Chemistry Excel Workshops (Spring 2015)

Excel Spreadsheet Tools that will be discussed:
- Data entry
- Use of built-in functions (AVG, LN, STDEV)
- Defining cells with algebraic formulas
  - Variable cell (i.e., F5) vs. same cell ($F$5)
  - $\Delta C = C_2 - C_1$ & $\Delta t = t_2 - t_1$, where $C$ = concentration & $t$ = time
  - Rate = $-(\Delta C/\Delta t) = k*C$; $\therefore k = -(\Delta C/\Delta t)/C$
- Graphing data (Are dependent and independent variables on correct axes?)
  - Dependent and independent variables on correct axes
  - Labeling graph & axes
  - Including multiple series of data in the same graph and labeling each
- Trendline fit
  - Linear (standard deviation of the slope)
  - Exponential
- Interpreting equations for trendlines
  - Changing significant figures displayed
  - Changing format between number & scientific notation

**Exercise #1** covers material relevant to CHM111 – entering and graphing data in Excel and determining a line of best fit. *This material is covered in the first part of the video.*

**Exercise #2** covers material relevant to CHM112 – entering functions in Excel and determining kinetic rate laws. *This material begins at 25:40 in the video.*

**Exercise #1:** When an aqueous solution of copper sulfate (10. mL, 0.20 M) and an aqueous solution of potassium iodide (20. mL, 0.18 M) are combined, a precipitate forms. After filtering and drying, 0.43 g of the product is dissolved in 100 mL water. The calibration data are shown in Table I.

**Additional Exercises to Review some CHM111 material**
- What is the mass % Cu in the sample?
- What is the concentration in units of molarity (M)?
- Identify the product by structure and name (both common and systematic).
- Provide a balanced reaction for the synthesis.
- Determine the following: Limiting reagent, Theoretical yield, and % yield for this reaction.

<table>
<thead>
<tr>
<th>Concentration (g Cu/100 mL)</th>
<th>Absorbance at 610 nm</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.1807</td>
<td>1.370</td>
</tr>
<tr>
<td>0.1084</td>
<td>0.796</td>
</tr>
<tr>
<td>0.0651</td>
<td>0.456</td>
</tr>
<tr>
<td>0.0390</td>
<td>0.340</td>
</tr>
<tr>
<td>0.0234</td>
<td>0.178</td>
</tr>
</tbody>
</table>

*Table I.* Calibration data for concentration of copper with respect to absorbance of 610 nm light.
Exercise #2: Heating ethyl chloride at a temperature of 650 °C results in thermal decomposition to ethylene and hydrogen chloride. The concentration versus time data for three trials are given in Table II.

**During the workshop**

- Make three plots to discover the order of the reaction, including all 3 trials in each graph as separate series
  - [Ethyl chloride] vs. time
  - ln ([Ethyl chloride]) vs. time
  - 1/[Ethyl chloride] vs. time
  [Refer to page 360 of Yoder text, Chapter Summary, Points 5 & 6]
- Graph data.
  - Label the graphs, axes, and each series.
  - Generate a trendline for each series.
  - Determine the average and standard deviation of the slope of the graph with the most linear lines (R² closest to 1.0). [This value would be the rate constant for the reaction.]

**Additional Exercises to Review CHM111 and for Kinetics coverage in CHM112**

- What is the balanced equation for the reaction?
- Draw the Lewis Structures of the reactants and the products of this reaction.
- What is the rate law for this reaction?
- What are the units of the rate constant?
- Calculate the individual rate constants for various intervals. Compare the average of these values with the specific rate constant. Calculate the standard error of the mean.
- What is the half-life of ethyl chloride under these conditions?

<table>
<thead>
<tr>
<th>Time (sec)</th>
<th>Trial 1</th>
<th>Trial 2</th>
<th>Trial 3</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>0.1010</td>
<td>0.1051</td>
<td>0.1102</td>
</tr>
<tr>
<td>10</td>
<td>0.0810</td>
<td>0.0738</td>
<td>0.0702</td>
</tr>
<tr>
<td>20</td>
<td>0.0532</td>
<td>0.0492</td>
<td>0.0470</td>
</tr>
<tr>
<td>30</td>
<td>0.0422</td>
<td>0.0381</td>
<td>0.0414</td>
</tr>
<tr>
<td>40</td>
<td>0.0282</td>
<td>0.0278</td>
<td>0.0288</td>
</tr>
<tr>
<td>50</td>
<td>0.0206</td>
<td>0.0188</td>
<td>0.0186</td>
</tr>
</tbody>
</table>

**Table II.** Results from three trials monitoring the decomposition of Ethyl chloride at 650 °C.