**Significant Digits**

The examples below illustrate the proper use of significant digits for the Chemistry 30 Diploma Examination in a response.

**Example 1**

A 10.0 mL sample of a Fe$^{2+}$(aq) solution of unknown concentration is titrated with a standardized 0.120 mol/L KMnO$_4$(aq) solution. The following data are recorded.

<table>
<thead>
<tr>
<th>Trial</th>
<th>I</th>
<th>II</th>
<th>III</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette reading (mL)</td>
<td>10.10</td>
<td>19.22</td>
<td>28.33</td>
</tr>
<tr>
<td>Initial burette reading (mL)</td>
<td>1.00</td>
<td>10.10</td>
<td>19.22</td>
</tr>
<tr>
<td>Titrant added (mL)</td>
<td><em>9.10</em></td>
<td>9.12</td>
<td>9.11</td>
</tr>
</tbody>
</table>

The concentration of the Fe$^{2+}$(aq) is ________.

**Reaction Equation**

MnO$_4^{-}$(aq) + 8 H$^{+}$(aq) + 5 Fe$^{2+}$(aq) → Mn$^{2+}$(aq) + 4 H$_2$O(l) + 5 Fe$^{3+}$(aq)

Average volume of titrant added is 9.11 mL

\[ [\text{Fe}^{2+} \text{(aq)}] = 9.11 \text{ mL MnO}_4^{-} \text{(aq)} \times 0.120 \text{ mol/L MnO}_4^{-} \text{(aq)} \times \frac{5 \text{ mol Fe}^{2+} \text{(aq)}}{1 \text{ mol MnO}_4^{-} \text{(aq)}} \times \frac{1}{10.0 \text{ mL Fe}^{2+} \text{(aq)}} \]

\[ [\text{Fe}^{2+} \text{(aq)}] = 0.547 \text{ mol/L} \]

* Final answer has 3 significant digits (least number present according to the multiplication/division rule)

**Example 2**

The pH of a 0.100 mol/L solution of ethanoic acid is ________.

\[ K_a = 1.8 \times 10^{-5} = \frac{x^2}{(0.100 \text{ mol/L} - x)} \]

The value of x can be ignored when compared to 0.100 mol/L in the case of such a weak acid.

\[ K_a \text{ is approximately } \frac{x^2}{0.100 \text{ mol/L}} \]

\[ x = [\text{H}_3\text{O}^+(\text{aq})] = 0.001342 \]

\[ \text{pH} = -\log(0.001342 \text{ mol/L}) = 2.87 \]

Additional digits carried through on an interim basis

Final answer has 2 significant digits
Example 3

A student conducts a calorimetry experiment to determine the energy transferred when Solution A is mixed with Solution B. The data collected are shown below. Assume the specific heat capacity for each solution is the same as that of water.

<table>
<thead>
<tr>
<th>Mass of Solution A (g)</th>
<th>100.0</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of Solution B (g)</td>
<td>100.0</td>
</tr>
<tr>
<td>Mass of final solution mixture (g)</td>
<td>200.0</td>
</tr>
<tr>
<td>Initial temperature of solutions A and B (°C)</td>
<td>20.0</td>
</tr>
<tr>
<td>Final temperature of the solution mixture (°C)</td>
<td>23.0</td>
</tr>
</tbody>
</table>

\[ \Delta H = mc\Delta t \]
\[ \Delta H = (200.0 \text{ g})(4.19 \text{ J/g.°C})(3.0 \text{ °C}) \]

The original data are limited to 3 significant digits.

\[ \Delta H = 2.5 \text{ kJ} \]

The resulting temperature has 2 significant digits.

The final answer should be rounded to the same number of significant digits contained in the input data for the calculation \( \Delta H = mc\Delta t \) that has the fewest number of significant digits.

The final answer has 2 significant digits because the input data for the \( \Delta H = mc\Delta t \) calculation is limited by the temperature difference of 3.0 °C, which has 2 significant digits.