



Student Performance Q&A: 2013 AP[®] Chemistry Free-Response Questions

The following comments on the 2013 free-response questions for AP[®] Chemistry were written by the Chief Reader, Larry Funck of Wheaton College, Wheaton, Illinois. They give an overview of each free-response question and of how students performed on the question, including typical student errors. General comments regarding the skills and content that students frequently have the most problems with are included. Some suggestions for improving student performance in these areas are also provided. Teachers are encouraged to attend a College Board workshop to learn strategies for improving student performance in specific areas.

Question 1

What was the intent of this question?

This equilibrium question focused on solubility chemistry and consists of three components: foundational knowledge regarding solubility chemistry, selective precipitation, and factors that can influence the solubility equilibrium. In part (a)(i) students wrote a chemical equation to represent the solubility reaction. Part (a)(ii) required stoichiometric knowledge and the ability to manipulate concentration calculations. In part (a)(iii) students used knowledge of solubility-product chemistry to calculate the equilibrium constant (K_{sp}) for the dissolution reaction. In part (b)(i) students judged the relative solubility of salts based on the values of solubility product constants. In part (b)(ii) students calculated the minimum F^- concentration necessary to initiate precipitation of CaF_2 using the solubility-product constant and a given value for $[Ca^{2+}]$. In part (b)(iii) students calculated the volume of NaF solution at a given concentration to be added to start the precipitation of CaF_2 . In part (c) students described a process, based on knowledge of what factors influence equilibria, to redissolve the precipitate formed in part (b).

How well did students perform on this question?

The mean score was 3.59 out of a possible 10 points. Students generally attempted all parts of the question. If parts were omitted it was usually parts (b)(ii) and (b)(iii). Of the points missed, the most common were parts (a)(i), (a)(ii), (b)(ii), and (b)(iii). Approximately 13.9 percent of students either did not address the question or earned no points for their responses, while 3.2 percent earned a perfect score of 10.

What were common student errors or omissions?

In part (a)(i) students wrote a chemical equation to represent the dissolving of MgF_2 . Common errors included not writing the following:

- A chemical equation that showed solubility: many students wrote a double displacement reaction ($\text{MgF}_2 + \text{H}_2\text{O} \rightarrow \text{MgO} + 2 \text{HF}$) that is not related to the dissolving of MgF_2 .
 - Students did not relate the statement about dissolving MgF_2 to solubility and therefore showed a chemical equation that was not representative: this also was true when they were working part (a)(iii), as they did not always relate the chemical equation to K_{sp}
- Equations that were both mass and charge balanced.
 - Forgetting charges on the ions
 - Using incorrect charges (mostly for Mg; using Mg^{1+} instead of Mg^{2+})
 - Including water on one side or both sides (a balanced net-ionic equation would not have included water)
 - Not balancing mass: many students wrote the equation with only one mole of F^-
- Phase symbols into the equation: using phase symbols in the equation can provide important information related to the conceptual and mathematical aspects of the solubility process.

In part (a)(ii) students calculated moles of MgF_2 , which required stoichiometric knowledge and the ability to manipulate concentration calculations. Common errors included the following:

- Assuming all of the MgF_2 dissolved.
 - Based on that assumption students completed a simple grams-to-moles calculation using molar mass to calculate the total moles of MgF_2 but did not consider that not all of the MgF_2 dissolved. The major error in this part was basing the answer entirely on this calculation
- Confusing moles and molarity: – students should have a better understanding of the differences between these concepts. Many students could not calculate the correct molar mass for MgF_2 .
- Confusing the chemical species in the problem: students often confused moles or molarity of F^- with moles or molarity of MgF_2 and Mg^{2+} .

In part (a)(iii) students used knowledge of solubility-product chemistry to calculate the solubility-product constant (K_{sp}) for the dissolution. Common errors included the following:

- Not knowing what K_{sp} meant and then writing incorrect equilibrium expressions.
 - For example, students were likely to write an expression for their reaction that included everything in the reaction, such as using the equation:



instead of writing a proper K_{sp} expression. If the students use phase symbols with the chemical equation and know what K_{sp} means, the solubility product expression would more likely include the appropriate species in the aqueous phase.

- Incorrectly calculating $[\text{Mg}^{2+}]$ when calculating K_{sp} .
- **Very important:** The given $[\text{F}^-]$ was not used when calculating K_{sp} .
 - Students need to read the problem carefully and interpret the given values.

In part (b)(i) students judged the relative solubility of salts based on the values of solubility product constants. Common errors included the following:

- Not relating a low K_{sp} value with a salt being less soluble; some students did not know that precipitation is related to the K_{sp} value.
- Confusing precipitating versus dissolving (“it precipitates first because it’s more soluble”).
- Confusing the exponents in the determination of which is a larger number.
 - Some students think that $3.5 \times 10^{-11} > 1.8 \times 10^{-6}$
- Making statements about the rate at which the salt will dissolve based on the equilibrium constant. For example, students stated that the more soluble the salt, the faster it dissolves. Equilibrium is not related to the rate at which a reaction occurs.

In part (b)(ii) students calculated the minimum values of F^- concentrations to initiate CaF_2 precipitation using solubility-product constants and a given value for $[Ca^{2+}]$. Common errors included the following:

- The given $[Ca^{2+}]$ in the problem was not used: Many students set up a problem involving the molar solubility, but inclusion of the $[Ca^{2+}]$ is necessary to calculate the minimum $[F^-]$.

In part (b)(iii) students calculated the volume of NaF solution at a given concentration that could be added to start the precipitation of CaF_2 . Common errors included the following:

- Not using the minimum $[F^-]$ calculated in part (b)(ii).
 - In this part, students incorrectly used the metal ion concentration.

