AP® Chemistry Mini-Test
Answer Key and Explanations

with Tyler DeWitt

For in-depth explanations of the answers and videos for each question, check out: www.tdwscience.com/apchem-resources

1) **C**
Given equal masses, there are more moles of sulfur present than moles of iron, since iron has a larger molar mass. Because the stoichiometry of the reaction requires one mole of iron to react with one mole of sulfur, iron is the limiting reagent, and is completely consumed in the reaction. This leaves only excess sulfur and product, FeS, in the reaction vessel.

2) **D**
The species that make up a Brønsted-Lowry acid-base conjugate pair differ by one proton. Only $H_2PO_4^-$ and HPO_4^{2-} meet these criteria. Choice (C) shows a pair that differs by one proton, but HNO_3 is a strong acid, so it cannot form a conjugate base.

3) **B**
Reading the peaks from right to left shows an element with an electron configuration of $1s^2\,2s^2\,2p^6\,3s^2\,3p^1$. This is the configuration of aluminum, with 13 electrons.

4) **A**
Ammonium phosphate dissociates fully according to the equation $(NH_4)_3PO_4 \longrightarrow 3NH_4^+ + PO_4^{3-}$. This shows a 1:3 ratio of ammonium phosphate to ammonium. If the concentration of $[NH_4^+] = 1.50$ M, the concentration of ammonium phosphate must be one-third this amount, or 0.50 M.

5) **A**
At the equivalence point of an acid-base titration, moles of H^+ must equal moles of OH^-. If it requires 20.0 mL of 0.100 HCl to neutralize the $Ca(OH)_2$, there must be:

mol = M•L = 0.100 M • 0.200 L = 0.00200 moles of H^+ ions.

Because 1 mole of $Ca(OH)_2$ dissociates into 2 moles of OH^- ions, there are only 0.00100 moles of $Ca(OH)_2$ present. Multiplying that by its molar mass gives 0.0741 grams.

6) **C**
For an oxyacid, the more electronegative oxygens attached to the central atom, the more powerful the pull on the O–H bond. This makes the bond *weaker*, which makes the acid *stronger*—the acid experiences a higher perfect ionization because it is easier to remove the hydrogen.

7) **C**
A reaction is automatically spontaneous if it is exothermic *and* increases entropy. Here, only choice (C) meets both requirements.

8) **B**
If the distilled water is left in the buret after cleaning, it will dilute the 0.001 M solution of $KMnO_4$ added to the buret. Therefore, the buret should be rinsed with potassium permanganate solution to clear out the water before starting the experiment.

9) **D**
The method of initial rates must be applied to each reactant in the table. Between Experiment 1 and Experiment 2, doubling [Q] doubles the rate, so the reaction is first order with respect to Q. Between Experiment 1 and Experiment 4, doubling [X] has no effect on the rate, so the reaction is zero order with respect to X. Between Experiment 1 and Experiment 3, doubling [M] quadruples the rate, so the reaction is second order with respect to M. This gives a rate law of Rate = $k[M][Q]^2$.

10) **D**
LeChâtelier's Principle states that a reaction at equilibrium will shift to counteract a change and return to equilibrium. The only change that would shift this exothermic reaction left is increasing the temperature. This is because heat is treated like a reactant or a product; so, for an exothermic reaction, raising the temperature is equivalent to adding more product. The reaction shifts left to counteract the disturbance.

11) **D**
Nitric acid is a strong acid and undergoes complete dissociation according to this equation: $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$. Only choice (D) shows fully dissociated hydrogen and nitrate ions.

12) **A**
The equation that "best represents the reacting species" is a *net ionic equation*. Here, that is:

$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
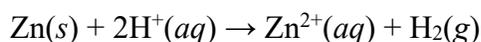
Choice (C) represents the full single replacement reaction, and is *not* a net ionic equation.

