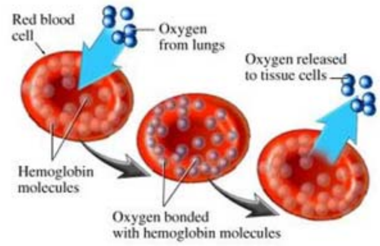
Equilibrium Systems

Section 7.1

Homework
Pg. 427 Practice #1-3
Pg. 428 #1-6a

What would happen if this reaction could only go one way?

$Hb + 4\ O_2 \rightarrow Hb(O_2)_4$



Most chemical reactions can be considered **reversible**.

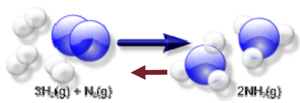
reactants → **products**
products → **reactants**

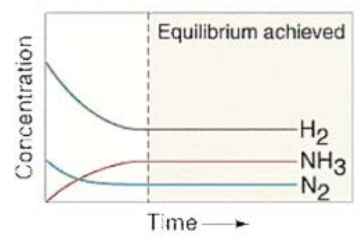
at the same time

reactants $\rightleftharpoons$ **products**

The Haber Process – Synthesis of Ammonia

$N_2 + 3\ H_2 \rightleftarrows 2\ NH_3$





"forward reaction"

reactants $\rightleftarrows$ products

"reverse reaction"

A state called **equilibrium** is achieved when **rate$_{forward}$ = rate$_{reverse}$**

The overall concentrations of all substances remains constant.

The process will APPEAR to have reached a constant state, although the forward and reverse processes are still occurring!

Equilibrium: **Key Terms**

- **Forward reaction**
  - The reaction that proceeds "to the right" – **Reactants → Products**
- **Reverse reaction**
  - The reaction that proceeds "to the left" – **Reactants ← Products**
- **Dynamic equilibrium** – Term that illustrates that molecules are still reacting (being "dynamic"), when a system is at equilibrium.
- **Equilibrium position** – The relative concentration of reactants and products in a system at dynamic equilibrium.

Equilibrium position can be described qualitatively and quantitatively.

$$H_2 + I_2 \overset{78\%}{\rightleftarrows} 2\,HI$$

Qualitative: "The equilibrium lies to the RIGHT."

- This means the FORWARD reaction is favoured.
- You can also say that the PRODUCTS are favoured.

Quantitative: "The percent reaction is 78%".

- This means that at equilibrium, 78% of the possible product is present.

For a closed system, the **same equilibrium concentrations** will be reached, whether equilibrium is approached in the forward or the reverse direction.

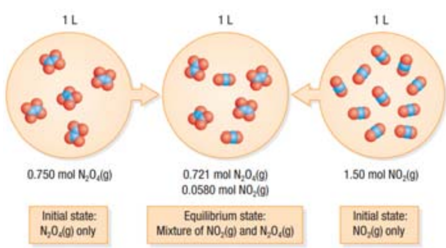
$$N_2O_4 \rightleftharpoons 2\,NO_2$$

Table 1 Changes in Concentrations of $NO_2(g)$ and $N_2O_4(g)$ by the Forward or Reverse Reactions

	Initial concentrations (mol/L)		Final concentrations (mol/L)	
	$N_2O_4(g)$	$NO_2(g)$	$N_2O_4(g)$	$NO_2(g)$
Experiment 1	0.750	0	0.721	0.0580
Experiment 2	0	1.50	0.721	0.0580

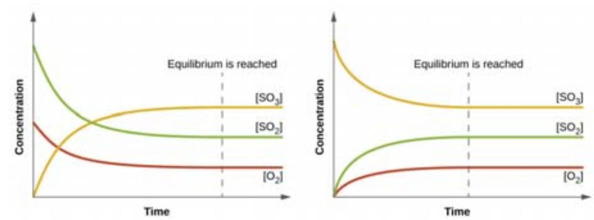
It doesn't matter if the system starts with ONLY REACTANT or ONLY PRODUCT, or even a COMBINATION…

The same concentrations are reached at equilibrium!