In part (c) students described a process, based on their knowledge of how factors influence equilibria, to redissolve the precipitate formed in part (b). Common errors included the following:

- Using terminology inappropriately as part of the justification (for example, heating the precipitate until it melts to help it dissolve).
- Making vague descriptions.
 - Students should use concise wording to directly justify an answer
- Not understanding the chemistry of the dissolving process. Dissolving is a disturbance of the equilibrium process which can occur in a variety of ways.

Based on your experience of student responses at the AP® Reading, what message would you like to send to teachers that might help them to improve the performance of their students on the exam?

- Construct knowledge about how conceptual chemistry topics connect to the mathematics of the topic. Many times it was clear that students could complete the mathematics, but there was no connection to the chemical concepts of the question.
 - Students wrote an incorrect chemical equation in part (a)(i) when asked to write an equation about dissolving a salt, then wrote the correct equation (or complete the correct calculation) when asked to solve for K_{sp} .
 - **Very important:** Base the calculations and explanations on the chemical equation; dissolving an ionic compound in water and the solubility-product refer to the same process.
 - Clarify that “soluble” or “insoluble” compounds represents a spectrum of solubility.
 - Discuss that solubility is an equilibrium described by K_{sp} (and the molar solubility for all compounds is on a continuum from very small to very large).

- Students wrote a K_{sp} equation without recognizing the involvement of ions
 - Solubility produces ions in solution which is described by K_{sp} .
 - Phase symbols in equations would help students to make this connection.
- Understand what a net-ionic equation means; not only is there a balance of mass, but also a balance of charges.
- Understand concepts, not just be able to complete the calculation involved.
 - Emphasize what K_{sp} means and how it relates to the chemical equation, solubility, precipitation
 - K_{sp} needs to be included as an equilibrium topic along with buffers.
- Explain concepts: they do better on the mathematical portions of a question. Keep asking students explanation questions so they can practice short, concise descriptions of concepts.
- Continue to practice with mole and concentration calculations.
- Practice explaining shifts in equilibria.
- Practice decoding questions and expressing explanations.
 - Reading comprehension practice (circle important values/questions)
 - Use concise wording to express ideas.
- Show **all** mathematical work (i.e., setups).
- Experiment with redissolving of a not-very-soluble (not insoluble) salt.
 - Observe the impact of adding an acid, base, complexing material, or heat to the solubility of a salt.
 - Observe the relationship between K_{sp} and ΔG .
 - Observe the temperature dependence of K_{sp} .

Question 2

What was the intent of this question?

This question assessed the students' ability to solve stoichiometry, electrochemistry, and thermodynamics problems and understand relationships between these concepts. Parts (a) through (d) involve answering questions about an electrolytic cell. In part (a) students calculated the number of moles of electrons transferred when a given mass of Al was produced in an electrolytic cell. Students used knowledge of stoichiometry and oxidation-reduction reactions to calculate this value. In part (b) students calculated the amount of time required to produce the Al in part (a) when given the current. This required an understanding of the concepts of charge and current. Part (c) asked students to calculate the volume of CO_2 gas produced in the process at a specific temperature and pressure. This part of the question assessed student understanding of stoichiometry and gas laws. Part (d) required students to explain the necessity for Al_2O_3 to be molten. This part evaluated students' understanding of ion mobility in an electrolytic cell. Part (e) asked students to calculate thermodynamic quantities for a different reaction. In part (e)(i) students calculated the value of E° for the reaction using the table of standard reduction potentials. This assessed student understanding of half-cell potentials. In part (e)(ii) students calculated the standard free energy change for the reaction, demonstrating the ability to recognize total moles of electrons transferred and to apply an appropriate relationship between electrochemical and thermodynamic quantities.

How well did students perform on this question?

Students did reasonably well on this question, earning a mean score of 4.09 of 10 possible points, showing a broad distribution with many students earning 4 or 5 points. Generally most students attempted all parts of the question. Ten percent of the students did not address the question or earned no points, while 15 percent of the responses earned a score of 8, 9, or 10 points. The most commonly missed points occurred in parts (b), (d), and (e)(ii). The large percentage of students earning a score of 5 on this question typically earned those points in parts (a), (c), and (e)(i).

Many students were successful in part (a) of this question and earned 1 point for calculating moles of Al and 1 point for multiplying this value by 3 moles of electrons transferred. Part (b) proved more challenging for students; the relationship between the number of electrons and the corresponding charge seemed to be one of the more difficult concepts. Students often were unable to calculate charge and did not earn the first point. Some students, however, incorrectly equated moles of electrons from part (a) with charge, substituted correctly, and solved for time earning 1 point. Part (c) was the most accessible part of the question, with many responses earning 2 points: 1 point for calculating moles of CO₂ and 1 point for calculating volume of CO₂. Part (d) was the most difficult point for students to earn. Students were required to discuss ion mobility to earn credit and many stated that the liquid allows current to flow or incorrectly explained that electrons were flowing through the solution. Many students calculated voltage correctly and earned 1 point for part (e)(i). Many students earned 2 points in part (e)(ii) for recognizing that six moles of electrons were transferred and correctly calculating the change in the standard free energy. Students who earned only 1 point in this part often listed the incorrect sign for ΔG° or calculated the answer in J/mol_{rxn} but incorrectly identified the value as kJ/mol_{rxn}.

What were common student errors or omissions?

Part (a)

- Identifying 3 or 12 moles as the number of electrons transferred ignoring the mass of Al.
- Multiplying or dividing the moles of Al by the coefficient from the balanced equation.

Part (b)

- Ignoring the answer in part (a) when calculating charge.
- Dividing moles of Al by current to obtain time.
- Using the charge of an electron instead of 96,500 C/(mol e⁻).

