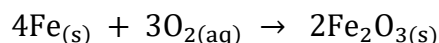


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Individual Exam Solutions

1. The process of the oxidation of iron is commonly known as rusting. It begins electromagnetically as electrons are transferred from iron to oxygen to produce iron (III) oxide. This reaction is given below:



- a) Assume that the reaction takes place as one elementary step in a beaker of water (this is unrealistic, but the real reaction is much more complicated), determine the rate law for the reaction using the data below. Remember to consider units.

Experiments	[Fe]	[O ₂]	Rate of reaction
1	5.0 g	$1.38 \times 10^{-4} \text{ M}$	$5.110 \times 10^{-6} \text{ M/s}$
2	6.0 g	$1.75 \times 10^{-4} \text{ M}$	$1.042 \times 10^{-5} \text{ M/s}$

Either of the rows can be used to determine the rate law. Assuming that the reaction takes place according to the elementary step $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$, the exponents on the reactant concentrations are simply the coefficients. However, iron does not factor into the rate law equation because it is in a solid state.

The rate law can be written as $\text{rate} = k[\text{O}_2]^3$.

Using the data in the table,

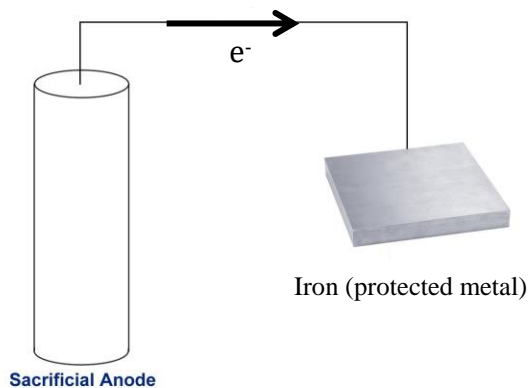
$$5.110 \times 10^{-6} \text{ M/s} = k(1.38 \times 10^{-4} \text{ M})^3 = k(2.63 \times 10^{-12} \text{ M}^3).$$

$$\text{Therefore, } k = \frac{5.110 \times 10^{-6} \text{ M/s}}{2.63 \times 10^{-12} \text{ M}^3} = \mathbf{1.94 \times 10^6 / \text{M}^2 \cdot \text{s}}$$

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- b) To prevent the rusting of iron, a sacrificial anode is a metal that is often used that oxidizes in the place of the iron.



The table below lists **oxidation potentials** of various metals. Based on the information, give two examples of metals that might be used as a sacrificial anode to prevent the rusting of iron? What makes these metals good sacrificial anodes?

Metal	Potential, Volts	Metal	Potential, Volts
Calcium	+2.20	Hydrogen	0.000
Magnesium	+1.87	Antimony	-0.190
Aluminum	+1.30	Arsenic	-0.320
Manganese	+1.07	Bismuth	-0.330
Zinc	+0.758	Copper	-0.345
Chromium	+0.600	Mercury	-0.799
Iron	+0.441	Silver	-0.800
Cadmium	+0.398	Platinum	-0.863
Nickel	+0.220	Gold	-1.100

Any metal with an oxidation potential higher than that of iron's could be used as a sacrificial anode for iron so, Calcium, magnesium, aluminum, manganese, zinc or chromium.

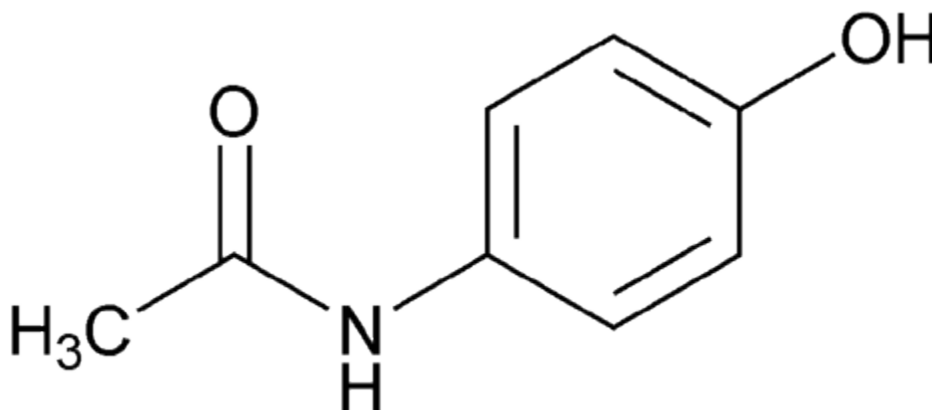
These metals are good sacrificial anodes because a higher potential means that electrons are likely to flow from that metal to the iron that was oxidized, preventing the rusting of the iron.

