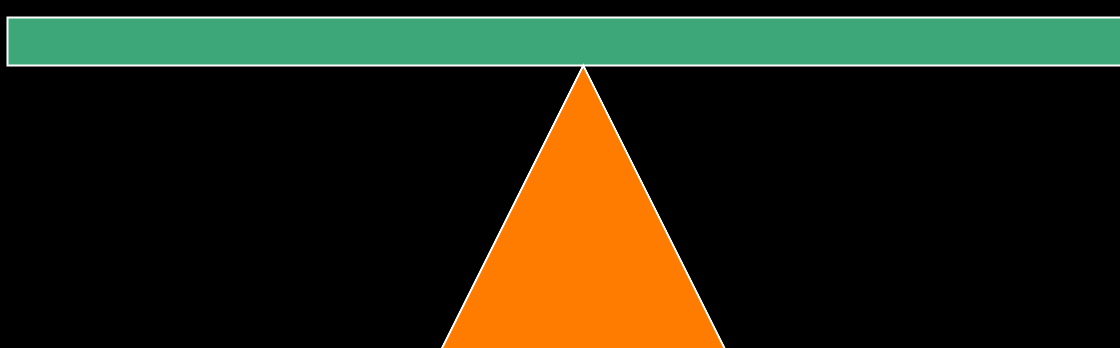
Equilibrium



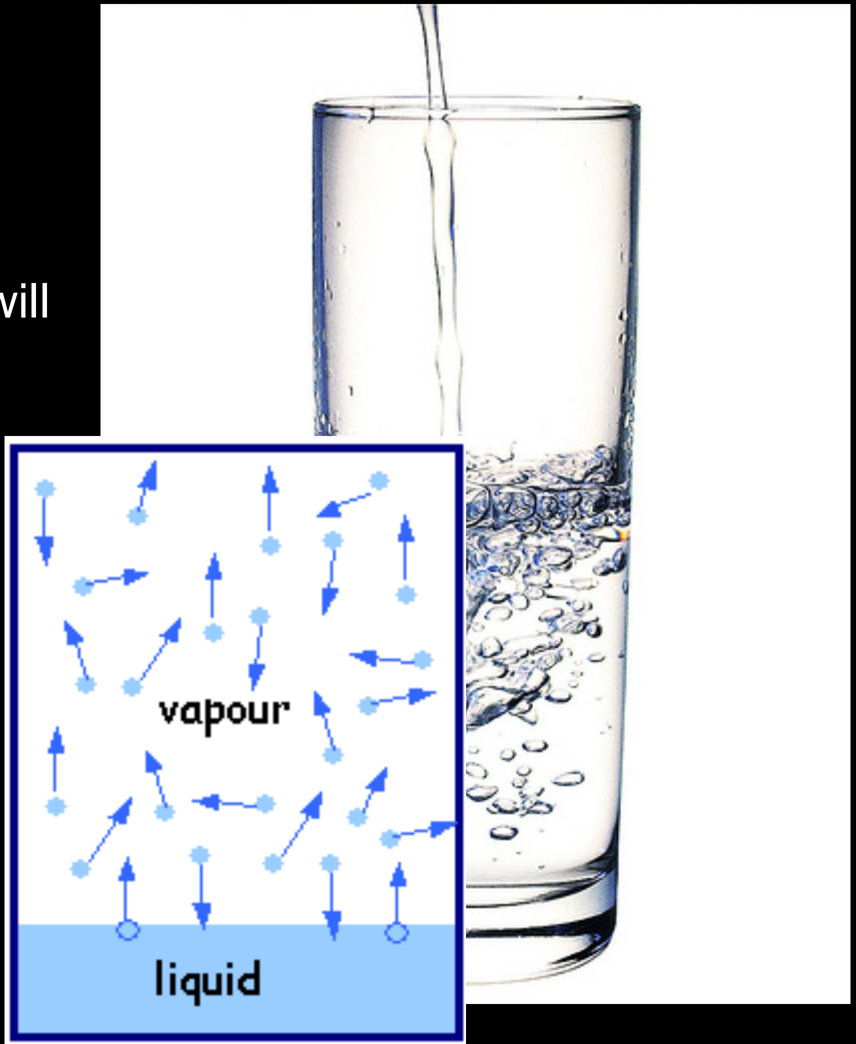
What is equilibrium?