Part (c)

- Using the coefficient of 3 for the number of moles of CO₂ instead of using the 8.71 moles of Al and the mole ratio from the balanced equation to find the moles of CO₂.
- Substituting an incorrect “R” value.

Part (d)

- Stating that free electrons must flow through the solution.
- Restating the prompt.
- Mentioning electrical neutrality of a salt bridge.
- Suggesting that the cell was a voltaic cell instead of an electrolytic cell.

Part (e)(i)

- Not evaluating the value of the voltage to determine anode and cathode and subtracting “top” – “bottom” to calculate a negative voltage.
- Multiplying half-reactions **and voltages** to make electrons equal.

Part (e)(ii)

- Using an incorrect number of electrons.

- Ignoring units and reporting the calculated value as $\text{kJ/mol}_{\text{rxn}}$ when it was really $\text{J/mol}_{\text{rxn}}$.
- Leaving off the negative sign for the reported value.

Based on your experience of student responses at the AP® Reading, what message would you like to send to teachers that might help them to improve the performance of their students on the exam?

Suggestions for teachers:

- Compare electrolytic and voltaic cells when teaching electrochemistry.
- Allow students an opportunity to perform experiments with electrolytic and voltaic cells.
- Provide students the opportunity to explain the process of charge transfer on a particulate level focusing on electron flow and ion movement in each type of cell.
- Make and stress connections between electrochemistry and thermodynamics for students.
- Teach students to recognize the significance of the values for reduction potentials when determining overall cell voltage.
- Require students to write appropriate units throughout calculations.

Suggestions for students:

- Use terms correctly distinguishing clearly between electrons, ions, and molecules.
- Use unit labels throughout a calculation to avoid incorrect final units.
- Check for reasonable answers when using the calculator.
- Focus on understanding the concepts, not memorization.
- Show all of your work and clearly label answers.
- Become familiar with the calculator to be used on the exam before the day of the exam.
- When asked to explain, be sure to answer the question “Why?” and write clear, concise answers.
- Double-check the sign of answers when making calculations.

Question 3

What was the intent of this question?

This question used laboratory data and served as the question where rules for significant figures are applied. This question assessed ability to use collected experimental data for calculations involving stoichiometry, thermodynamics, error analysis, and reporting values with the correct number of significant digits. In part (a) students used experimental data to determine the limiting reactant for the reaction given. In part (b) students examined the data to determine which of the four trials contained inconsistent data and justified their choice. Part (c) required students to use the experimental data to determine the heat released (q) in the reaction. In part (d) students then used the data from part (c) to calculate the experimental value of ΔH° . In part (e) students calculated the accepted value of ΔH° from enthalpies of formation given in a table. Part (f) asked students to examine the experimental and accepted ΔH° values and a possible source of error to determine if that error could have accounted for the discrepancy between the accepted and experimental values.

How well did students perform on this question?

This question discriminated among the student responses, with a broad distribution of scores reflecting the range of understanding among the students’ papers. The average score was 3.10 out

of 9 points. This question had a high percentage of zero scores (combined “earned” zeroes and no-response zeroes) equal to 17 percent.

Students did reasonably well on part (a), earning one point for calculations that identified the limiting reactant.

Part (b) was challenging because it required students to compare the data from all trials and then choose the inconsistent trial. Students needed to recognize that the ratio of ΔT to g MgO was different in trial 1 when compared to the other three trials.

One of the most frequent errors in part (c) was determining the total mass of the solution in the calorimeter. Many students were able to earn the consistency point in part (c) for calculating q using the incorrect mass. In addition, part (c) had many significant figure and unit errors.

In part (d) students frequently earned 0 of the 2 possible points. A useful equation was not provided in the Equations and Constants pages of the exam. As a result, students often picked the equation $C_p = \Delta H/\Delta T$, which is not applicable here, and thus earned 0 points. Another common error was not recognizing $q_{\text{reaction}} = -q_{\text{calorimeter}}$. This error resulted in a positive value for ΔH° and the loss of 1 point.

The most frequently earned points were in part (e). Many students successfully calculated ΔH° from the enthalpies of formation given.

The point in part (f) was a difficult point to earn. Students did not relate their explanation to the numbers calculated in parts (d) and (e). Often students simply restated the information given in the question. Other students used arguments that confused experimental and accepted values.

What were common student errors or omissions?

In part (a) many students earned credit for concluding that MgO was the limiting reactant by either comparing the molar calculations using the stoichiometry of the reaction or by recognizing that the temperature change depended on the amount of MgO added.

Common errors included:

- Forgetting to use the reaction stoichiometry.
- Comparing the amounts of reactants and not the moles of reactant.
- Stating that the smaller number of grams was the limiting reactant.
- Stating that because grams of MgO (and not HCl) was changing without referencing the temperature change, then MgO was limiting.

The response in part (b) was independent of any of the other parts of the question and was often missed. Often students compared only trial 1 to trial 3 and this was insufficient to get the comparative argument for the ratio of ΔT to g MgO. Students needed to reference trials 2 and 4 in order to demonstrate that trial 1 had a ratio of ΔT to g MgO that was inconsistent with the other trials.

Common errors included:

- Selecting trial 1 and stating that the final temperature was lower than it should have been because the final temperature should be 28.1°C or 29.1°C like the other three trials.

- Selecting trial 1 because it has the smallest ΔT .
- Comparing trials 1 and 3 only and randomly picking one or the other without any comparison to trials 2 and 4.

In part (c) most students recognized they needed to use the formula $q = mc\Delta T$. However, many students made errors in substituting correct values into this formula. Students often missed 1 of the 2 possible points for incorrectly computing the total mass of the solution in the calorimeter. If the student used their computed mass for the calorimeter correctly with no math, significant figure, or unit errors, they could earn 1 point.

