

## Practice Exam for Exam 2

1. Potassium permanganate solution reacts with sodium oxalate solution in the presence of sulfuric acid in an oxidation-reduction reaction. Two of the products of this reaction are the manganese(II) ion and carbon dioxide.

a. Write a *balanced complete chemical equation* for this reaction.



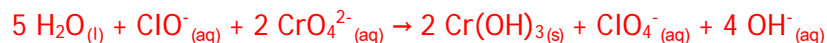
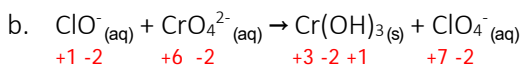
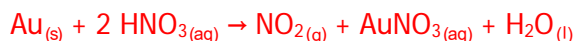
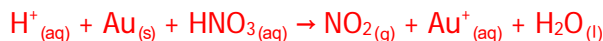
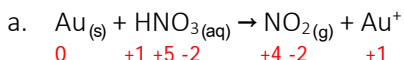
- b. If 25.43 mL of 0.02532 M potassium permanganate solution reacts with 14.64 mL of 0.09383 M sodium oxalate solution, how many *liters of carbon dioxide gas* are produced at 749.2 mmHg and 23.5°C?

$$\begin{aligned} ?\text{L CO}_2 = 25.43 \text{ mL KMnO}_4 &\times \frac{0.02532 \text{ mol KMnO}_4}{1000 \text{ mL KMnO}_4} \times \frac{10 \text{ mol CO}_2}{2 \text{ mol KMnO}_4} \\ &\times \frac{(62.364 \text{ L mmHg mol}^{-1} \text{ K}^{-1})(296.7 \text{ K})}{(749.2 \text{ mmHg})} = 0.07950 \text{ L CO}_2 \end{aligned}$$

$$\begin{aligned} ?\text{L CO}_2 = 14.64 \text{ mL Na}_2\text{C}_2\text{O}_4 &\times \frac{0.09383 \text{ mol Na}_2\text{C}_2\text{O}_4}{1000 \text{ mL Na}_2\text{C}_2\text{O}_4} \times \frac{10 \text{ mol CO}_2}{5 \text{ mol Na}_2\text{C}_2\text{O}_4} \\ &\times \frac{(62.364 \text{ L mmHg mol}^{-1} \text{ K}^{-1})(296.7 \text{ K})}{(749.2 \text{ mmHg})} = 0.06785 \text{ L CO}_2 \end{aligned}$$

0.06785 L of CO<sub>2</sub> will be produced.

2. Balance the following *oxidation-reduction reactions*. Give the molecular equation if possible.



3. A researcher determines the molar mass of a gas by using the time it takes for 1.00 mole of the gas to effuse compared to the amount of time it takes for 1.00 mol of argon to effuse at the same temperature in the same apparatus. She measures the time for argon to effuse as 6.34 minutes. 1.00 mole of the unknown gas takes 8.25 minutes to effuse.
- a. What is the *density, in g L<sup>-1</sup>*, of the unknown gas at STP?

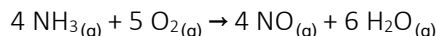
$$\frac{t_{unk}}{t_{Ar}} = \sqrt{\frac{M_{unk}}{M_{Ar}}} \Rightarrow M_{unk} = \left(\frac{t_{unk}}{t_{Ar}}\right)^2 M_{Ar} = \left(\frac{8.25 \text{ min}}{6.34 \text{ min}}\right)^2 (39.948 \text{ g mol}^{-1}) = 67.6 \text{ g mol}^{-1}$$

$$d = \frac{MP}{RT} = \frac{(67.6 \text{ g mol}^{-1})(1.00 \text{ atm})}{(0.082058 \text{ L atm mol}^{-1} \text{ K}^{-1})(273.15 \text{ K})} = 3.02 \text{ g L}^{-1}$$

- b. At what *Kelvin temperature* would the unknown gas have an rms speed of 456.2 m s<sup>-1</sup>?

$$u_{rms} = \sqrt{\frac{3RT}{M}} \Rightarrow T = \frac{Mu_{rms}^2}{3R} = \frac{(67.6 \times 10^{-3} \text{ kg mol}^{-1})(456.2 \text{ m s}^{-1})^2}{3(8.314 \text{ J mol}^{-1} \text{ K}^{-1})} = 564 \text{ K}$$

4. Use the data in the table provided to calculate the amount of energy, in kJ, produced or consumed when 100.00 grams of each of the reactants below is converted into products.