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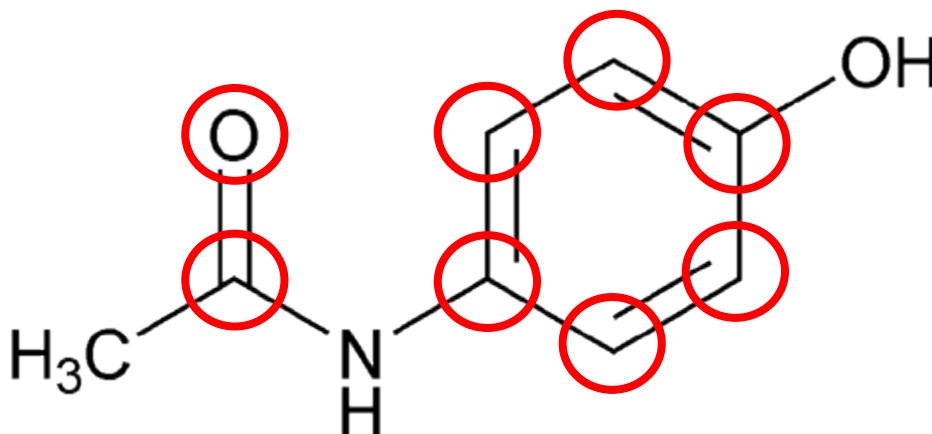
2. Acetaminophen is a medication used as a pain reliever and fever reducer, found in common brands like Tylenol and Panadol. As with many medications, the right dosage can be effective but too much may lead to symptoms ranging from nausea to the induction of a coma (The latter can be avoided by at least remaining in the ballpark of the recommended dosage).

- a. Below is the structure of an acetaminophen molecule. Write the number of atoms in the molecule that have AT LEAST one lone pair of valence electrons.



3 atoms (both of the oxygen atoms and the nitrogen atom) have at least one lone pair of valence electrons. The oxygen atoms each have 2 lone pairs while the nitrogen atom has one lone pair.

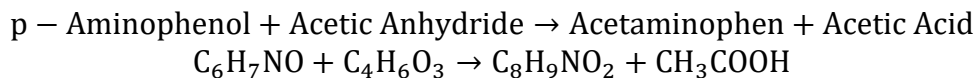
- b. Circle all the sp^2 hybridized atoms and draw a box around all the sp hybridized atoms in the below structure of acetaminophen.



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- c. Acetaminophen can be produced in a reaction between p-Aminophenol (C_6H_7NO) and acetic anhydride ($C_4H_6O_3$). Acetic acid (MW = $60.05 \frac{g}{mol}$; $K_a = 1.74 \times 10^{-5}$) is a by-product of the reaction. Acetic acid. The reaction is written below using common names and molecular formulas.



You have isolated some acetaminophen using the correct molar ratios of reactants, and to check the yield, you analyze the left-over acetic acid. You add enough water to obtain a 500 mL solution of acetic acid, and add 2.34 g solid NaOH (assume no volume change). The pH after the NaOH dissolves is measured to be 4.76. Calculate the original mass of acetic acid generated. Hint: a buffer system must be considered.

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right) \Rightarrow 4.76 = 4.76 + \log\left(\frac{[A^-]}{[HA]}\right)$$

$$\frac{[A^-]}{[HA]} = 1$$

$$(2.34 \text{ g NaOH}) \cdot \frac{1 \text{ mol NaOH}}{40 \text{ g NaOH}} \cdot \frac{1 \text{ mol } CH_3COO^-}{1 \text{ mol NaOH}} = 0.0585 \text{ mol } CH_3COO^-$$

$$0.117 \text{ mol } CH_3COOH \text{ originally} \cdot \frac{60.05 \text{ g } CH_3COOH}{1 \text{ mol } CH_3COOH} = 7.026 \text{ g acetic acid}$$

- d. In a separate instance, 7.18 g of acetic acid was formed in a reaction between p-Aminophenol and acetic anhydride. Assume the reaction went to completion. The acetaminophen (MW = 151.16 g/mol) formed from this same reaction was used to make 9 bottles of Tylenol. How many grams of acetaminophen were needed for each bottle of Tylenol, assuming an equal amount went into each bottle?

