If a man neglects education, he walks lame to the end of his life. --Plato

Factors Affecting the Rates of Reactions

To begin to understand how chemical reactions actually occur on the molecular level it helps to have some kind of model for what the invisible molecules are doing. Such a model should account for the changes in rate that we observe in every day life as well as in some of the more obscure reactions that we sometimes see in the lab. For example, most people recognize that food is kept in a refrigerator to help retard spoiling. And while it is true that the low temperatures do slow down the action of bacteria, there are also significant purely chemical processes which contribute to the spoiling of food. And these are slowed at low temperature (not to mention the chemical life processes of the bacteria!).

Similarly, anyone who has done photo processing knows that the temperature of the developer and other chemical solutions influences the time for each part of the process. It seems apparent that temperature is an important factor in determining the rate of chemical processes. Why?

The particle model for matter and elementary kinetic theory should lead us to suspect that at higher temperatures the molecules in a system are moving more rapidly. The opposite is true at lower temperatures. How is *motion* involved in reactions? The simplest guess is that molecules that will interact (even sometimes exchange parts) must come in contact with one another. This is the basic premise of **Collision Theory**.

Collision Theory is a very mathematical description of what molecules are doing during a reaction. We need not go into the mathematics, but the main concepts are pretty basic:

1. collisions must occur if a reaction is to take place
2. not all collisions result in reactions

In addition to the effects that temperature has on both of these, there are other factors which contribute to collisions and their effectiveness. Seeing what some of them are and how they affect reaction rates is what this experiment is all about.

Adapted from: A Demonstration Illustrating the Factors Determining Rates of Chemical Reactions, L.H. Holmes, Jr., *J. Chem. Educ.*, 1991, vol. 68, No. 6, p. 501

Preparing to experiment

You will be provided with the following materials:

1. acidified 0.2 M $Na_2C_2O_4$ (sodium oxalate) [use about 10 drops]
2. acidified 0.1 M $FeSO_4$ [use about 10 drops]
3. 0.02 M $KMnO_4$ [use one drop]
4. 0.01 M $MnSO_4$ [use one drop]
5. 24-well microplate
6. small test tube
7. stopwatch

Design an experiment to:

a. compare the rates of reaction for $Na_2C_2O_4$ with $KMnO_4$ and $FeSO_4$ with $KMnO_4$ [see **Technique**]
b. compare the rates of reaction for $Na_2C_2O_4$ with $KMnO_4$ with:
  (i) the same mixture with a 50% dilution of the $Na_2C_2O_4$
  (ii) the undiluted mixture at a higher temperature
  (iii) the undiluted mixture with a drop of $MnSO_4$ added before the $KMnO_4$ is added
  (iv) the original *reacted* mixture from (a) with another drop of $KMnO_4$ added [oxalate is in great excess, so there will be some left]

Technique

1. how much to use?

This is a problem in any experiment. Taking into account factors like cost and environmental impact, the answer is always the same: as little as necessary. In this experiment you can easily observe the reactions by using about 10 drops of either the oxalate or the iron(II) solutions (this should be enough to cover the bottom of the well completely--if not, use a little more). Consistency is also important because you are comparing several different factors and seeing how they affect the rate of reaction. So close to the same amount each time eliminates one very important variable.

2. using a control for comparing rates

Both reactions are followed by the disappearance of the purple color of the MnO_4^- ion. You may find it helpful to make a "control" mixture of water and one drop of $KMnO_4$ in one well so that you have something to compare with the color of the reacting solutions.

3. heating

Placing some solution in a small test tube and standing it in a small beaker of the hottest tap water for a few minutes should work fine. Just pour the heated solution into one of the empty wells and proceed as you would at room temperature.

Analysis

1. Both of the reactions observed in this experiments are redox reactions. In one, iron(II) ions are oxidized to iron(III) ions by permanganate ions which are reduced to manganese(II) ions [in acidic solution]. In the other, oxalate ions ($C_2O_4^{2-}$) are oxidized to carbon dioxide by permanganate ions which are reduced to manganese(II) ions [again, in acidic solution]. Write balanced equations for these processes, showing your balanced half-reactions.

2. The "nature of the reactants" is often given as one factor that affects the rate of a chemical process. In both of these reactions, MnO_4^- is converted to Mn^{2+}. But the oxidations are very different. In light of collision theory, and comparing the oxidation half-reactions, suggest a reason for the difference in the observed rates of these two processes.

3. Explain briefly the observed difference between the undiluted and diluted $C_2O_4^{2-}$ rate with reference to collision theory.

4. Explain briefly the observed difference between the room temperature and heated reactions with reference to collision theory.

5. Substances which increase the rate of a reaction without being consumed are known as *catalysts*. In the oxalate/permanganate system, Mn^{2+} acts as an *autocatalyst* (i.e., it is a product which catalyzes its own reaction). What experimental evidence is there that Mn^{2+} is a catalyst for this reaction system?

6. In some reaction systems, surface area (for solids) and pressure (for gases) are often cited as factors which also affect the rate of reaction. Explain briefly, with reference to collision theory, why these should be included with the effects observed in this experiment