Equilibrium

Rates of reaction

- Different reactions happen at different speeds.
- There are ways to speed up or slow down a reaction.
- Changing the temperature
  - Warmer reactions tend to go faster
- Changing the amount/concentration of reactants
  - More reactant speeds up the reaction
- Changing the surface area of the reactants
  - Powders react more quickly than "chunks"
- Amount of rate change depends on the reaction.

Collision Theory

- This theory explains how reactions happen at a molecular level.
- In essence, reactant atoms/molecules collide with one another with enough force to break their bonds.
- The pieces from this then reform into the product atom/molecules.

Explaining Why…

- Temperature relates to the rate of motion of the particles.
- Faster moving particles should collide more and more forcefully speeding up the reaction.
- Higher concentration increases the chance of a collision
- More surface area gives more places for collisions to take place
- Which phase is the most reactive?

Relating to phase

- The most reactive phase should be liquid.
- Solids have particles that can only vibrate in place. Collisions will be rare. These should be the least reactive.
- Gases are spaced very far apart, which will decrease the rate of collisions.

Catalysts and Inhibitors

- Catalyst- a substance that increases the rate of a reaction without changing the products of the reaction.
- Catalytic converter speeds up the reaction of emissions of a car to less dangerous products
- Inhibitor- something that slows or stops a reaction
- -food preservatives

Catalyst example

- $2\ O_3 \rightarrow 3\ O_2$
- Ozone will decompose into elemental oxygen, however this process is very slow.
- Chlorine acts as a catalyst as shown
- $O_3 + Cl \rightarrow ClO + O_2$
- $ClO + O_3 \rightarrow Cl + 2\ O_2$
- ClO is an intermediate, something formed in the middle of the reaction that is later consumed.
- Chlorine is a catalyst because it is a reactant in the first step, but a product in the last step. So it isn't used up during the reaction.

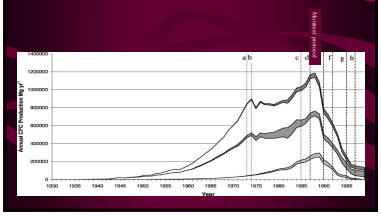
Ozone layer

- This catalyzed reaction was the concern with the ozone layer.
- Ozone in the stratosphere (layer of the atmosphere above us) absorbs UV radiation from the sun.
- Cl atoms (produced from CFC's) was catalyzing the destruction of the ozone.
- CFC's are chlorofluorocarbons. They had a variety of uses included in aerosol cans and refrigerants.

Montreal Protocol 1987

- Holes were forming in the ozone layer that could have been potentially deadly to the planet
- This is a great example of science and political groups working together to fix a problem.
- The Montreal Protocol was signed in 1987, this is a global agreement to phase out the use of all ozone damaging chemicals.
- Many consider it to be the most effective international environmental treaty ever.

Global Production of CFC's



It is not fixed yet, but getting better

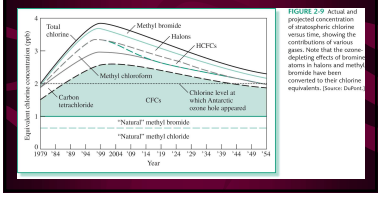


FIGURE 2-9 Actual and projected concentration of stratospheric chlorine versus time, showing the contributions of various gases. Note that the ozone-depleting effects of bromine atoms in halons and methyl bromide have been converted to their chlorine equivalents. [Source: DuPont.]

How a catalyst works

- Our method of representing a reaction skips all of the intermediate steps.
- The catalyst is reacts in these intermediate steps making certain reactions occur with less energy.
- Prior to getting to product the catalyst comes out of the compound.
- Inhibitors are the same but the increase the energy required for certain reactions

Forwards and backwards

- Most reactions can go forwards or backwards
- Neutralization equation
- $H_3O^+ + OH^- \rightarrow 2\ H_2O$
- Self ionization of water
- $2\ H_2O \rightarrow H_3O^+ + OH^-$

Equilibrium

- In water, both of those reactions are occurring simultaneously.
- Equilibrium is when the forward and backward reactions are occurring at the same rate.
- This will cause a stable amount of product and reactant to be present. No net change is occurring when it is at equilibrium. (dynamic equilibrium)
- The amount of product and reactant do NOT have to be equal!

Representing equilibrium

- It is normally represented with a double arrow
- $2\ H_2O \rightleftharpoons H_3O^+ + OH^-$
- This reaction comes to equilibrium when $[H_3O^+] = 1 \times 10^{-7}$ M and $[OH^-] = 1 \times 10^{-7}$ M (assuming the solution is neutral)

Le Châtelier's Principle

- ~whenever stress is applied to a system at equilibrium, a new equilibrium will be obtained to relieve this stress.
- stress is a change in temperature, pressure, or concentration of some component.
- This will change the rate of reaction of either the forward or backward reaction
- So you will see an increase in the concentration of the substances on one side of the equation, and a decrease on the other.
- This will "shift" the equation to the right or left.

Examples

- Endothermic reactions absorb heat, i.e. they need heat to react.
- If the solution is heated prior to the reaction (stress)…
- It will react more quickly
- So the equation will be forced to the right (product side)
- If the reaction is cooled, it will be forced to the left (reactant side)