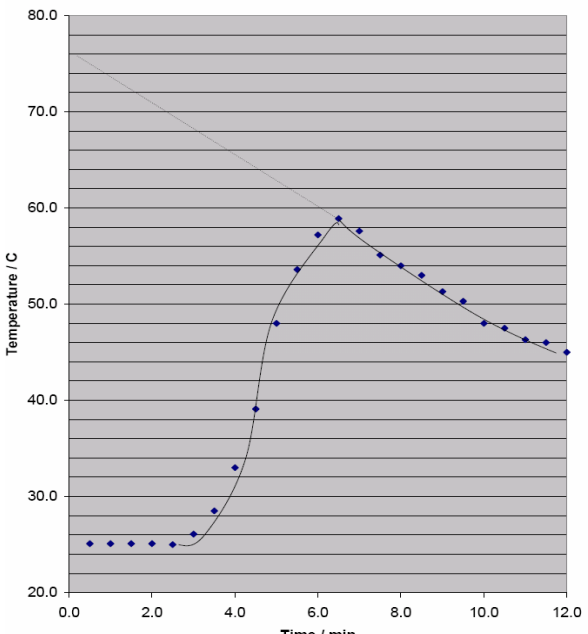


Sample Conclusions

Example 1

Report	Evaluation
<p>Conclusion & Evaluation</p> <p>The relationship between the concentration of the solution and the moles of molecules absorbing the light should be directly proportional. Thus, an increase of concentration increases proportionately the absorbency of light (assuming that all other variables are held constant). This is according to the Beer-Lambert Law:</p> $A = \epsilon C L$ <p>where A is the absorbance and C the concentration. (ϵ is molar extinction coefficient and L is path length). Overall, the data generated from the experiment appears to support this relationship and the theory behind it. As we can see from the graph there is a clear directly proportional linear relationship between absorbency (A) and concentration (C). The trend line crosses the y-axis at 0, so there appears to be no systematic error. In the hypothesis it was stated that according to the prediction the graph would approximately resemble the graph drawn below (to the right). The real graph is shown below (to the left). They are very similar:</p>	<p>Concluding</p> <p>Complete</p> <p>A valid conclusion was reached based on the data gathered and processed. The student correctly noted that the graph is linear within the limits of random error. The conclusion is correctly justified through reference to accepted theory.</p>
<p style="text-align: center;">Graph: The Relationship between Concentration and Absorbency</p> <div style="display: flex; justify-content: space-around;"> <div data-bbox="162 735 974 1302"> <p>Y-axis: Absorbency</p> <p style="text-align: right;">$y = 1.008x - 0.001$</p> <p style="text-align: center;">Y-axis: Concentration of $\text{CuSO}_4(\text{aq})$ [mol dm^{-3}]</p> </div> <div data-bbox="1023 924 1494 1260"> <p style="text-align: center;">Y-INTERCEPT = 0</p> <p style="text-align: right;">GRADIENT = ϵL</p> </div> </div>	
<p>The accuracy of the real results does not detract from the probable veracity of this conclusion. The errors are relatively small and appear to be random. Even so the trend line still falls within the error bars on the graph - a testament to the remarkable accuracy of the results obtained in this investigation. Of course, these small random errors are not ideal, although more repetitions of the data would most likely have eliminated these after multiple repetitions. These are probably small human errors which would not be eliminated by altering the apparatus used. In all this experiment was a success as we could draw firm conclusions from the data generated.</p>	<p>Evaluating procedure</p> <p>Partial</p> <p>The results are evaluated and few weaknesses were identified. The uneven spread of data points or limited range of concentrations (copper (II) sulfate can be made up to nearly 2 M at room temperature) should have been commented on. The student has seen the linear relationship confirmed and has then ceased to look critically.</p>
<p>Further work could be done to expand our knowledge of transition metal complexes and their light absorbing qualities. Variables such as path length could be varied in further experiments to confirm the linear relationship with absorbency as predicted by the Beer-Lambert Law. It would also be interesting to take results for different wavelengths of filter in the colorimeter for a fixed concentration in order to determine an approximate, if not exact range for the wavelengths absorbed by complex ions. Work could also be done to investigate the effects of using different solutes to dissolve the copper sulfate or other transition metal salts in. In this way we could find out the relative ability of ligands to split d-orbitals.</p>	<p>Improving the investigation</p> <p>Partial</p> <p>The student suggested further research away from the focused research question under consideration rather than modifications to reduce random error (only suggesting repeats) or to increase the range over which conclusions could be drawn.</p>

