

1. A gaseous substance is found to have the following percent composition: 37.48% C, 12.58% H, and the rest is O. When 1.074 g of this substance is heated to a temperature of 71.6°C at 743.8 mmHg it occupies a volume of 969.0 mL. What is the *molecular formula* of the compound?

$$37.48 \text{ g C} \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 3.12055 \text{ mol C} \div 3.12055 = 1.0000$$

$$12.58 \text{ g H} \times \frac{1 \text{ mol H}}{1.00794 \text{ g H}} = 12.48090 \text{ mol H} \div 3.12055 = 3.9996$$

$$(100.00 - 37.48 - 12.58) \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 3.12136 \text{ mol O} \div 3.12055 = 1.0003$$

The empirical formula is CH₄O, which has an empirical mass of 32.0419 g mol⁻¹.
To find the molecular formula we need the molar mass.

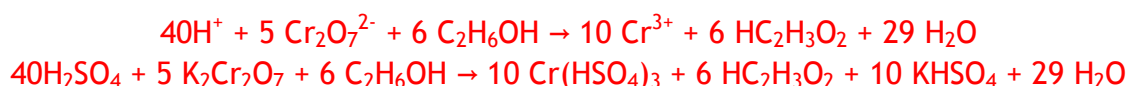
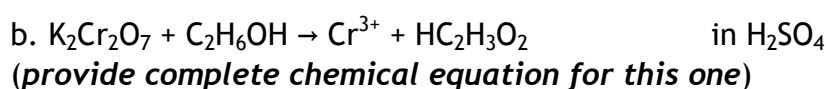
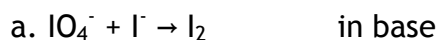
$$MP = dRT$$

$$M = \frac{dRT}{P} = \frac{\left(\frac{1.074 \text{ g}}{969.0 \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)(62.364 \text{ L mmHg mol}^{-1} \text{ K}^{-1})(344.8 \text{ K})}{743.8 \text{ mmHg}} = 32.04 \text{ g mol}^{-1}$$

$$n = \frac{\text{molar mass}}{\text{empirical mass}} = \frac{32.04 \text{ g mol}^{-1}}{32.0419 \text{ g mol}^{-1}} = 1$$

The molecular formula is CH₄O.

2. Balance the following oxidation-reduction reactions.



3 Acetone vapor ($\text{C}_3\text{H}_6\text{O}$) burns to produce carbon dioxide gas and water vapor. What **volume, in L, of carbon dioxide** gas at STP can be produced when 145.6 L of acetone vapor at 76.5°C and 1.036 atm reacts with 613.2 L of oxygen gas at 86.9°C and 732.5 mmHg?

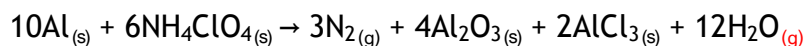


$$\frac{(145.6 \text{ L})(1.036 \text{ atm})}{(0.082058 \text{ L atm mol}^{-1} \text{ K}^{-1})(349.7 \text{ K})} \times \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_6\text{O}} \times \frac{22.414 \text{ L CO}_2}{1 \text{ mol CO}_2} = 353.5 \text{ L CO}_2$$

$$\frac{(613.2 \text{ L})(732.5 \text{ mmHg})}{(62.364 \text{ L mmHg mol}^{-1} \text{ K}^{-1})(360.1 \text{ K})} \times \frac{3 \text{ mol CO}_2}{4 \text{ mol O}_2} \times \frac{22.414 \text{ L CO}_2}{1 \text{ mol CO}_2} = 336.2 \text{ L CO}_2$$

336.2 L CO_2 can be produced at STP.

4. Calculate the **amount of energy, in J**, released when 25.0 metric tons (1 metric ton = 1000 kg) of Aluminum metal reacts with excess ammonium perchlorate according to the reaction



Substance	ΔH_f° (kJ mol ⁻¹)
$\text{NH}_4\text{ClO}_{4(s)}$	-295.3
$\text{Al}_2\text{O}_{3(s)}$	-1675.7
$\text{AlCl}_{3(s)}$	-704.2
$\text{H}_2\text{O}_{(g)}$	-241.818