When a glass of water is left out in the open, when will it reach equilibrium?

The liquid will eventually evaporate leaving the glass to dryness.

If the container is sealed, The liquid inside will reach equilibrium with the vapour.
rate of evaporation = rate of condensation

This is **dynamic** because even though the amount of vapour and liquid, for example, do not change there is constant conversion between the two states.



What is equilibrium?

Chemists have a notation to indicate this dynamic equilibrium. They use a double half arrows like this:

or <===> or ↔

Watch the video…or the demo

What is equilibrium?

In Chemical Equilibrium, is where there is no tendency for the quantities of reactants and products to change.

As a result, equilibrium can only be reached in closed systems.

Example

Let's consider the conversion of nitrogen dioxide gas, NO2, into dinitrogen tetraoxide, N2O4 by the reaction:

$$2\ NO_2(g) \rightarrow N_2O_4(g)$$

Nitrogen dioxide is the exhaust of automobiles and is converted to dinitrogen tetraoxide by ultraviolet light form the sun.

The brownish tint of the smog is caused by the presence of the nitrogen dioxide.

Nitrogen dioxide is a brown gas while dinitrogen tetraoxide
is colourless.

Example

Let's consider the conversion of nitrogen dioxide gas, NO2, into dinitrogen tetraoxide, N2O4 by the reaction:

$$2\ NO_2(g) \rightarrow N_2O_4(g)$$

If this reaction took place in a closed container, the brown tint will disappear but not completely.

As the reactants are converted to products, some of the products are converting back to the reactants.

This reaction is reversible

Animation of NO_2 and N_2O_4 molecules Here

Example

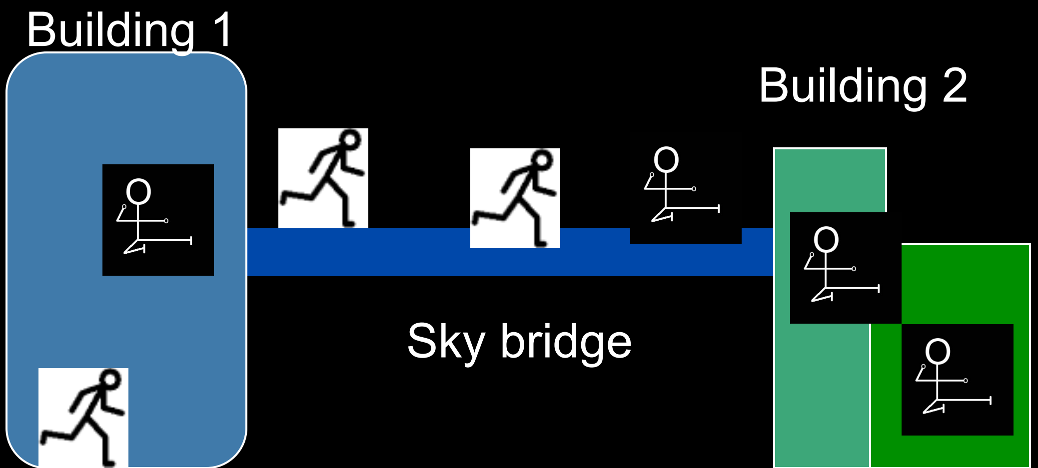
Let's consider the conversion of nitrogen dioxide gas, NO2, into dinitrogen tetraoxide, N2O4 by the reaction:

$$2\ NO_2(g) \rightarrow N_2O_4(g)$$

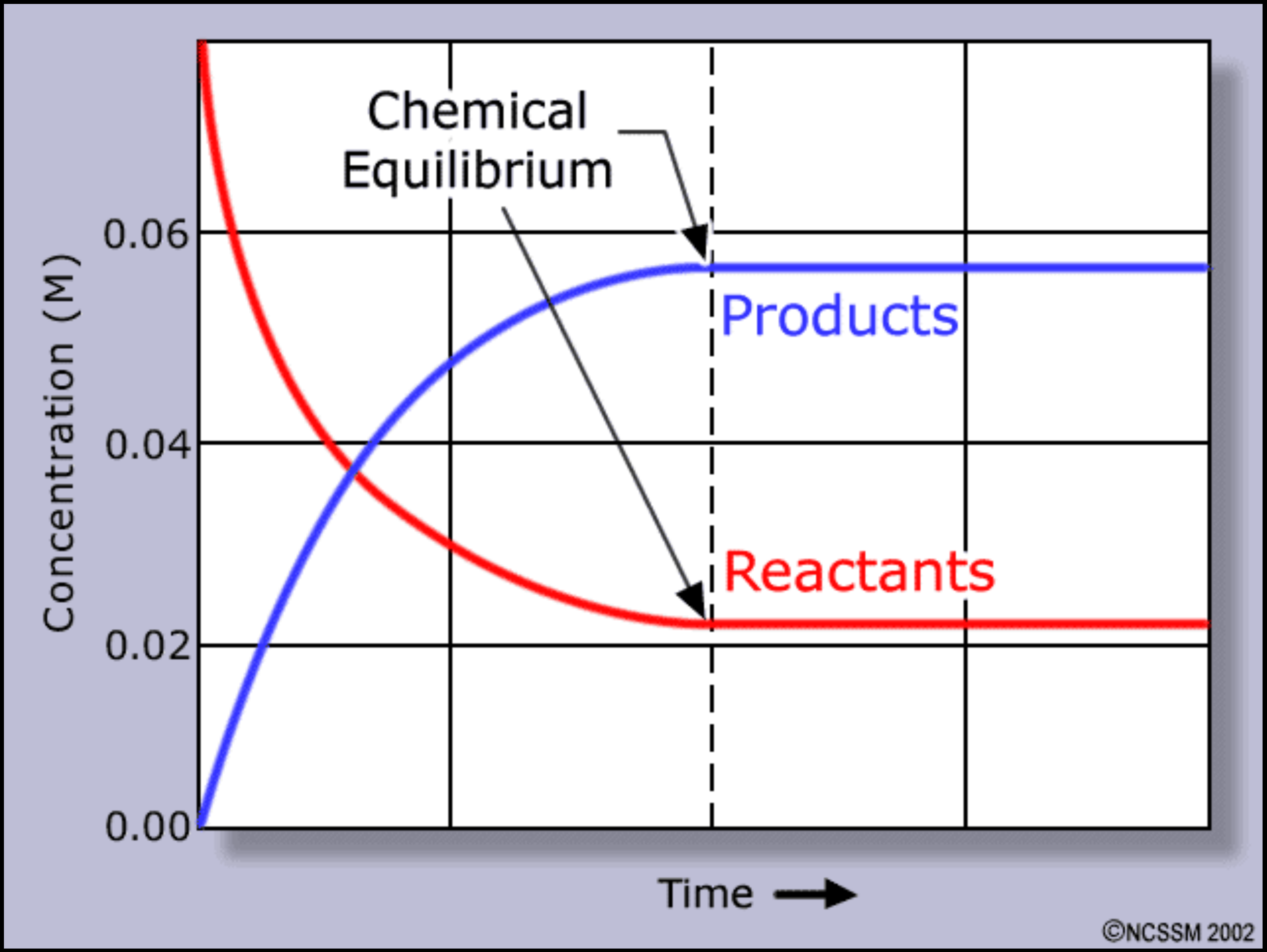
The reactant molecules are the same as the number of the products formed.

At equilibrium, molecules are still in motion.

Analogy:



Equilibrium Graphed



Equilibrium starts when there is a plateau, due to the same concentrations.