Example 2

Report	Evaluation
<p style="text-align: center;">Time vs Temp. of $\text{CuSO}_4(\text{aq}) + \text{Zn}(\text{s})$ Reaction</p>  <p>Conclusion and Evaluation:</p> <p>The purpose of this investigation was achieved in that the value of ΔH for the reaction between $\text{CuSO}_4(\text{aq})$ and $\text{Zn}(\text{s})$ was found to be -213 kJ/mol. To account for the error due to heat loss, the graph of temperature vs. time was drawn. The rate of cooling was extrapolated using a line of best fit, which was then extended back to the time of mixing to determine the true rise in temperature. The graph also included values for temperature of the solution for three minutes before the time of mixing to ensure that the CuSO_4 was at a constant temperature.</p>	<p>Processing raw data Partial</p> <p>The student has made a major error in processing. He extrapolated the temperature back (although the trend line does not appear to be completely appropriate which in itself could reduce the award to partial) but then based ΔT on the temperature at $t=0$ mins, whereas it should have been based on the temperature at $t=3$ mins, that is, when mixing occurred. Although this is a “major” error (as stated in the aspect descriptor) the subsequent calculations to calculate ΔH should be rewarded.</p> <p>Concluding Complete</p> <p>The conclusion is the final experimental value for ΔH and the justification for its veracity is the quantitative comparison with the literature value, that is, the percentage experimental error (it is important that the teacher’s instructions did not direct the student to do this), and the conclusion as to whether the percentage error is explained by solely random error or whether systematic error is implied. For this type of investigation where a quantitative property is being experimentally determined this discussion of random and systematic error is required for complete to be awarded. The only concerns are that the student incorrectly switched the accepted and experimental values when calculating the percentage error and also did not cite the source of the literature value, which would have been the proper academic practice. However, there is more than sufficient evidence to award complete.</p>
<p>However, even though the percent uncertainty in the value obtained was only 0.79%, the percent difference when compared with the accepted value (-217 kJ/mol) was 1.84%. Clearly, since percent difference is greater than percent uncertainty, random error alone cannot account for the discrepancy between the two values. There was a systematic error in assuming the density of the solution to be equal to 1.00 g/ml, such that 20.0 ml of solution had a mass of 20.0 g. Also, there was systematic error in assuming that the specific heat of the solution was $4.184 \text{ J/g}^\circ\text{C}$. Since both assumptions are rough approximations at best, these assumptions skewed the experimental value.</p>	<p>Evaluating procedure Complete</p> <p>The student has not simply listed the weaknesses but shows an appreciation of their importance. Challenging the initial assumptions regarding density and specific heat capacity is a good example of critical thinking and compensates for missing the more obvious heat losses to the external environment.</p>
<p>In the future, the mass and the specific heat of $\text{CuSO}_4(\text{aq})$ should be measured directly to account for these two errors. Also, despite extrapolating with the graph, there were heat losses that were not accounted for in the procedure. The heat produced by the reaction was absorbed not only by the 20.0g of solution, but also by the solid metals on the bottom of the calorimeter (the excess zinc and the precipitated copper), and by the calorimeter (the paper cup). However, the latter two heat losses were not accounted for in calculating Q. Thus, Q was calculated to be lower than the true value because these two absorptions of heat were ignored. The error in Q was propagated when calculating ΔH, and caused it to become smaller than the accepted value. In the future, to account for this error, the mass and specific heats of the solid metals should be included, or calculated and accounted for. Less excess zinc could be added to minimize the amount of metals on the bottom of the cup absorbing heat. Also, a styrofoam cup would serve as a better calorimeter because its excellent insulating properties would ensure that very little heat would be absorbed by it. Also, the results should be duplicated in order to improve the quality of measurement to obtain more precise value.</p>	<p>Improving the investigation Complete</p> <p>The improvements relate to the evaluation of the procedure and are realistic. A more precise suggestion as to how to calculate specific heat capacity of the copper sulfate solution would have been appropriate but overall there is sufficient evidence to say that the aspect has been fulfilled.</p>