Common errors included:

- Using something other than $4.18 \text{ J}/(\text{g}\cdot^\circ\text{C})$ for the value of c .
- Using some random temperature value, such as average ΔT or the final temperature converted to Kelvin, for the value of ΔT .
- Recording the final answer with an unrounded number containing more than two significant figures.
- Recording the final answer with a math error.
- Recording the final answer with the wrong unit for q .
- Doing calculations for all trials when only one was required.
- Averaging masses for the mass of the calorimeter contents.
- Using the mass of MgO for the mass of the calorimeter contents.

Part (d) presented more challenges because the student needed to calculate ΔH° from the relationship: $-q_{\text{calorimeter}}/(\text{mol MgO})$. The student needed to recognize that the q of the reaction was equal in magnitude, but opposite in sign to the q of the calorimeter. If students used an incorrect setup or formula, and did not have moles calculated for the trial, they earned 0 of the 2 possible points. In addition, students needed to use the value in part (c) and use an appropriate mole calculation for the trial to earn points. Mathematical and significant figure errors were common.

Common errors included:

- Forgetting to record the calculated value of ΔH° with a negative sign for an exothermic reaction.
- Mixing up the data for the chosen trial. For example, students would use a q value from trial 2 but a mole calculation from trial 1.
- Using the formula from the exam, $C_p = \Delta H/\Delta T$, rearranging that formula to $\Delta H = C_p \cdot \Delta T$ and then calculating ΔH .
- Using the stoichiometric factor of 2 from the reaction for the number of moles in $\Delta H^\circ = -q_{\text{calorimeter}}/n$.
- Entering some random number for moles or ΔH° with no formula or no indication of where the number came from.
- Using the q value from part (c) and dividing it by 1,000 to get the final answer.

The points in part (e) could have been earned independently from any of the other parts. Students needed to know or be able to choose the correct formula from the Equations and Constants pages, substitute the ΔH_f° values from the given table into the formula, and then complete the calculation. This part required no interpretation of data. The points in part (e) were the most commonly earned points in the question.

Common errors included:

- Mixing up the order of the products and reactants when calculating ΔH_{rxn}°
Using $\Delta H_{rxn}^{\circ} = \Sigma\Delta H_f^{\circ}(\text{reactants}) - \Sigma\Delta H_f^{\circ}(\text{products})$ instead of
$$\Delta H_{rxn}^{\circ} = \Sigma\Delta H_f^{\circ}(\text{products}) - \Sigma\Delta H_f^{\circ}(\text{reactants})$$
- Computing the summation and subtraction incorrectly (math error).
- Using Hess's Law, but multiplying rather than adding the ΔH_f° values for the reactants and the products terms.
- Randomly entering the ΔH_f° values into Hess's Law which led to an incorrect answer.

Part (f) was the hardest point to earn in this question. The intent of the part of the question was to look at the calculated values in parts (d) and (e) and then decide if their values were consistent with the calorimeter leaking heat energy. This problem became very difficult for students when the experimental answer for part (d) was positive (endothermic) and the accepted value in part (e) was exothermic. The point was earned when students recognized the magnitude difference between the two values. Some students earned the point by recognizing the experimental value was endothermic so it would make no sense to have heat energy leak out of the calorimeter. Common errors included:

- Restating the question without providing evidence.
- Making no attempt to relate the differences in the values to the lower recorded value of ΔT .
- Not taking a stand as to whether the loss of heat would account for the discrepancy.
- Stating that yes, it would make a difference because the temperature change would be different, but not declaring if the temperature difference would be greater or less than the temperature difference in the experimental or accepted values of ΔH° .
- Confusing the experimental and accepted values of ΔH° .
- Confusing reference terms. Using ambiguous words such as calculated and actual rather than experimental and accepted. Terms like experimental, actual, or calculated paired with terms like accepted, theoretical, and literature earned credit.
- Making one statement of comparison and then contradicting the statement in the next sentence.

Based on your experience of student responses at the AP[®] Reading, what message would you like to send to teachers that might help them to improve the performance of their students on the exam?

Suggestions for teachers:

- Have students show all work, with units, and express answers with the correct number of significant figures.
- Use data throughout the course. Have students write claims and support those claims with evidence from the data.
- Tell the students to be specific with claims and justifications.
- Caution students against providing multiple answers to a question part. When one answer is requested, provide only one answer.
- Recommend that students circle the answer so it is easy to find.
- Do experiments in the laboratory so that students have experience.

Question 4

What was the intent of this question?

This question tested students' ability to convert the names of substances into chemical formulas, predict the products from a given set of reactants, and write balanced net-ionic equations. A follow-up question to each part assessed knowledge and understanding of fundamental concepts from the classroom and laboratory, namely performing solution stoichiometry calculations, rationalizing changes in solution pH, and justifying a difference in reaction rate.

How well did students perform on this question?

The mean score was 7.31 out of a possible 15 points. The distribution was somewhat unusual because it had three maxima, one at 5, a second at 9, and a third at 13–14. There does not appear to be any simple explanation for this unusual distribution curve. There were relatively few zeroes either through blank papers or failure to earn points on any part; the sum of these represented only 2.5 percent. Scores were generally higher on parts (a) and (b) than on part (c). The point earned most often was writing the formulas for the reactants carbon dioxide and water in part (b)(i). The least often earned point was the explanation of the slower reaction rate of zinc in acetic acid in part (c)(ii).

What were common student errors or omissions?