Different reaction; same principle, represented graphically

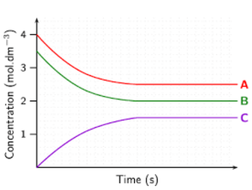
$$2\,SO_2 + O_2 \rightleftharpoons 2\,SO_3$$



Only SO_2 and O_2 initially: Reaction moves to the right

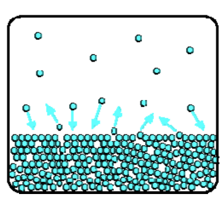
Only SO_3 initially: Reaction moves to the left

Types of Equilibrium



Chemical: Equilibrium between forward and reverse reactions

Solubility: Equilibrium between dissolving & crystallizing



Phase: Equilibrium between changes of state

Learning Checkpoint

- Most chemical and many physical processes are reversible.
- Equilibrium describes a condition in which the rate of the forward reaction equals the rate of the reverse reaction.
  - On a macroscopic level, it appears as though the system is unchanging, since the concentrations will remain constant.

ICE tables are used to analyze equilibrium systems.

ICE stands for…

Initial concentration
Change in concentration
Equilibrium concentration

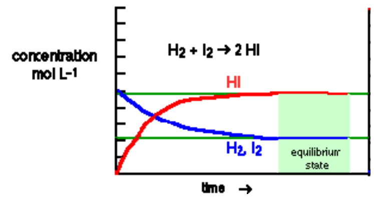
<u>Sample ICE table</u>

$$H_2 + I_2 \rightleftharpoons 2\,HI$$

	H_2	I_2	$\rightleftharpoons$ 2 HI
Initial concentration (mol/L)	0.100	0.100	0.000
Change in concentration (mol/L)	-0.080	-0.080	+ 0.160
Equilibrium concentration (mol/L)	0.020	0.020	0.160

$\Delta[HI]$
$= 2\,\Delta[H_2]$
$= 2\,\Delta[I_2]$

why??



<u>Sample ICE table 2:</u> Calculating unknown equilibrium concentrations

	H_2	I_2	$\rightleftharpoons$ 2 HI
Initial concentration (mol/L)	0.600	0.600	0.000
Change in concentration (mol/L)	- x	- x	+ 2x
Equilibrium concentration (mol/L)	0.600 - x	0.600 - x	0.960

<u>Strategy</u>:
1. Let "x" = the change in concentration per 1 mole of your substances.
2. Use the balanced equation to write expressions for the change in concentration of each substance.
3. Solve for x.
4. Find unknown concentrations.

Example 1.

Calculating Concentrations at Equilibrium

If the reaction begins with 1.00 mol/L concentrations of H_2 and F_2, and no HF, use an ICE table to **calculate the concentrations of H_2 and HF at equilibrium.** The equilibrium concentration of F_2 is measured to be 0.24 mol/L.

	H_2	F_2	$\rightleftharpoons$ 2 HF
Initial concentration (mol/L)			
Change in concentration (mol/L)			
Equilibrium concentration (mol/L)			

Example 2.

4.0 mol of NH_3 is introduced into a 2.0 L container, and heated. After a while, the concentration of NH_3 is found to be constant, at 1.0 mol/L. Determine the equilibrium concentrations of the other two entities.

	2 NH_3 $\rightleftharpoons$	N_2	3 H_2
Initial concentration (mol/L)			
Change in concentration (mol/L)			
Equilibrium concentration (mol/L)			

Summary

- Dynamic equilibrium refers to a state in which the rate of a forward reaction equals the rate of its reverse reaction.
  - chemical reaction; change of state; dissolving
- Equilibrium systems can be analyzed mathematically using ICE tables.
- Changes in concentration reflect stoichiometric relationships.
  - If A → 2 B, then $\Delta[A]$ = x and $\Delta[B]$ = 2x