Example 3

Report	Evaluation
<p>Conclusion and Evaluation</p> <p>The experimental value of pKa was 4.67 (+/- 0.02) and the literature value (<i>source: IB Chemistry Data Book, Table 16</i>) is 4.76</p> <p>The % error is $[(4.67 - 4.76)/4.76] * 100 = 1.89 = 2\%$</p> <p>This error is only a bit higher than the total random error of the experiment so I am pleased with my result. There must have been some systematic errors though.</p>	<p>Concluding</p> <p>Complete</p> <p>The conclusion is the numerical result for the pKa value along with the comparison with the literature value. It is good practice that the source of the literature value has been given. The justification of the conclusion is the discussion of whether the deviation from the literature value is explainable solely by random errors or whether systematic errors have also taken place. This has been done.</p>
<p>The biggest problem was with the dropping syringe taps blocking. Sometimes the flow rate slowed right down. This is a major problem because the drop sizes might have changed which would mean my volume calibration was not exactly right. Changing the ethanoic acid solution should not have affected the result because pKa does not change with acid concentration.</p>	<p>Evaluating procedure</p> <p>Partial</p> <p>The important issue of the reliability of the calibration due to the tap blocking has been raised. However, the other system errors could have been discussed such as whether the drop size varies as the liquid level in the drop counter syringe (the burette) gets lower or whether the tap position gets moved during the experiment. The teacher present during the action phase will be best able to judge other reasonable sources of error.</p>
<p>To modify the experiment I would clean the dropping syringe very well and make sure the glassware is very clean for making up the ethanoic acid. I would also do the drop calibration up to a volume of 10 cm³ because a bigger collected volume and more drops will reduce the uncertainty in the volume readings.</p>	<p>Improving the investigation</p> <p>Partial</p> <p>A valid improvement regarding calibration is given, however, little else was offered. For example, in determining a numerical quantity such as pKa the number of repeats is important since it gives a clear empirical basis for uncertainty and reproducibility. This improvement and some other possibilities were missed.</p>

Example 4

Report	Evaluation
<p>Conclusion: As expected the solvents showed major cooling which is due to the high energy particles evaporating through the liquid surface. The more volatile liquid propanone, which has weaker dipole-dipole intermolecular forces than the hydrogen bonding in ethanol, showed greater cooling since there is a greater proportion of particles with enough energy to break the surface and evaporate. We can see the effect of composition of solvent mixtures in graphs 1 and 2. It appears that the relationship between molar composition and degree of evaporative cooling in the propanone/ethanol mixture is not linear and is a shallow negative curve. The R2 correlation coefficient of 0.96 is very close to the 0.95 of the linear plot. Because of the random error in plotting the points (about 2% of the molar composition and +/- 0.6 °C on the temperature scale), and because the repeat run of pure ethanol showed a difference of 1.8°C indicating some systematic error, this means that these relative correlations are too close to confidently whether the relationship is linear or not.</p>	<p>Concluding Complete</p> <p>The statement in the subject guide’s clarification to this aspect, that “analysis may include comparisons of different graphs or descriptions of trends shown in graphs”, is well demonstrated. The student also meets the clarification’s requirement that “the conclusion must take into account any systematic or random errors and uncertainties”.</p>
<p>Evaluation of Experimental Procedure: The estimated random error does not completely explain the error which appeared when I repeated the pure ethanol measurement. So there were probably sources of systematic error. One source of systematic error related to the data logger temperature readings. By the time I had removed the two probes from the solvent and clamped them and then pressed start on the software the temperature had already started dropping. This gave an error in the maximum temperature meaning that all values for the amount of cooling were too low. Also the data logger stopped collecting data in some trials just before the minimum temperature was reached. These errors can be reduced by clamping probes and starting data logger before immersing them in the liquid. As the probes are gently drawn out of the solvent the software will detect the maximum temperature. If I had enough time I would have repeated all trials with a longer time setting on the software so that I would definitely have recorded the minimum temperature. Another source of error is that I assumed that the amount of liquid in each drop was the same. This is important because maybe a greater amount of liquid will evaporate for longer and will then allow time for more cooling to occur. Although it will be difficult to control drop size I could test whether this is an important control variable by wrapping twice as much tissue paper around the probe for one sample of pure propanol so I could see if more liquid made a big difference. Another error is that the ethanol may not have been pure. I used the supplied stock bottle but that probably was only 95% ethanol with 5% water. This would reduce the ethanol composition in the mixture and introduce a third liquid, water, into the mixture. This source of error can be corrected by using 99% ethanol. I think other sources of system error were small. The cardboard box should have stopped draughts, I took great care in exposing the same area of tissue paper and the trials were all carried out in one lesson so room temperature should have been constant. The experiment would also have benefited from testing each sample more than once and also by preparing samples of more different compositions so my data points would have covered a greater range of composition values.</p>	<p>Evaluating procedure Complete</p> <p>The student has identified three appropriate sources of systematic error.</p> <p>Improving the investigation Complete</p> <p>The student suggested modifications to the three sources of systematic error. The student also has addressed a clear failing of the original data, its poor distribution of compositions, and since clarification to aspect 3 states that the “data range can be addressed here” it is only now that aspect 3 can be considered complete. In other cases, where a more appropriate range of data had been collected, this statement may not have been necessary and the complete could have been awarded without it. The data directs the marking of this criterion.</p>