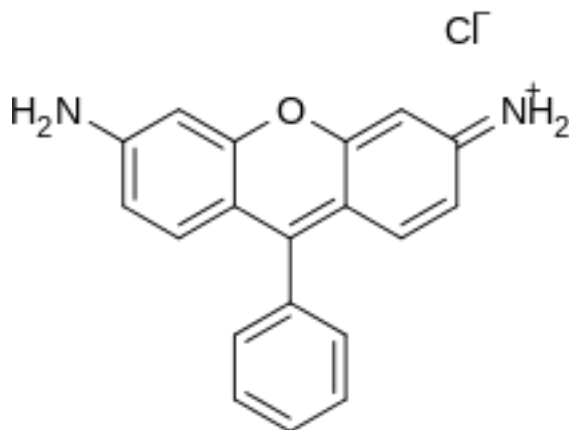
$$7.18 \text{ g acetic acid} \cdot \frac{1 \text{ mol acetic acid}}{60.05 \text{ g acetic acid}} = 0.1196 \text{ mol acetic acid}$$

$$(0.1196 \text{ mol acetic acid}) \cdot \frac{1 \text{ mol acetaminophen}}{1 \text{ mol acetic acid}} \cdot \frac{151.16 \text{ g acetaminophen}}{1 \text{ mol acetaminophen}} = 18.079 \text{ g acetaminophen} = 2.009 \text{ g per bottle}$$

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3. Rhodamine (322 g/mol) is a dye commonly used in cancer research due to its fluorescent properties. Its chemical structure is shown below.



- a. Identify the regions of rhodamine that are hydrophilic and the regions that are hydrophobic. Which regions lead to rhodamine's high solubility in water?

The regions that are hydrophilic are the amine groups and the oxygen atom. The hydrophobic regions are the aromatic rings. The parts that lead to the high solubility are the hydrophilic groups

- b. If a scientist has 1.5 grams of Rhodamine, how many milliliters of water would he need to dissolve the Rhodamine to make a 0.01M solution?

$$\frac{1.5 \text{ grams of rhodamine}}{322 \text{ grams per mol of Rhodamine}} = 0.0047 \text{ moles of Rhodamine}$$

$$\frac{0.0047 \text{ moles of Rhodamine}}{X \text{ mL of Water}} = 0.01 \text{ moles per liter}$$

470 mL of water

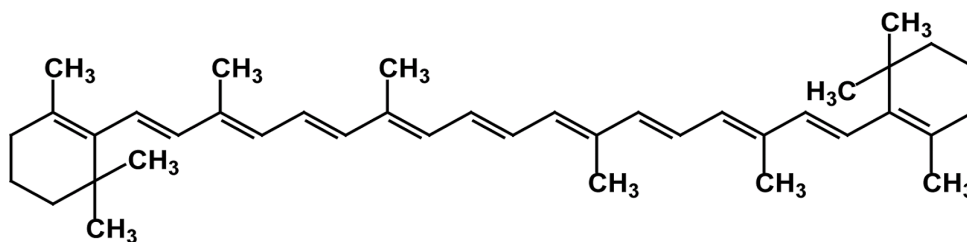
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- c. Rhodamine has two forms, Rhodamine-6G and Rhodamine-B. Rhodamine-6G has a green color and emits a wavelength of 576 nm and Rhodamine-B emits purple-red light with a wavelength of 590 nm. Which form emits the highest energy photons? Justify your answer.

The Rhodamine 6G will have higher energy photons, because the wavelength of the photons are lower, meaning that the photons emitted from the rhodamine-6G are higher in energy.

- d. Beta carotene is another pigment that has a strong orange color. The structure is listed below. What structural component of both compounds makes them strongly colored?