13) **A**
This is an example of the common ion effect applied to an "insoluble" salt at equilibrium. The presence of KCl in Beaker B supplies additional Cl^- ions, which act to depress the solubility of AgCl. Because K_{sp} must remain constant in both beakers, Beaker A has a greater concentration of silver ions.

14) **B**
Carbon dioxide has two electronegative oxygens bonded to a central carbon. These bonds are polar, but the linear shape of the molecule cancels them out.

15) **B**
Atomic radius *increases* moving down a period. Ionization energy and electronegativity *decrease* moving down a period. Only choice (B) meets these criteria.

16) B

Dalton's Law of Partial Pressures states that the total pressure of a system is equal to the sum of the partial pressures. Because the system maintains a constant temperature and volume, the Ideal Gas Law indicates that each partial pressure is directly related to the number of moles of gas. Given equal masses of both gases, there is twice as much neon present. (Neon's ~20 grams/mol compared to argon's ~40 grams/mol.) If the total pressure of 12 atm, the partial pressure of neon must be 8 atm, and the partial pressure of argon is 4 atm.

17) B

This electron configuration depicts an atom with 36 electrons. Only the strontium ion, Sr^{2+}, meets this condition.

18) B

This is a specific heat/calorimetry problem in which magnesium is heated and transfers its heat to water. To calculate the mass of magnesium, the heat lost by the magnesium must first be established. This is done by assuming the heat *lost* by the magnesium is equal to the heat *gained* by the water:

For water: $q = mc\Delta T = (100.0 \text{ g})(4.2 \text{ J/g°C})(1.0°C) = 420 \text{ J}$

This value is then substituted in the equation for the mass of the magnesium:

For magnesium: $m = \dfrac{q}{c\Delta T} = \dfrac{420 \text{ J}}{(1.0\frac{J}{g}°C)(60.0°C)} = 7.0 \text{ grams Mg}$

19) A

The standard cell potential is calculated by $E°_{cell} = E°_{cathode} - E°_{anode}$. In this electrochemical reaction, Fe^{3+} ions are being reduced to Fe^{2+} ions at the cathode, and Pb atoms are being oxidized to Pb^{2+} ions at the anode. Plugging in these values results in $E°_{cell} = 0.77 - (-0.13)$. This is mathematically identical to choice (A).

20) A

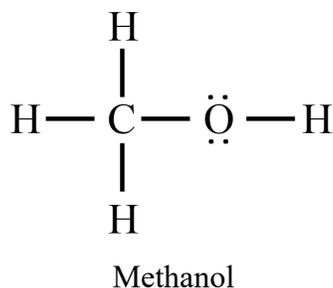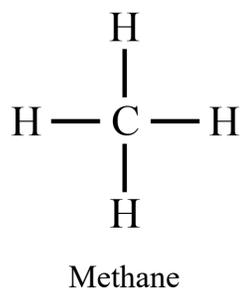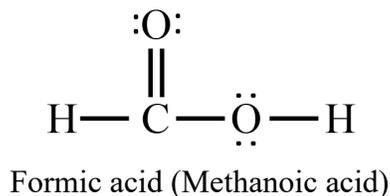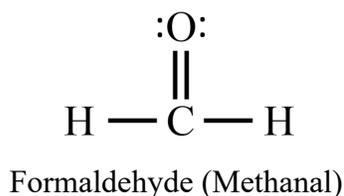Electrode 1 is a lead anode. As lead atoms are oxidized to Pb^{2+} ions, the mass decreases. Electrode 2 is a platinum cathode. As Fe^{3+} ions are reduced to Fe^{2+} ions, there's no change to the inert platinum electrode.