Equilibrium Law

Equilibrium law is the ratio of product concentrations to reactants concentrations.
The value from this ratio is the equilibrium constant.

For the reaction $aA + bB \rightleftharpoons cC + dD$

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

$$K_c = \frac{\text{concentration of products}}{\text{concentration of reactants}}$$

Writing Equilibrium Law

Homogeneous equilibria are those in which the reactants and products are all in the same phase, gases (g) or aqueous (aq).

Example 1. Write the equilibrium law for the following equation:

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

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

$$K_c = \frac{\text{concentration of products}}{\text{concentration of reactants}}$$

$$K_c = \frac{[NH_3]^2}{[N_2]\times[H_2]^3} \quad or \quad K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Writing Equilibrium Law

In Heterogeneous equilibria, reactants and products are not in gaseous or aqueous phase. Solids and liquids are omitted

Example 2. Write the equilibrium law for the following equation:

$$C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$$

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

$$K_c = \frac{\text{concentration of products}}{\text{concentration of reactants}}$$

OMIT

$$K_c = \frac{[CO][H_2]}{[H_2O]}$$

Equilibrium Constant

The Kc or equilibrium constant can indicate if the reaction will favor the products or the reactants at equilibrium. (The likelihood if a reaction will occur)

Case 1:
If **K = 1**, [products] = [reactants].
Neither reactants nor products are favoured.

Case 2:
If **K > 1**, the amount of product is greater than reactant.

We say “the products are favoured” or “the equilibrium position lies to the right”.

Case 3:
If **K < 1**, the product concentrations are less than reactant concentrations.
We say “the reactants are favoured” or “the position of equilibrium lies to the left”

Examples

Example 1. For the reaction

N2(g) + 3 H2(g) ←→ 2 NH3(g)

at 225°C, a 2.0 L container holds 0.040 moles of N2, 0.15 moles of H2 and 0.50 moles of NH3. If the system is at equilibrium, calculate K_C.

Step 1. Find concentration

$$\text{Concentration} = \frac{\text{moles}}{\text{Litres}}$$

$$C_{N_2} = \frac{0.040 \text{ moles}}{2.0 \text{ L}} = 0.020 \text{ mol/L}$$

$$C_{H_2} = \frac{0.15 \text{ moles}}{2.0 \text{ L}} = 0.075 \text{ mol/L}$$

$$C_{NH_3} = \frac{0.50 \text{ moles}}{2.0 \text{ L}} = 0.25 \text{ mol/L}$$

Step 2: Get equilibrium law

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Step 3: Calculate

$$K_c = \frac{(0.25)^2}{(0.020)(0.075)^3}$$

$$= \frac{0.0625}{(0.020)(4.219 \times 10^{-4})}$$

$$K_c = 7.4 \times 10^3$$

Examples

Example 2. For the following reaction at 210°C, the KC is 64.0:

$N_2(g) + O_2(g) \leftrightarrow 2\,NO(g)$

The equilibrium concentrations of N2 and O2 are 0.40 mol/L and 0.60 mol/L, respectively. Calculate the equilibrium concentration of NO.

Step 1. Get equilibrium law

$$K_c = \frac{[NO]^2}{[N_2][O_2]}$$

Step 2: Solve for the unknown

$$[NO]^2 = K_c[N_2][O_2]$$

$$[NO] = \sqrt{K_c[N_2][O_2]}$$

Step 3

$$[NO] = \sqrt{(64.0)(0.40)(0.60)} = \sqrt{15.36}$$

$$[NO] = 3.9 \text{ mol/L}$$

I-C-E Tables

I-C-E tables or Initial Concentration, Change in Concentration and Equilibrium Concentration tables.

Purpose: Uses stoichiometry to calculate the initial concentration, the change in concentration or the equilibrium concentrations.

Example: To calculate Equilibrium Constant

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

1.00 moles of hydrogen and 1.00 moles of fluorine are sealed in a 1.00 L flask at 150°C and allowed to react. At equilibrium, 1.32 moles of HF are present. Calculate the equilibrium constant.