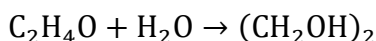


The large number of double bonds is responsible for the coloration of these compounds

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4. Colligative properties, with respect to solutions, are properties that depend on the concentration of solute dissolved. Freezing point depression and boiling point elevation are two such properties and are commonly applied in everyday life. For instance, internal combustion engines in automobiles use a solution of ethylene glycol $(\text{CH}_2\text{OH})_2$ and water in order to lower the freezing point of the water coolant for cold weather and raise the boiling point of the coolant for hot weather. Ethylene glycol is prepared by reacting ethylene oxide $(\text{C}_2\text{H}_4\text{O})$ with water in the following chemical equation:



- a. Suppose that you react ethylene oxide with water to obtain a certain mass of ethylene glycol. When all of the ethylene glycol obtained is dissolved in 1.00 kg of pure water, the boiling point is raised by 2.017°C . What mass of ethylene oxide did you start with? Assume that the reaction and dissolution go to completion, and remember that $K_{b,\text{H}_2\text{O}} = 0.512 \frac{^\circ\text{C}}{\text{molal}}$.

$$\begin{aligned}\Delta T_b &= K_b * m \\ \Delta T_b = 2.017^\circ\text{C} &= [0.512^\circ\text{C kg H}_2\text{O mol}^{-1} \text{CH}_2(\text{OH})_2] * m \\ m &= 2.017 / [0.512^\circ\text{C kg H}_2\text{O mol}^{-1} \text{CH}_2(\text{OH})_2] \\ &= 3.94 \text{ mol CH}_2(\text{OH})_2 \text{ kg}^{-1} \text{H}_2\text{O}\end{aligned}$$

$$\begin{aligned}\text{Moles of CH}_2(\text{OH})_2 \text{ in solution} &= 3.94 \text{ mol CH}_2(\text{OH})_2 \text{ kg}^{-1} \text{H}_2\text{O} * 1 \text{ kg H}_2\text{O} \\ &= 3.94 \text{ mol CH}_2(\text{OH})_2\end{aligned}$$

$$\begin{aligned}\text{Moles of C}_2\text{H}_4\text{O required} &= 3.94 \text{ mol CH}_2(\text{OH})_2 * (1 \text{ mol C}_2\text{H}_4\text{O} / 1 \text{ mol CH}_2(\text{OH})_2) \\ &= 3.94 \text{ mol C}_2\text{H}_4\text{O}\end{aligned}$$

$$\begin{aligned}\text{Mass of C}_2\text{H}_4\text{O required} &= 3.94 \text{ mol C}_2\text{H}_4\text{O} * 44 \text{ g C}_2\text{H}_4\text{O} / \text{mol C}_2\text{H}_4\text{O} \\ &= 173.36 \text{ g C}_2\text{H}_4\text{O}\end{aligned}$$

- b. Realistically, the preparation of ethylene glycol has a 90% yield by mass. Given that an antifreeze mixture is usually 50% ethylene glycol, how much ethylene oxide should be added to 1 kg of water to produce the antifreeze mixture? Assume the mass of ethylene oxide added does not change the volume of aqueous solution.

Assume that you have x moles of $\text{C}_2\text{H}_4\text{O}$ added to 1 kg (55.56 mol) water. Upon reacting, there will be $0.1x$ mol $\text{C}_2\text{H}_4\text{O}$, $55.56 - 0.9x$ moles of water, and $0.9x$ moles of ethylene glycol. Since ethylene glycol should be 50% by mass, then:

$$\begin{aligned}(0.1x \text{ mol C}_2\text{H}_4\text{O}) \left(44 \frac{\text{g}}{\text{mol}}\right) + (55.56 - 0.9x \text{ mol water}) \left(18 \frac{\text{g}}{\text{mol}}\right) \\ = (0.9x \text{ mol (CH}_2\text{OH)}_2) \left(62 \frac{\text{g}}{\text{mol}}\right)\end{aligned}$$

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b cont.) Solving for x, you find that $x=14.8$, meaning that 14.8 mol of ethylene oxide were added initially, or 651.2 g