Question 1

This question refers to the following compounds.

$$:O:$$
$$\|$$
$$H-C-H$$

Formaldehyde (Methanal)

$$:O:$$
$$\|$$
$$H-C-\ddot{O}-H$$

Formic acid (Methanoic acid)

$$H$$
$$|$$
$$H-C-H$$
$$|$$
$$H$$

Methane

$$H$$
$$|$$
$$H-C-\ddot{O}-H$$
$$|$$
$$H$$

Methanol

(a) Indicate the hybridization of the carbon atom in formaldehyde. (1 point)

The carbon atom in formaldehyde is *sp²* hybridized because it has two single bonds and one double bond.	1 point is earned for a correct explanation

(b) Indicate the approximate bond angle of the C–O–H bond in methanol.

Because the geometry around the oxygen atom is roughly *sp3* tetrahedral, the C–O–H bond angle is approximately 109.5°.	1 point is earned for a correct explanation

(c) The boiling points of the four compounds are shown in the table below. Use concepts of intermolecular forces to explain why methanol has a higher boiling point than formaldehyde. (1 point)

Compound Name	Boiling Point (1 atm)
Formic Acid	100.8 ºC
Methanol	64.7 ºC
Formaldehyde	−19.3 ºC
Methane	−162.2 ºC

Methanol has a higher boiling point than formaldehyde because the hydrogen attached to methanol's electronegative oxygen is prone to hydrogen bonding with the electron pairs on adjacent methanol molecules. This is a very strong intermolecular attraction, requiring much energy to break. By contrast, formaldehyde only has a carbon-oxygen bond, which can only participate in dipole-dipole attractions, which are weaker than hydrogen bonds.	1 point is earned for a correct explanation

(d) A student argues that methane molecules experience no intermolecular forces. Explain why the student is or is not correct. (1 point)

The student is incorrect. All particles experience London dispersion forces, which arise from the random motions of shared electron pairs.	1 point is earned for a correct explanation

(e) Methanol can be used as a fuel for a variety of applications. The balanced equation for its combustion is given below.

$$2CH_3OH(g) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$$
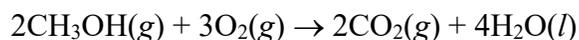
Calculate the $\Delta H°$ for the reaction above, using the data from the table below. (2 points)

Compound	$\Delta H_f°$ (kJ/mol)
$CH_3OH(g)$	–201.0
$O_2(g)$	0.0
$CO_2(g)$	–393.5
$H_2O(l)$	–285.8

$$\Delta H° = \sum \Delta H_f°_{products} - \sum \Delta H_f°_{reactants}$$ $$\Delta H° = \left(2\left(\Delta H_f°_{CO_2}\right) + 4\left(\Delta H_f°_{H_2O}\right)\right) - \left(2\left(\Delta H_f°_{CH_3OH}\right) \right.$$ $$\left. + 3\left(\Delta H_f°_{O_2}\right)\right)$$ $$\Delta H° = \left(2(-393.5 \text{ kJ}) + 4(-285.8 \text{ kJ})\right)$$ $$- \left(2(-201.0 \text{ kJ}) + 3(0.0 \text{ kJ})\right)$$ $$\Delta H° = (-1930.2 \text{ kJ}) - (-402.0 \text{ kJ})$$ $$\Delta H° = -1528.2 \text{ kJ}$$	1 point for subtracting the reactants from the products. 1 point for the correct calculation

(f) Methanol has a density of 0.792 g/mL and a molar mass of 32.04 g/mol. How much heat will be released when a 150.0 mL sample of methanol is combusted with an excess of oxygen? (2 points)

$$mass_{CH_3OH} = density \times volume$$ $$mass_{CH_3OH} = 0.792 \frac{g}{mL} \times 150.0 \text{ mL}$$ $$mass_{CH_3OH} = 119 \text{ g}$$ $$mol_{CH_3OH} = \frac{119 \text{ g}}{32.04 \frac{g}{mol}}$$ $$mol_{CH_3OH} = 3.71 \text{ mol}$$	1 point for correctly calculating the mass of methanol