$$\Delta H = \sum_{\text{products}} n\Delta H_f^\circ - \sum_{\text{reactants}} n\Delta H_f^\circ$$

$$= \left[ \left( \frac{4 \text{ mol NO}}{\text{mol rxn}} \right) \left( \frac{90.29 \text{ kJ}}{\text{mol NO}} \right) + \left( \frac{6 \text{ mol H}_2\text{O}}{\text{mol rxn}} \right) \left( \frac{-241.8 \text{ kJ}}{\text{mol H}_2\text{O}} \right) \right] - \left[ \left( \frac{4 \text{ mol NH}_3}{\text{mol rxn}} \right) \left( \frac{-45.90 \text{ kJ}}{\text{mol NH}_3} \right) + \left( \frac{5 \text{ mol O}_2}{\text{mol rxn}} \right) \left( \frac{0.00 \text{ kJ}}{\text{mol O}_2} \right) \right]$$

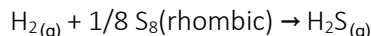
$$= -906.04 \text{ kJ mol}^{-1}$$

$$100.0 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.0305 \text{ g NH}_3} \times \frac{1 \text{ mol rxn}}{4 \text{ mol NH}_3} \times \frac{-906.04 \text{ kJ}}{\text{mol rxn}} = -1.330 \times 10^3 \text{ kJ}$$

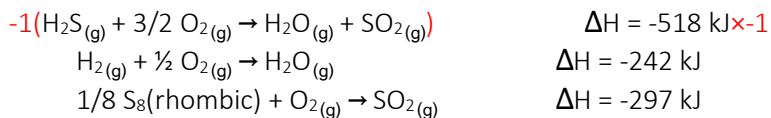
$$100.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{31.9988 \text{ g O}_2} \times \frac{1 \text{ mol rxn}}{5 \text{ mol O}_2} \times \frac{-906.04 \text{ kJ}}{\text{mol rxn}} = -566.3 \text{ kJ}$$

566.3 kJ of energy will be produced by the mixture.

5. Hydrogen Sulfide, H<sub>2</sub>S, is a poisonous gas with the odor of rotten eggs. The reaction for the formation of H<sub>2</sub>S from the elements is



Use Hess's Law to obtain the enthalpy change for this reaction from the following enthalpy changes:



$$\Delta H = (-1)(-518 \text{ kJ}) + (1)(-242 \text{ kJ}) + (1)(-297 \text{ kJ}) = -21 \text{ kJ}$$

6. A mixture contains calcium carbonate, CaCO<sub>3</sub>, and magnesium carbonate, MgCO<sub>3</sub>. A sample of this mixture weighing 7.85 g was reacted with excess hydrochloric acid. The reactions are



If the sample reacted completely and produced 1.94 L of carbon dioxide, CO<sub>2</sub>, at 25°C and 785 mmHg, what were the *percentages of CaCO<sub>3</sub> and MgCO<sub>3</sub>* in the mixture?

$$m_{\text{CaCO}_3} + m_{\text{MgCO}_3} = 7.85 \text{ g} \Rightarrow m_{\text{CaCO}_3} = 7.85 - m_{\text{MgCO}_3}$$

$$n_{\text{CO}_2} = \frac{PV}{RT} = \frac{(785 \text{ mmHg})(1.94 \text{ L})}{(62.364 \text{ L mmHg mol}^{-1} \text{ K}^{-1})(298 \text{ K})} = 0.08194 = m_{\text{CaCO}_3} \left( \frac{1 \text{ mol CaCO}_3}{100.09 \text{ g CaCO}_3} \right) \left( \frac{1 \text{ mol CO}_2}{1 \text{ mol CaCO}_3} \right) + m_{\text{MgCO}_3} \left( \frac{1 \text{ mol MgCO}_3}{84.314 \text{ g MgCO}_3} \right) \left( \frac{1 \text{ mol CO}_2}{1 \text{ mol MgCO}_3} \right)$$

$$0.08194 = \frac{(7.85 - m_{\text{MgCO}_3})}{100.09} + \frac{m_{\text{MgCO}_3}}{84.314} = \left( (7.85 - m_{\text{MgCO}_3})(84.314) \right) + m_{\text{MgCO}_3}(100.09) = (0.08194)(100.09)(84.314) = 691.4906980244$$

$$691.4906980244 = 661.8649 - 84.314 m_{\text{MgCO}_3} + 100.09 m_{\text{MgCO}_3} = 661.8649 + 15.776 m_{\text{MgCO}_3}$$

$$691.4906980244 - 661.8649 = 29.6438 = 15.776 m_{\text{MgCO}_3}$$

$$m_{\text{MgCO}_3} = \frac{29.6438}{15.776} = 1.88 \text{ g} \Rightarrow m_{\text{CaCO}_3} = 7.85 \text{ g} - 1.88 \text{ g} = 5.97 \text{ g}$$

$$\% \text{MgCO}_3 = \frac{1.88 \text{ g}}{7.85 \text{ g}} \times 100 = 23.95\% \quad \text{and} \quad \% \text{CaCO}_3 = 100 - 23.95 = 76.05\%$$