Chem12 The Reaction Rate-20

Some reactions proceed slowly. The rusting of iron is an example. Other reactions such as the explosion of gunpowder proceed very rapidly. Chemists may want to slow down or speed up a reaction. It is important to be able to measure the rate a reaction proceeds.

e.g. $3Fe(s) + 2O_2(g) \rightarrow Fe_3O_4(s)$

If the above slow reaction (rusting) is observed (in a closed container) the rate may be measured in several ways. The mass of Fe will decrease. The mass and pressure of O_2 will decrease. The mass of Fe_3O_4 will increase.

The **reaction rate** is : $\frac{\text{The rate of change of mass (or moles)}}{\text{Time}}$

For the equation : $A + B \rightarrow C + D$ the rate $= k [A]^x[B]^y$

The rate depends on the concentrations of the reactants. k is a constant. The exponents x and y are determined experimentally. The quantity (x + y) is called the order of the reaction.

Factors affecting the reaction rate

Reactions result from the collisions between reactant molecules. This is the basic premise of **collision theory** but not all collisions lead to reactions). There are two other conditions necessary for a reaction to occur.

a) The reactant molecules must have enough kinetic energy. (greater than the energy of activation).
b) The molecules must collide with the correct collision geometry. (orientation)

In general, the reaction rate is proportional to the frequency of collisions of the reactant molecules.

There are several factors which determine the speed of a reaction :

- Nature of the reactants

  -Ionic substances react very rapidly
  -simpler reactions proceed faster than complex ones
  -gas phase reactions occur more rapidly than liquid or solid phase reactions.

- Homogeneous vs. Heterogeneous reactions

  If the reactant molecules are in the same phase the reaction is faster.

- Concentration/temperature/surface area/catalysts

  Increasing concentration or temperature or the surface area of the reactants increases the number of collisions and the rate. Adding a **<u>catalyst</u>** increases the reaction rate by lowering the activation energy. The catalyst is not used up in the reaction.

<u>Important note</u> : Breaking a bond requires the input of energy (PE increases). Forming a bond releases energy (PE decreases).

Exercise 1: Give 2 ways to measure the rate of the reaction:

$2Ag^{+}(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq)$

(note: $Ag^{+}(aq)$ is colorless, $Cu^{2+}(aq)$ is blue)

Exercise 2) Rank the order of the following reactions from fastest to slowest.

a) $H_2(g) + F_2(g) \rightarrow 2HF(g)$
b) $Ag^{+1}(aq) + Cl^{-1}(aq) \rightarrow AgCl(s)$
c) $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

Exercise 3) Rank the order of the following reactions from fastest to slowest.

a) $Fe(s) + I_2(s) \rightarrow FeI_2(s)$
b) $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$
c) $Ba^{2+}(aq) + SO_4^{-2}(aq) \rightarrow BaSO_4(s)$
d) $H_2(g) + Cl(g) \rightarrow 2HCl(g)$

Exercise: 4) Give 4 ways to increase the reaction rate of:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Exercise 5) Explain: a) Gasoline vapor burns faster than liquid gas.
b) Paper burns faster than wood. (masses being equal)

Exercise 6) The following reaction takes place in a closed container:

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

In terms of collision theory explain what happens to the reaction rate when :

a) T is decreased

b) O_2 concentration is increased

c) a catalyst is added

d) S is ground to a powder

Exercise 7) Explain how a catalyst works.

Answers :1) measure the color change (use a spectrophotometer), find the rate at which the mass of Cu(s) decreases. 2) b, a, c, 3) c, d, b, a, 4) powder $CaCO_3(s)$, add a catalyst, increase temperature, add more HCl, 5)a) gas phase reactions occur at a faster rate, b) surface area is greater, 6)a) reaction slows because of less collisions as particles are moving more slowly; b), d), reactions speed up because collisions between reactant molecules increase; c) reaction rate increases as the activation energy is decreased, 7) A catalyst reduces the activation energy so that more collisions are "successful" and the reaction rate increases.