- c. You have 2 kg of a solution of ethylene glycol that is 50% by mass. How much heat (in kJ) is required to bring the antifreeze mixture from room temperature (21°C) to its boiling point? The specific heat of the solution is $3.62 \frac{\text{J}}{\text{g}\cdot^{\circ}\text{C}}$, $K_{b,\text{H}_2\text{O}} = 0.512 \frac{^{\circ}\text{C}}{\text{molal}}$, and assume ideal behavior of solutions.
Hint: You will need to calculate the molality to find the boiling point.

$$1\text{kg CH}_2(\text{OH})_2 * (1000\text{g} / 1\text{kg}) * (1 \text{ mol CH}_2(\text{OH})_2 / 62.7\text{g}) = 15.95 \text{ mol CH}_2(\text{OH})_2$$

$$m = \text{molality of solutes in solution}$$

$$= (15.95 \text{ mol (CH}_2(\text{OH}))_2) / 1\text{kg H}_2\text{O}$$

$$= 15.95 \text{ molal}$$

$$\Delta T_b = K_b * m = 0.512 \text{ }^{\circ}\text{C kg H}_2\text{O mol}^{-1} \text{ solute} * 15.95 \text{ molal solute}$$

$$= 8.17 \text{ }^{\circ}\text{C} \rightarrow \text{Boiling Point} = 8.17 \text{ }^{\circ}\text{C} + 100 \text{ }^{\circ}\text{C} = 108.17 \text{ }^{\circ}\text{C}$$

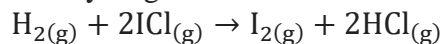
$$Q = mc\Delta T = (2 * 10^3 \text{ g of solution}) * 3.62 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1} * (108.17 - 21) \text{ }^{\circ}\text{C}$$

$$= 631.1 \text{ kJ}$$

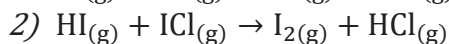
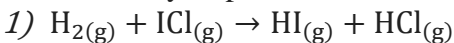
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5. The reaction between the gases iodine monochloride and hydrogen gas produces the two gases iodine and hydrogen chloride in the overall reaction:



This reaction is done in two elementary steps:



- a. If 2.4 moles of $\text{H}_2(\text{g})$ and 3.1 moles of $\text{ICl}(\text{g})$ react, which of the two gases is the limiting reagent?

From the overall reaction $\text{H}_2(\text{g}) + 2\text{ICl}(\text{g}) \rightarrow \text{I}_2(\text{g}) + 2\text{HCl}(\text{g})$, 2 moles of $\text{ICl}(\text{g})$ react for every 1 mole of $\text{H}_2(\text{g})$. Therefore, if 2.4 moles of $\text{H}_2(\text{g})$ and 3.1 moles of $\text{ICl}(\text{g})$ react, **$\text{ICl}(\text{g})$ must be the limiting reactant.**

- b. Name an intermediate in the overall reaction.

$\text{HI}(\text{g})$ is the intermediate in the overall reaction since $\text{HI}(\text{g})$ is a product of the first step of the reaction and a reactant in the second step of the reaction. It is synthesized and used up within the overall reaction.

- c) Suppose that the second step is slow and the first step is fast. Write a rate law using the rate constant k and concentrations of the appropriate gases. (3 points)

The second step is rate-determining since it is slow, and the rate law for this step can be written as $\text{rate} = k_2[\text{HI}][\text{ICl}]$. However, the first step must also be considered in order to substitute for the concentration of the intermediate HI in the rate law. Considering both sides of the reaction in the first step, $k_1[\text{H}_2][\text{ICl}] = k_{-1}[\text{HI}][\text{HCl}]$ can be written with k_1 being the rate constant for the forward reaction while k_{-1} is the rate constant for the backwards reaction. $[\text{HI}]$ can be substituted for with the expression $\frac{k_1[\text{H}_2][\text{ICl}]}{k_{-1}[\text{HCl}]}$. This makes the rate law $= k_2\left(\frac{k_1[\text{H}_2][\text{ICl}]}{k_{-1}[\text{HCl}]}\right)[\text{ICl}]$. Simplifying this expression and using an overall rate constant k , the rate law becomes **$k\frac{[\text{H}_2][\text{ICl}]^2}{[\text{HCl}]}$**