$\Delta H° = -1528.2 \text{ kJ/rxn}$ 2 moles methanol per mole of reaction $\Delta H° = 3.71 \text{ mol}_{CH_3OH} \times \dfrac{1 \text{ rxn}}{2 \text{ mol}_{CH_3OH}} \times \dfrac{-1528.2 \text{ kJ}}{1 \text{ rxn}} = -2830 \text{ kJ}$	1 point for correctly calculating the amount of heat released

Question 2

Formic acid (CHOOH) reacts with water according to the following equation:

$$CHOOH + H_2O \leftrightharpoons H_3O^+ + CHOO^-$$
$$K_a = 1.78 \times 10^{-4} \text{ at } 25°C$$

(a) Write the equilibrium constant expression, K_a, for formic acid. (1 point)

$K_a = \dfrac{[H^+][COO^-]}{[COOH]}$	1 point is earned for a correct equilibrium expression

(b) Calculate the pH of a 0.150 M aqueous solution of formic acid. (2 points)

	[COOH]	[H⁺]	[COO⁻]	
				1 point is earned for calculating x
Initial	0.15	0	0	
Change	$-x$	$+x$	$+x$	
Equilibrium	$0.15-x$	$+x$	$+x$	

$K_a = \dfrac{[H^+][COO^-]}{[COOH]} = \dfrac{x^2}{0.15-x} \approx \dfrac{x^2}{0.15}$

(Assuming x is small relative to 0.15)

$x^2 = K_a \times 0.15$

$x = \sqrt{K_a \times 0.15} = \sqrt{(1.78 \times 10^{-4}) \times 0.15} = 0.0052$

$[H^+] = x = 0.0052$

$pH = -\log[H^+] = -\log(0.0052) = 2.3$

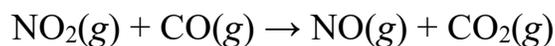1 point is earned for a correct calculation of pH

(c) To create a buffer solution, sodium formate (NaCHOO) can be added to formic acid.

How many moles of sodium formate would need to be added to 250 mL of a 0.10 M formic acid solution to create a buffer with a pH of 3.90? (2 points)

$pH = pK_a + \log\frac{[A^-]}{[HA]} = pK_a + \log\frac{[CHOO^-]}{[CHOOH]}$ $\log\frac{[CHOO^-]}{[CHOOH]} = pH - pK_a$ $pK_a = -\log(1.78 \times 10^{-4}) = 3.75$ $\log\frac{[CHOO^-]}{[0.10]} = 3.90 - 3.75 = 0.15$ $\frac{[CHOO^-]}{[0.10]} = 10^{0.15}$ $[CHOO^-] = 0.10\ M \times 10^{0.15}$ Moles $CHOO^- = 0.14\ M \times 0.250\ L = 0.035$ mol	1 point is earned for a correct calculation of $[CHOO^-]$ 1 point is earned for a correct calculation of moles