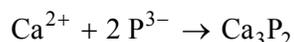
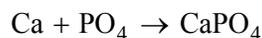
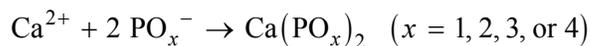
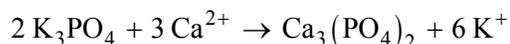
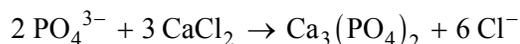
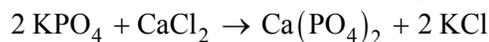
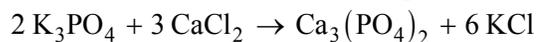
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- Providing a full or partial molecular equation instead of a net-ionic equation.
- Including extraneous charges on molecular reactants/products (e.g., H_2^+ or HCl^-).
- Retaining some or all of the spectator ions in the final equation.
- Failing to balance equations for both mass and charge.
- Using vague or nonspecific terminology in verbal answers.
- Leaving the reaction box blank, but answering the question elsewhere.

Part (a)(i)

- Providing a full or partial molecular equation instead of a net-ionic equation.
- Writing an incorrect symbol, formula, or charge for phosphate ion (e.g., Ph, P, PO, PO_2 , or PO_3 , with -1 , -2 , or -3 charge).
- Failing to identify the reaction as a precipitation reaction (i.e., omitting the calcium ion and/or the phosphate ion from the net-ionic equation).
- Omitting parentheses in the formula $\text{Ca}_3(\text{PO}_4)_2$.

- **Some common incorrect responses:**



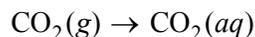
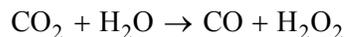
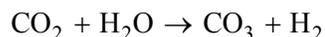
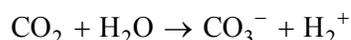
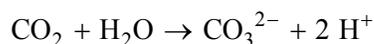
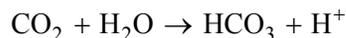
Part (a)(ii)

- Using the stoichiometric coefficients of products in the equation, rather than the volume and molarity of each solution of reactants, to calculate number of moles of product
e.g., $3 \text{Ca}^{2+} + 2 \text{PO}_4^{3-} \rightarrow \text{Ca}_3(\text{PO}_4)_2$ leading to “1 mole of product”
 $2 \text{K}_3\text{PO}_4 + 3 \text{CaCl}_2 \rightarrow \text{Ca}_3(\text{PO}_4)_2 + 6 \text{KCl}$ leading to “7 moles of product.”
- Calculating moles of product by adding together the number of moles of each reactant
e.g., $0.0020 \text{ mol K}_3\text{PO}_4 + 0.0030 \text{ mol CaCl}_2$ leading to “0.0050 mole of product.”
- Making computational errors by one or two orders of magnitude (e.g., calculating 0.10 mol or 0.010 mol of product formed, rather than 0.0010 mol).
- Ignoring or incorrectly calculating the ratio of Ca^{2+} to PO_4^{3-} that is necessary to make one mole of $\text{Ca}_3(\text{PO}_4)_2$.
- Missing mathematical justification to show the basis of the answer.

Part (b)(i)

- Omitting H_2O as a reactant or writing it over the reaction arrow.
- Writing an incorrect formula for carbon dioxide and/or water.
- Including an ionic charge for CO_2 and/or H_2O .
- Omitting or incorrectly assigning ionic charges on products, especially the bicarbonate ion.
- Dissociating H_2CO_3 completely into $2 \text{H}^+ + \text{CO}_3^{2-}$, which occurs only to a negligible extent.
- Proposing incorrect products, including C, CO, CO_3 , O_2 , OH^- , H_2 , and H_2O_2 .

- **Some common incorrect responses:**

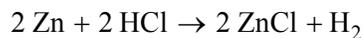
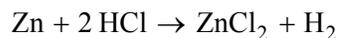
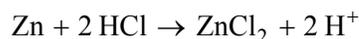
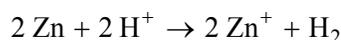


Part (b)(ii)

- Referring to “an acid” being produced in the reaction, or simply stating that acids generally lower the pH of a solution, without sufficient support. It was necessary to write an acidic product in the reaction box and/or mention carbonic acid by name specifically.
- Misinterpreting a lower solution pH as the production of a base or decrease in $[H^+]$.
- Employing circuitous mechanistic logic (such as CO_2 acting as an oxygen atom acceptor) to rationalize the change in pH.
- Using a dilution argument to explain the lower pH.
- Making a rationale based upon Lewis acidity.

Part (c)(i)

- Assigning an incorrect charge to the zinc metal and/or the zinc ion.
- Writing the strong acid HCl as a non-ionized species.
- Misidentifying the reaction as a precipitation reaction that forms $ZnCl_2$.
- **Some common incorrect responses:**

**Part (c)(ii)**

- Failing to state that the slower rate of reaction of Zn in acetic acid is a consequence of the reduced concentration of H^+ reactant.
- Identifying hydrochloric acid as a strong acid and acetic acid as a weak acid, but unsuccessfully connecting this concept to the reaction rate being dependent upon $[H^+]$.
- Mentioning that 1.0 M HCl has a lower pH than 1.0 M acetic acid, but without describing the importance of H^+ as a reactant.
- Identifying chloride and/or acetate ions as being relevant to the reaction.
- Proposing that differences in bond strength, molecule polarity, solubility, molecular weight, or rate of acid ionization explain the slower rate of the Zn reaction in acetic acid.

Based on your experience of student responses at the AP[®] Reading, what message would you like to send to teachers that might help them to improve the performance of their students on the exam?