I-C-E Tables

The Setup:

Since the flask is 1.0 L, $[H_2] = [F_2] = 1.0$ mol/L and the initial concentration of HF is zero.

We set up the table by first rewriting the equation and ICE down the left side. We also insert the initial concentrations.

	$H_2(g)$	$+ F_2(g)$	$\rightleftharpoons 2HF(g)$	
I	1.00	1.00	0.00	
C				
E				

Find the change in concentration based on stoichiometry

The hydrogen and fluorine will react stoichiometrically to form and equilibrium concentration of 1.32 mol/L HF

According to stoichiometry,

$$1.32 \text{ mol/L HF}\left(\frac{1 \text{ mole } H_2}{2 \text{moles HF}}\right) = 0.66 \text{ mol/L } H_2$$

I-C-E Tables

Calculate the equilibrium concentration

0.66 mol/L of H_2 and F_2 will react. We show this in the change, C, as subtracting 0.66 mol/L from H_2 and F_2 and adding 1.32 mol/L HF.

	$H_2(g)$	$+ F_2(g)$	$\rightleftharpoons 2HF(g)$
I	1.00	1.00	0.00
C	−0.66	−0.66	+1.32
E	0.34	0.34	1.32

Plug and play

This means the equilibrium concentrations of H_2 and F_2 are 0.34 mol/L and HF is 1.32 mol/L

Next, we substitute the equilibrium concentrations into the equilibrium law and solve for K_C:

$$K_c = \frac{[HF]^2}{[H_2][F_2]} = \frac{(1.32)^2}{(0.34)(0.34)} = \frac{1.742}{0.1156}$$

$$K_c = 15.1$$

I-C-E Tables

Example 2: Given equilibrium constant but calculating concentration.

$N_2(g) + O_2(g) \; 2\,NO(g)$

The equilibrium constant is 6.76.

If 6.0 moles of nitrogen and oxygen gases are placed in a 1.0 L container, what are the concentrations of all reactants and products at equilibrium?

Step 1: Form your ICE table

	$N_2(g)$	$+ O_2(g)$	$\rightleftharpoons 2\,NO(g)$
I	6.0	6.0	0.00
C			
E			

I-C-E Tables

Example 2: Given equilibrium constant but calculating concentration.

$N_2(g) + O_2(g)\ \ 2\ NO(g)$

The equilibrium constant is 6.76.
If 6.0 moles of nitrogen and oxygen gases are placed in a 1.0 L container, what are the concentrations of all reactants and products at equilibrium?

Step 2: Now what can we fill in? Algebra! We don't know any exact amounts, but we do have stoichiometry in how they react. In the reaction, for every one H2 or F2, there are 2HF produced, so we use "x" for the moles.

	$N_2(g)$	$+ O_2(g)$	$\rightleftharpoons 2\ NO(g)$
I	6.0	6.0	0.00
C	−x	−x	+2x
E	6.0 − x	6.0 − x	2x

I-C-E Tables

Example 2: Given equilibrium constant but calculating concentration.

$N_2(g) + O_2(g)$ 2 NO(g)

The equilibrium constant is 6.76.

If 6.0 moles of nitrogen and oxygen gases are placed in a 1.0 L container, what are the concentrations of all reactants and products at equilibrium?

Step 3: Write what we know in the equilibrium constant.

$$K_c = \frac{[NO]^2}{[N_2][O_2]}$$

$$6.76 = \frac{(2x)^2}{(6.0 - x)(6.0 - x)}$$

I-C-E Tables

Example 2: Given equilibrium constant but calculating concentration.

$$N_2(g) + O_2(g) \;\; 2\,NO(g)$$

The equilibrium constant is 6.76.
If 6.0 moles of nitrogen and oxygen gases are placed in a 1.0 L container, what are the concentrations of all reactants and products at equilibrium?

