4\times10^{-13} = x^5$

$x = 0.00375$

$[Al^{3+}] = 2x = 0.0075\ \frac{mol}{L}$

Ksp

$AuCl_3$ has a solubility product of $3.2x10^{-25}$. Calculate the solubility of $AuCl_3$ in a 0.2mol/L $MgCl_2$ solution.

$$AuCl_{3(s)} \rightleftharpoons Au^{3+}_{(aq)} + 3Cl^-_{(aq)}$$

3.2×10^{-25} | $\varnothing$ | $\varnothing$

-3.2×10^{-5} | $+3.2\times10^{-5}$ | $+9.6\times10^{-5}$

$\varnothing$ | 3.2×10^{-5} | 9.6×10^{-5}

$$Ksp = [Au^{3+}][Cl^-]^3$$
$$= (3.2\times10^{-5})(9.6\times10^{-5})^3$$
$$Ksp = 2.95\times10^{-13}$$

Ksp

Calculate the Ksp for bismuth sulfide (Bi_2S_3), which has a solubility of 5.1 x 10^{-13} g/L.

$$Bi_2S_3 \rightleftharpoons 2Bi^{3+} + 3S^{2-}$$

9.91×10^{-16} | $\varnothing$ | $\varnothing$

X | $+1.98\times10^{-15}$ | $+2.97\times10^{-15}$

 | 1.98×10^{-15} | 2.97×10^{-15}

$$5.1\times10^{-13}\ \frac{g}{L} \times \frac{1\ mol}{514.3\ g} = 9.91\times10^{-16}\ \frac{mol}{L}$$

$$Ksp = [Bi^{3+}]^2[S^{2-}]^3$$
$$= (1.98\times10^{-15})^2(2.97\times10^{-15})^3$$
$$Ksp = 1.03\times10^{-73}$$

Ksp

If 2 x 10^{-3} g of an imaginary insoluble salt $Fe(OH)_3$ is all that can be dissolved in 100 ml of water at 25°C, calculate the Ksp.

$$2\times10^{-3}\,g \times \frac{1\,mol}{106.83\,g} = \frac{1.87\times10^{-5}\,mol}{0.1\,L} = 1.87\times10^{-4}\,\frac{mol}{L}$$

$$K_{sp} = (1.87\times10^{-4})(5.6\times10^{-4})^3$$

$$K_{sp} = 3.28\times10^{-14}$$

Common Ion, Ksp

Which will dissolve more Lead (II) Chloride: Pure water or 0.1mol/L $FeCl_3$?

Pure water: no common ion

Find the solubility of lead (II) chloride in a 0.1mol/L solution of $FeCl_3$ if the Ksp of lead (II) chloride is 2×10^{-8}

$$FeCl_3 \rightarrow Fe^{3+} + 3Cl^-$$

0.1 0.1 0.3

$$PbCl_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$$

	Pb^{2+}	Cl^-
	∅	0.3
	+x	+2x
	x	0.3+2x

$$K_{sp} = [Pb^{2+}][Cl^-]^2$$

$$2\times10^{-8} = (x)(0.3+2x)^2 \quad \text{assume}$$

$$2\times10^{-8} = (x)(0.3)^2$$

$$x = 2.2\times10^{-7}\,\frac{mol}{L}$$

Common Ion

Determine the molar solubility of lead (II) iodide (Ksp = 9.8×10^{-9})

a) in pure water

$$PbI_2 \rightleftharpoons Pb^{2+} + 2I^-$$

	Pb^{2+}	$2I^-$
	0	0
	+x	+2x
	x	2x

$K_{sp} = [Pb^{2+}][I^-]^2$

$9.8 \times 10^{-9} = (x)(2x)^2$

$9.8 \times 10^{-9} = 4x^3$

$x = 0.0013 \frac{mol}{L}$

b) in 0.1 mol/L magnesium iodide solution

$$MgI_2 \rightarrow Mg^{2+} + 2I^-$$

0.1 → 0.1, 0.2

$$PbI_2 \rightleftharpoons Pb^{2+} + 2I^-$$

	Pb^{2+}	$2I^-$
	∅	0.2
	+x	+2x
	x	0.2+2x

$K_{sp} = [Pb^{2+}][I^-]^2$

$9.8 \times 10^{-9} = (x)(0.2 + 2x)^2$ assume

$9.8 \times 10^{-9} = (x)(0.2)^2$

$x = 2.45 \times 10^{-7} \frac{mol}{L}$