- Encourage students to commit the formula and charge of common ions (especially polyatomic ones) to memory. The chloride ion is not Cl_2^- or ClO_3^- , for example, nor is P an acceptable abbreviation for the phosphate ion.
- Follow the conventions of superscripts and subscripts in chemical formulas. Unorthodox notations like H^2 or Ca_{2+} are confusing.
- Require students to write a balanced net-ionic equation for reactions that they perform, observe, or read about. Provide the reactants in the chemical equation but ask the students to predict the product(s) and write a complete chemical equation. Challenge students to apply fundamental logic and pattern recognition to reactions that they have never seen

before. Practice repeatedly throughout the course to develop the skill of thinking like a chemist.

- Perform lab experiments or demonstrations that involve specific reactions found in the AP[®] Chemistry curriculum to familiarize students with the common types of reactions.
- Tell students to read the instructions and follow them carefully. Remind students that the answer (and only the final answer) should be written inside the reaction box. AP[®] Readers generally score only the last answer or the circled portion inside the box.
- Ask students to avoid any note-taking or scratch work (like oxidation state or ionic charge, unless they are required) inside the answer box. Ambiguous formulas like $\text{Ca}_3^{2+}(\text{PO}_4^{3-})_3$ or $\text{H}_2^+\text{CO}_3^{2-}$ are difficult to evaluate and often result in reduced credit or no credit.
- Reinforce the importance of eliminating spectator species from a net-ionic equation.
- Provide strategies for predicting which materials will ionize in aqueous solution and which will remain in molecular form. The ability to use solubility rules and identify strong acids is particularly important.
- Encourage students to inspect their chemical equations for balance of both mass and charge. The balance of net charge seems particularly easy for students to overlook.
- Students should reduce stoichiometric coefficients to the lowest whole numbers.
- Emphasize the importance of clear, legible handwriting.

Question 5

What was the intent of this question?

This question assessed students' ability to provide molecular level descriptions to explain experimental observations and examine a proposed reaction mechanism. In part (a) students applied principles of kinetic molecular theory to explain why the pressure of a gas sample increases with temperature. In part (b) students used the concept of gas stoichiometry to link the decreased number of molecules to the observed decrease in the pressure after the reaction occurs. Part (c) asked students to write a rate law that is consistent with the proposed reaction mechanism. Part (d) required students to identify the intermediates in the proposed mechanism. In parts (e) and (f), students constructed a reaction-energy profile illustrating a two-step, exothermic reaction, and labeled the E_a for the first step.

How well did students perform on this question?

The mean score was 3.70 out of 8 possible points. Students generally attempted all parts of the question. Very few responses (about 1.5 percent) omitted this question entirely. Of those who attempted to answer this question, only 4.9 percent earned a score of zero points. Students who earned only 1 or 2 points typically earned them in parts (a), (d), or (e). Students who earned 6 or 7 points typically did not earn points in parts (a) or (e).

What were common student errors or omissions?

In part (a), the entry point was accessible and was earned by a large majority of papers: students earned 1 point by citing increased molecular speed (or velocity, energy, etc.) as the cause of the increased pressure. Many responses went on to tie the increased speed of the particles to the increased frequency of collisions or the increased energy of collisions (strengthening the first point, but not yet earning the second); in these cases, students often thought they were providing two

different reasons (as in, “the molecules move faster, and also collide more frequently”). Most responses did not make a clear distinction between the two effects (frequency of collisions and force of collisions) to earn the second point.

Common misconceptions included:

- Basing increased pressure on collisions between molecules of gas, rather than collisions between molecules of gas and the container walls. Many student responses listed molecule-wall collisions and molecule-molecule collisions as two separate contributions to the increased pressure.
- Stating that as temperature increases, molecules get farther apart (expand, take up more room, etc.), leading to increased pressure on the container. The prompt specifies a rigid container, so there is no change in volume. (Many responses claimed that the molecules get farther apart, but since the volume can't change, the pressure increases.)
- Asserting that the molecules themselves get bigger, “fatter,” or heavier.
- Explaining the increase in pressure in terms of bonds breaking and the creation of more gas particles.
- Indicating that molecules “begin to collide” or “begin to move” when temperature is increased, implying that the molecules weren't moving to begin with, or that the molecules (or, worse, electrons) are “excited” by a temperature increase.
- Invoking intermolecular forces, non-ideal gas concepts, or phase changes (vapor pressure) as the cause of the increased pressure

In part (b), the reaction stoichiometry question, several unsuccessful approaches were common:

- Invoking Le Châtelier's principle (not applicable to the case, and also a bulk-level approach).
- Implying that the gas is “used up,” that there are *no* molecules colliding (or no gas molecules present) after the reaction is complete, or that only the reactants contribute to pressure. Also, indicating that only reactant molecules collide (apparently confusing reaction rate arguments with kinetic molecular theory of gases).
- Using ΔH to explain why the pressure decreases. Common examples include, “the reaction is exothermic, so it's losing heat, so the molecules slow down” or “the reaction is exothermic, so temperature decreases, so pressure decreases”. This was often, but not always, linked to attributing the pressure decrease to the product molecules being heavy, large (or small), or slow.
- Implying that the two reactant gases occupy separate volumes in the reactant container, and failing to recognize that each gas in a mixture actually occupies the entire volume of the container. For example, “the volumes of the reactant gases became smaller, therefore the pressure decreased.” This was often linked to a statement that the reaction caused the molecules to be closer together.

The point for the rate law in part (c) was earned quite often.

Common errors included:

- Writing an answer in the form of an equilibrium constant expression rather than a rate law.
- Omitting the rate constant k , or using another variable in its place (no point was earned for this error).
- Including intermediates or products in the rate law expression.
- Using variable exponents, even though the mechanism was provided.
- Supplying a generic integrated rate law or an Arrhenius expression.