Step 3: Solve for x.

$$\sqrt{6.76} = \sqrt{\frac{(2x)^2}{(6.0-x)^2}}$$

$$2.60 = \frac{2x}{6.0-x}$$

$$(6.0-x)(2.60) = \left(\frac{2x}{6.0-x}\right)(6.0-x)$$

$$15.6 - 2.60x = 2x$$

$$15.6 = 2x + 2.60x$$

$$\frac{15.6}{4.6} = \frac{4.6x}{4.6}$$

$$3.4 = x$$

I-C-E Tables

Example 2: Given equilibrium constant but calculating concentration.

$N_2(g) + O_2(g)$ 2 NO(g)

The equilibrium constant is 6.76.
If 6.0 moles of nitrogen and oxygen gases are placed in a 1.0 L container, what are the concentrations of all reactants and products at equilibrium?

Step 4: Plug and play

Since x equals the loss in concentrations of nitrogen and oxygen,

Equilibrium [N2] = [O2] = 6.0 – x = 6.0 – 3.4 = 2.6 mol/L

Equilibrium [NO] = 2x = 2(3.4 mol/L) = 6.8 mol/L

Therefore, at equilibrium nitrogen and oxygen are both 2.6 mol/L and the concentration of NO is 6.8 mol/L.

I-C-E Tables

Step 4: Plug and play

	$N_2(g)$	$+ O_2(g)$	$\rightleftharpoons 2 NO(g)$
I	6.0	6.0	0.00
C	−x	−x	+2x
E	6.0 − x	6.0 − x	2x

Since x equals the loss in concentrations of nitrogen and oxygen,

Equilibrium [N2] = [O2] = 6.0 – x = 6.0 – 3.4 = 2.6 mol/L

Equilibrium [NO] = 2x = 2(3.4 mol/L) = 6.8 mol/L

Therefore, at equilibrium nitrogen and oxygen are both 2.6 mol/L and the concentration of NO is 6.8 mol/L.

1. Write the equilibrium law (mass action expression) for each of the following reactions:
    a. $SO_2(g) + NO_2(g) \rightleftharpoons SO_3(g) + NO(g)$
    b. $2\ C(s) + 3\ H_2(g) \rightleftharpoons C_2H_6(g)$
    c. $3\ O_2(g) \rightleftharpoons 2\ O_3(g)$
    d. $MgCO_3(s) \rightleftharpoons CO_2(g) + 2\ MgO(s)$
    e. $2\ Bi^{3+}(aq) + 3\ H_2S(g) \rightleftharpoons 2\ Bi_2S_3(s) + 6\ H^+(aq)$
    f. $I_2(aq) \rightleftharpoons I_2(s)$
    g. $Cl_2(g) + PCl_3(g) \rightleftharpoons PCl_5(g)$
    h. $I_2(g) + Cl_2(g) \rightleftharpoons 2\ ICl(g)$
    i. $2\ NO_2(g) \rightleftharpoons 2\ NO(g) + O_2(g)$
    j. $2\ SO_2(g) + O_2 \rightleftharpoons 2\ SO_3(g)$
    k. $Cl_2(g) + PCl_3(g) \rightleftharpoons PCl_5(g)$

2. Comment on the favorability of product formation in each of the reactions.
    a. $H_2(g) + F_2(g) \rightleftharpoons 2\ HF(g)$ $K_C = 1.0 \times 10^{13}$
    b. $SO_2(g) + NO_2(g) \rightleftharpoons NO(g) + SO_3(g)$ $K_C = 1.0 \times 10^{2}$
    c. $2\ H_2O(g) \rightleftharpoons 2\ H_2(g) + O_2(g)$ $K_C = 6.0 \times 10^{-28}$

3. Chemists have determined the equilibrium constants for several reactions. In which of these reactions are the products favoured over the reactants?
    a. $K_C = 1.0 \times 10^{2}$
    b. $K_C = 3.5$
    c. $K_C = 0.003$
    d. $K_C = 6.0 \times 10^{-4}$