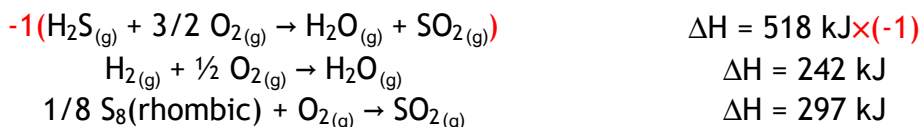
$$\begin{aligned} \Delta H &= \sum_{\text{products}} n\Delta H_f^\circ - \sum_{\text{reactants}} n\Delta H_f^\circ \\ &= \left[\left(\frac{3 \text{ mol N}_2}{\text{mol rxn}} \right) \left(\frac{0.00 \text{ kJ}}{\text{mol N}_2} \right) + \left(\frac{4 \text{ mol Al}_2\text{O}_3}{\text{mol rxn}} \right) \left(\frac{-1675.7 \text{ kJ}}{\text{mol Al}_2\text{O}_3} \right) + \left(\frac{2 \text{ mol AlCl}_3}{\text{mol rxn}} \right) \left(\frac{-704.2 \text{ kJ}}{\text{mol AlCl}_3} \right) + \left(\frac{12 \text{ mol H}_2\text{O}}{\text{mol rxn}} \right) \left(\frac{-241.818 \text{ kJ}}{\text{mol H}_2\text{O}} \right) \right] \\ &\quad - \left[\left(\frac{10 \text{ mol Al}}{\text{mol rxn}} \right) \left(\frac{0.00 \text{ kJ}}{\text{mol Al}} \right) + \left(\frac{6 \text{ mol NH}_4\text{ClO}_4}{\text{mol rxn}} \right) \left(\frac{-295.3 \text{ kJ}}{\text{mol NH}_4\text{ClO}_4} \right) \right] \\ &= -9241.2 \text{ kJ (mol rxn)}^{-1} \end{aligned}$$

$$25.0 \text{ mt Al} \times \frac{10^3 \text{ kg}}{1 \text{ mt}} \times \frac{10^3 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol Al}}{26.9898 \text{ g Al}} \times \frac{1 \text{ mol rxn}}{10 \text{ mol Al}} \times \frac{-9241.2 \text{ kJ}}{\text{mol rxn}} \times \frac{10^3 \text{ J}}{1 \text{ kJ}} = -8.56 \times 10^{11} \text{ J}$$

5. Hydrogen sulfide, H₂S, is a poisonous gas with the odor of rotten eggs. The reaction for the formation of H₂S from the elements is

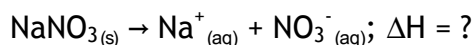


Use Hess' law to obtain *the enthalpy change* for this reaction from the following enthalpy changes:



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

6. When 15.3 g of sodium nitrate, NaNO₃, was dissolved in water in a calorimeter, the temperature fell from 25.00°C to 21.56°C. If the heat capacity of the solution and the calorimeter is 1071 J/°C, what is the *enthalpy change* when 1 mol of sodium nitrate dissolves in water? The solution process is



$$q_{cal} = C\Delta T = (1071 \text{ J/}^\circ\text{C})(21.56^\circ\text{C} - 25.00^\circ\text{C}) = -3684.24 \text{ J}$$

$$q_{rxn} = -q_{cal} = 3684.24 \text{ J}$$

$$\Delta H = \frac{q_{rxn}}{n_{\text{NaNO}_3}} = \frac{3684.24 \text{ J}}{15.3 \text{ g NaNO}_3} \times \frac{84.9947 \text{ g NaNO}_3}{1 \text{ mol NaNO}_3} \times \frac{1 \text{ mol NaNO}_3}{1 \text{ mol rxn}} = 2.05 \times 10^4 \text{ J (mol rxn)}^{-1}$$

7. A particular gas is found to have a root-mean-square speed, u_{rms} , of 337.1 m s^{-1} at standard thermodynamic temperature.

a. What is the *molar mass of the gas in g mol^{-1}* ?

$$u_{rms} = \sqrt{\frac{3RT}{M}} \Rightarrow M = \frac{3RT}{u_{rms}^2} = \frac{3(8.314 \text{ J mol}^{-1} \text{ K}^{-1})(298.2 \text{ K})}{(337.1 \text{ m s}^{-1})^2}$$
$$= 6.5451772 \times 10^{-2} \text{ kg mol}^{-1} \times \frac{10^3 \text{ g}}{1 \text{ kg}} = 65.45 \text{ g mol}^{-1}$$

b. If 1.00 mol of this gas effuses in 15.67 min , how many *minutes* will it require for 1.00 mol of Nitrogen dioxide to effuse in the same apparatus?

The rate of effusion is inversely proportional to the time and the molar mass of the compound. Therefore, the time of effusion is directly proportional to the molar mass.

$$\frac{t_{N_2}}{t_{gas}} = \sqrt{\frac{M_{NO_2}}{M_{gas}}} \Rightarrow t_{NO_2} = \sqrt{\frac{M_{NO_2}}{M_{gas}}} t_{gas} = \sqrt{\frac{46.0055 \text{ g mol}^{-1}}{65.45 \text{ g mol}^{-1}}} (15.67 \text{ min}) = 13.14 \text{ min.}$$

Nitrogen dioxide will take 13.14 min to effuse in the same apparatus.