In part (d), the identification of an intermediate, a large majority of students earned credit. Some students gave the formulas of the products or the reactants, or gave a reaction equation rather than identifying a single species. A few students, after explaining their choice, commented that the intermediate was not involved in the reaction.

In part (e), many students clearly had not seen a multi-step reaction-energy diagram.

Common errors included:

- Graphs drawn with the shape of titration curves, heating/cooling curves, gas-law curves, Boltzmann distributions, reaction rate or reactant concentration versus time, etc. However, a graph earned 1 point if it was drawn as exothermic. This applied to many graphs that were not, strictly speaking, classic reaction-energy diagrams.
- Curves in the reaction-energy diagram format that were drawn as a single-step reaction instead of a two-step reaction.
- Graphs in the form of a comparison of a catalyzed and an uncatalyzed reaction, which students attempted (unsuccessfully) to pass off as a two-step reaction-energy diagram.
- A handful of students drew a reasonable reaction-energy curve (single- or multi-step) and put the intermediates, C_2H_5^+ and Cl^- , at the top of the curve, in the position of the activated complex. Points are not deducted or earned for labeling the energy or the intermediates.

Among the papers with a reasonable reaction-energy curve, many identified E_a appropriately on their graphs, and earned the point for part (f).

The most common errors included:

- Identifying a single point, rather than an energy difference, as E_a .
- Writing E_a next to the curve somewhere, without clearly indicating an energy range.
- Tying the top or bottom of the E_a range to the products, the intermediates in a two-step curve, the x-axis of the graph, or nothing in particular.

Based on your experience of student responses at the AP® Reading, what message would you like to send to teachers that might help them to improve the performance of their students on the exam?

- Students need to distinguish between molecular particle-level behavior and bulk properties, and to link the two appropriately (in particular, to correctly assign cause and effect). Models or computer simulations may help students develop a molecular particle-level view of matter.
- Students would benefit from more practice generating reaction-energy diagrams for exothermic, endothermic, multi-step, and catalyzed reactions. Early and repeated exposure to reaction-energy diagrams will help students to organize their understanding of what is occurring in reactions (e.g., going up the graph, roughly speaking, corresponds to breaking or weakening bonds; bond formation relates to going back down the graph). After observing many cases of imprecise labeling of E_a , dashed lines, clearly assigning the E_a range to the top of the curve and the energy of the reactants, would have helped. Many students were able to appropriately label a single-step reaction-energy curve, but had no idea what to do with a multi-step process.
- In learning about reaction mechanisms, students should be able to:
 - Write the rate law for each step

- Identify reactants, products, intermediates, and catalysts
- Consider what an intermediate might look like, how it is different from an activated complex or catalyst, why it is not stable, and why it doesn't appear in the overall reaction
- Correlate the mechanism to the overall reaction
- Generate an appropriately labeled reasonable reaction-energy curve
- The temperature of a gas sample (bulk property) is a measure of the average kinetic energy of the molecules (molecular-level property). If a reaction is carried out under isothermal conditions, the average kinetic energy of the sample does not change, regardless of the ΔH of the process.
- Reactions involving ions may not necessarily take place in aqueous solution. Direct or indirect reference to solution conditions appeared in parts (b), (c), and (d), despite the given phase labels.
- Gas-phase reactions are generally presented as happening in containers. Students need some exposure to the physics of gases: What is pressure? What is temperature? What is the energy of a collision? Additionally, students should know that gases do not separate by type inside a container; in a mixture of gases, each component moves freely within the entire volume of the container.
- AP[®] students only rarely encounter gases that do not behave ideally, and should assume ideal behavior unless given explicit evidence or direction to the contrary.
- Le Châtelier arguments can only be used in the context of an existing equilibrium with an applied stress. Le Châtelier arguments (and those involving $PV = nRT$) are generally bulk-level observations, and are an effect of the molecular-level kinetics of the reaction, not a cause or explanation of the molecular-level behavior.

Question 6

What was the intent of this question?

This question provided students with data to draw Lewis electron-dot diagrams and to use the concept of intermolecular forces to explain observations about solubility and vapor pressure. Part (a)(i) asked students to complete a Lewis electron-dot diagram, including any lone (non-bonding) pairs of electrons when given an array of atoms arranged to represent ethanol. Part (a)(ii) assessed knowledge of structural isomers by asking students to draw a complete Lewis electron-dot diagram for the isomer of the compound drawn in part (a)(i). In part (b) students used the information in the data table to determine which of the two compounds was represented by the Lewis electron-dot structure in part (a)(i), justifying answers in terms of the intermolecular forces present in each compound. Part (c) asked students to examine the Lewis electron-dot diagrams of dichloromethane and carbon tetrachloride and to use intermolecular forces to explain why dichloromethane is more soluble than carbon tetrachloride. In part (d) students explained why dichloromethane has a higher vapor pressure than carbon tetrachloride in terms of intermolecular forces. Part (e) asked students to draw a correctly oriented water molecule to illustrate a hydrogen bond formed between a molecule of water and a molecule of methanal.

How well did students perform on this question?

The mean score was 2.85 out of a possible 9 points. Students generally attempted all parts of this question. The most commonly earned points were in part (a) for the Lewis structures and, to a lesser extent, part (c) for a simplistic explanation of the relative solubilities of dichloromethane and carbon tetrachloride using their polarities. The least often earned point was in part (d) for an

explanation of how the intermolecular forces of dichloromethane and carbon tetrachloride affected their vapor pressures.

What were common student errors or omissions?

In part (a) students used a variety of unacceptable representations for both bonding and lone pairs of electrons. Some used a combination of lines and dots to represent the same pair of electrons.