Question 3

$$NO_2(g) + CO(g) \rightarrow NO(g) + CO_2(g)$$

(a) Explain how the following graphs suggest the reaction above is second order with respect to NO_2.
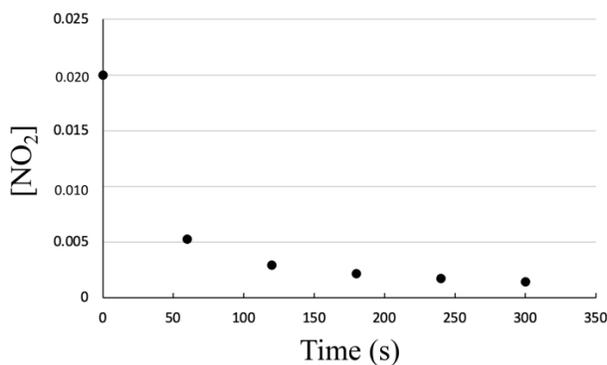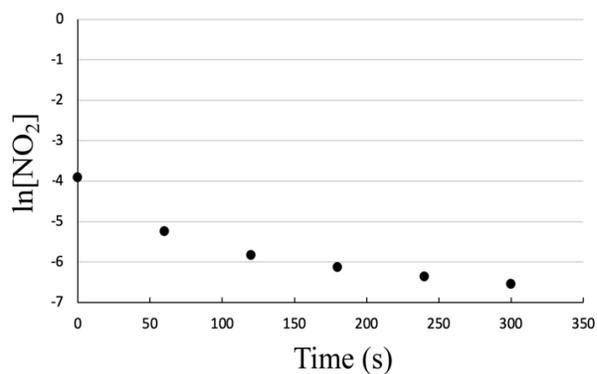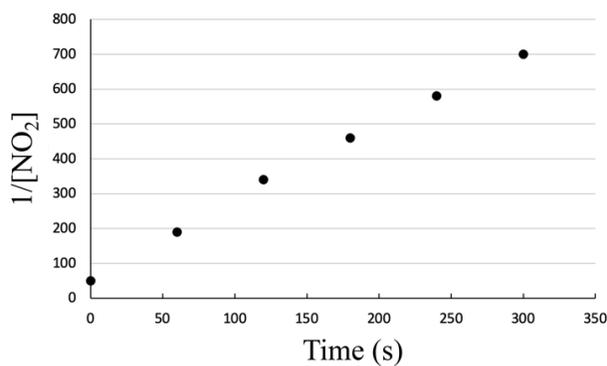
Because the plot of time vs. $1/[NO_2]$ is a linear fit, the reaction is second order with respect to NO_2.	1 point is earned for a correct explanation

(b) Write a rate law for the full reaction, assuming it is zero-order with respect to CO. (1 point)

Rate = $k\,[NO_2]^2$	1 point is earned for the correct rate law

(c) The initial rate of reaction is 1.4×10^{-4} mol L^{-1} s^{-1}. Using the graphs above, determine the rate constant, k. (1 point)

Rate $= k \, [NO_2]^2$ $k = \dfrac{Rate}{[NO_2]^2}$ From graph: At 0 sec, $[NO_2] = 0.02$ M $k = \dfrac{1.4 \times 10^{-4} \; mol \; L^{-1} \, s^{-1}}{(0.02 \; M)^2} = 0.35$ L mol^{-1} s^{-1}	1 point is earned for the correct calculation of k

(d) The reaction is expected to proceed according to the mechanistic steps below:

$$NO_2 + NO_2 \rightarrow NO_3 + NO \text{ (slow)}$$

$$NO_3 + CO \rightarrow NO_2 + CO_2 \text{ (fast)}$$

Explain whether this mechanism is consistent with your proposed rate law from part 2. (1 point)

The mechanism is consistent with a rate law of Rate = k[NO_2]^2. The slow step is the rate-determining step, and it only has NO$_2$. Also, this step is bimolecular with respect to NO$_2$, which explains why it's second order.	1 point is earned for a correct explanation

(e) The reaction proceeds spontaneously. Sketch a relative potential energy diagram for the overall reaction. (1 point)
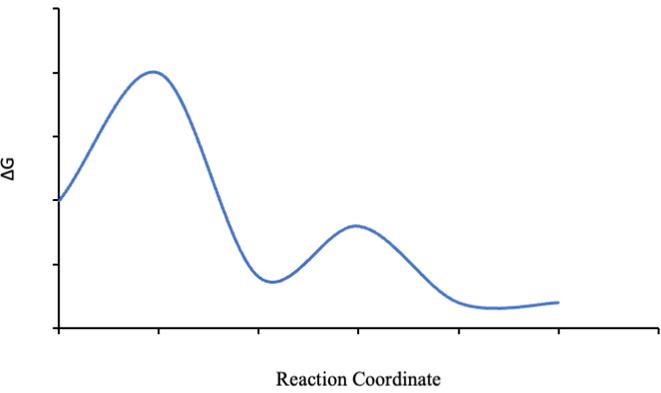
	1 point is earned for a correct potential energy diagram