Common errors included:

- Students did not understand the concept of an isomer. A significant number simply redrew the ethanol as its mirror image or with a ninety-degree rotation of the hydroxyl group.
- Diagrams that included five bonds on a central carbon atom.

In part (b) the majority of students were able to determine that the alcohol would have a higher boiling point and a greater solubility in water than the ether. Of those who successfully identified the alcohol many mentioned hydrogen bonds in their justification for choosing it; however, it was apparent that few had a good understanding of what a hydrogen bond actually is.

Common errors and misconceptions included:

- Students referring to the covalent bond between the oxygen and the hydrogen in the alcohol's structure as the hydrogen bond.
- Students did not recognize that the ether is a bent structure and is a polar molecule. This misconception led many students not to list all of the intermolecular forces involved.
- Students did not recognize that methoxymethane (dimethyl ether) can form hydrogen bonds with water.

In part (c) the most common misconception was that the hydrogen atoms in dichloromethane's structure were responsible for its solubility.

Common errors included:

- A lack of understanding of what the term "dissolving" means. The use of terms like "reacts with"; "bonds with"; "dissociates in"; "breaks up in" to describe what happens when dichloromethane dissolves in water demonstrated this deficiency.
- Many students simply stated the adage "like dissolves like" accompanied by the relative polarities of each of the three species. A full and complete explanation was required to earn points with this approach.

Part (d) provided a final chance for students to focus on hydrogen bonding. Usually, the misconceptions were similar to those demonstrated in part (c). Also, as in part (c), many did not discuss intermolecular forces at all, but defaulted to a simple comparison of polarities.

Common errors and misconceptions included:

- Misunderstanding the concepts of vapor pressure and hydrogen bonding. Students who correctly identified the intermolecular forces often incorrectly associated them with vapor pressure stating that strong intermolecular forces lead to a high vapor pressure.
- Very few students actually noted that carbon tetrachloride's relatively large London dispersion forces were actually greater than the dipole-dipole and London forces of dichloromethane, assuming instead that the dipole-dipole forces were greater and that this was responsible for the large vapor pressure.
- Many students seemed to confuse vapor pressure with heat of vaporization, indicating that

particles with large intermolecular forces require more energy to vaporize.

- Students associated vapor pressure with the “speed” or “root mean square velocity” of the particle.
- Students may have performed an experiment where they collected a gas over water because they sometimes cited the need to consider the intermolecular forces between the two liquids and water occasionally referring to the species as being “in solution.”

In part (e) about 50 percent of the students conveyed their misunderstanding of hydrogen bonding visually.

Common errors included:

- Showing a bond between the oxygen atom in the water molecule with a hydrogen atom in the methanol.
- Students not reading the entire question or ignoring the instructions as they didn't show a “dashed line” in their illustration to represent the hydrogen bond.

In general, students continued to confuse *intermolecular* and *intramolecular* forces. While intermolecular forces may indeed be considered weak bonds of a sort, many students are not communicating a clear distinction between these types of forces and strong covalent or ionic bonds. In fact, many did a good job of describing the latter when they likely meant the former. Often students were unable to demonstrate a mastery of chemical vocabulary.

Based on your experience of student responses at the AP[®] Reading, what message would you like to send to teachers that might help them to improve the performance of their students on the exam?

- Remind students to read the entire question and answer what is being asked. They should be taught to underline key phrases and box or outline exactly what they are being asked to do.
- Remind students to communicate clearly and precisely. The value of concise answers that use precise terminology cannot be overemphasized. It is important to provide frequent opportunities for students to practice writing skills.
- When asked to provide a justification for an answer, it is not adequate to simply provide a statement of fact or to repeat the information given as justification. Students should give an explanation, sometimes with evidence, as to why something occurs.

Suggestions for teachers:

- Insist that students write legibly and place all answers in the lined spaces or boxes provided instead of squeezing words in between the question parts.
- Students should be diligent in reading through the entire question in order to avoid losing points, such as in part (e) in this question.
- Insist that when comparing two substances, students always refer to both substances and clearly indicate which one they are referring to in turn (avoid ambiguities like “**It** dissolves in water”). Discussing one substance will not earn credit.
- Do not accept incorrect terminology from students. Practice defining and applying chemical terms in a variety of contexts.

- Encourage students to explain and justify answers for all types of questions and to support claims with evidence. Explanation and justification are extremely important assessment tools.
- Get students to write (not type) legible explanations of phenomena, discrepant and non-discrepant. Provide students with tabular data and expect the same.
- Intramolecular and intermolecular interactions are difficult for students to distinguish and describe. You may assist students by:
 - Teaching the major bonding forces (ionic, covalent, metallic) separately from the intermolecular forces.
 - Once all force types have been mastered, have students consider both force types and put them into categories.
 - Have students illustrate attractions within and between molecules and clearly indicate which attractions are intramolecular forces and which are intermolecular forces
 - Have students practice explaining differences in data using the various types of interactions.
- Avoid using the term “Van der Waals forces” as synonymous to “London dispersion forces.” The term “Van der Waals forces” is more correctly used as a descriptor to encompass all weak intermolecular interactions.
- Avoid using molar mass as a predictive tool for the strength of London dispersion forces. While mass often works, it is not always successful (a good counter example is neon versus methane). A better argument would be to focus on the cause and effect of the size of an electron cloud on the polarizability of a molecule and hence the strength of London dispersion forces.
- Use molecular three-dimensional software and/or simulations to illustrate molecular representations, rather than just two-dimensional drawings
- When drawing Lewis electron-dot diagrams, have students:
 - Determine the total number of electrons in the structure first.
 - Use the open areas of the page to refine the structure.
 - Draw dots large enough to be obvious and try to avoid erasures.
 - Clearly indicate only the final structure in the box.