Chem1A, General Chemistry I

1.) Dimethylhydrazine is a carbon-hydrogen-nitrogen compound used in rocket fuels. When complete combusted, a 0.505 g sample of dimethylhydrazine yields 0.741 g CO_2 and 0.605 g H_2O. The molecular weight for dimethylhydrazine is 60.099 g/mol. What is its molecular formula?

Determine how much of the original sample was N by converting CO_2 to g C and H_2O to g H.

$0.741 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.202\underline{2}13361 \text{ g C}$

$0.605 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.067\underline{6}84795 \text{ g H}$

0.505 g compound - 0.202213361 g C – 0.067684795 g H = 0.235101844 g N

Convert each mass to moles and divide by the smallest.

$0.202 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.0168\underline{1}9317 \text{ mols C}/0.0168 \rightarrow 1 \text{ mol C}$

$0.067\underline{7} \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.067\underline{1}62698 \text{ mols H}/ 0.0168 \rightarrow 4 \text{ mol H}$

$0.23\underline{5} \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.016\underline{7}73733 \text{ mols N}/ 0.0168 \rightarrow 1 \text{ mol N therefore } CH_4N$

Calculate the empirical weight. Divide the molecular weight by the empirical.
(1)(12.01 g/mol) + (4)(1.008 g/mol) + (1)(14.01 g/mol) = 30.0$\underline{5}$2 g/mol
(60.099 g/mol)/(30.052 g/mol) = 2 so 2 x $(CH_4N) \rightarrow$ **$C_2H_8N_2$**

2.) A 293 mL sample of 2.1 M magnesium chloride is mixed with 109 mL of 1.5 M sodium hydroxide. A double displacement reaction is observed to occur.

    a.) Write the **balanced molecular equation** for this reaction.

        $CaCl_2(aq) + 2 \text{ NaOH}(aq) \rightarrow Ca(OH)_2(aq) + 2 \text{ NaCl}(aq)$

    b.) Write the **total (complete) ionic equation** for this reaction.

    $Ca^{2+}(aq) + 2 \text{ Cl}^-(aq) + 2 \text{ Na}^+(aq) + 2 \text{ OH}^-(aq) \rightarrow$

                $Ca^{2+}(aq) + 2 \text{ OH}^-(aq) + 2 \text{ Na}^+(aq) + 2 \text{ Cl}^-(aq)$
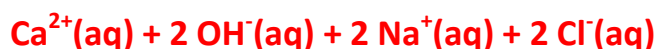
    c.) Write the **net ionic equation** for this reaction.

    No net ionic equation (no reaction)

    d.) Determine the theoretical yield of solid product, in g (molar mass = 74.10 g/mol).
       Convert each of the reactants to mass of $Ca(OH)_2$ produced. The least amount is the theoretical yield.

$293 \text{ mL } CaCl_2 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{2.1 \text{ mol } CaCl_2}{1 \text{ L } CaCl_2} \times \frac{1 \text{ mol } Ca(OH)_2}{1 \text{ mol } CaCl_2} \times \frac{74.10 \text{ g } Ca(OH)_2}{1 \text{ mol } Ca(OH)_2} = 45.\underline{5}9373 \text{ g } Ca(OH)_2$

$109 \text{ mL NaOH} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1.5 \text{ mols NaOH}}{1 \text{ L NaOH}} \times \frac{1 \text{ mol } Ca(OH)_2}{2 \text{ mol NaOH}} \times \frac{74.10 \text{ g } Ca(OH)_2}{1 \text{ mol } Ca(OH)_2} = 6.05\underline{7}675 \text{ g } Ca(OH)_2$

                Theoretical Yield: 6.06 g $Ca(OH)_2$

    e.) What is the limiting reactant? **NaOH – produces least amount of product.**

3.) Hydrogen can be produced by "water splitting" according to the following reaction:

$$H_2O(l) \rightarrow H_2(g) + \frac{1}{2} O_2(g)$$
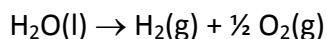
a.) What element is being oxidized?   **O – going from -2 to 0 oxidation state**

b.) What element is being reduced?   **H – going from +1 to 0 oxidation state**

c.) In a container, 0.0183 mols H_2 (g), and 0.0294 mols of O_2 (g) are held over water at 25°C. The vapor pressure of water at this temperature is 23.5 mmHg. The total pressure inside the container is 892 mmHg. Calculate the partial pressure of $H_2(g)$ and $O_2(g)$ in mmHg.
Subtract out the vapor pressure due to water from the total.
892 mmHg – 23.5 mmHg = 86<u>8</u>.5 mmHg due to $H_2 + O_2$

Calculate the mole fraction of H_2. Multiply it by the total pressure.
$\dfrac{0.0183 \text{ mols } H_2}{0.0183 \text{ mols } H_2 + 0.0294 \text{ mols } O_2}$ = 0.38<u>3</u>647799 ; (0.38<u>4</u>)(86<u>8</u>.5 mmHg) = 33<u>3</u>.1981132 $\rightarrow$

P$_{H2}$ = 333 mmHg

Find the partial pressure of the O_2 by subtracting P_{H2} from the total.
86<u>8</u>.5 mmHg – 333 mmHg = 53<u>5</u>.5 $\rightarrow$ **P$_{O2}$ = 536 mmHg**

4.) A 1.620 g sample of naphthalene ($C_{10}H_8$), is completely burned in a bomb calorimeter, resulting in a temperature increase of 8.44°C. If the heat of combustion of naphthalene is -5156 kJ/mol, calculate the heat capacity of the bomb calorimeter in kJ/°C.

Calculate the heat released by the sample.
1.620 g $C_{10}H_8$ × $\dfrac{1 \text{ mol } C_{10}H_8}{128.164 \text{ g } C_{10}H_8}$ × $\dfrac{-5156 \text{ kJ}}{1 \text{ mol } C_8H_{10}}$ = -65.1<u>7</u>212322 kJ released = 65.1<u>7</u> kJ absorbed by the calorimeter

Use q = C_{cal} x ΔT to solve for C_{cal}
65.1<u>7</u> kJ = C_{cal} (8.44 °C)
C_{Cal} = 7.7<u>2</u>1815547 $\rightarrow$ **7.72 kJ/°C**

5.) Calculate the change in internal energy (ΔE or ΔU) for a gas that releases 32.5 kJ of heat and has 52.3 kJ of work done on it by the surroundings, in kJ.

Releases heat: q = - 32.5 kJ ; work done by the surroundings: w = + 52.3 kJ

ΔE = -32.5 kJ + 52.3 kJ = 19.<u>8</u> $\rightarrow$ **19.8 kJ**

6.) Use the following heats of formation to solve for the $\Delta H°$ for the below reaction.

$$4\ NH_3(g) + 7\ O_2(g) \rightarrow 4\ NO_2(g) + 6\ H_2O(l)$$
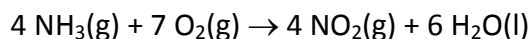
Compound	$\Delta H_f°$
$NH_3(g)$	- 46.19 kJ/mol
$NO_2(g)$	33.84 kJ/mol
$H_2O(l)$	- 285.83 kJ/mol

Recall: $\Delta H° = \sum n \cdot$ (products) - $\sum m \cdot$ (reactants)

$\Delta H° = [(4)(33.84\ kJ/mol) + (6)(-285.83\ kJ/mol)] - [(4)(-46.19\ kJ/mol)]$

$\Delta H° = [135.\underline{3}6\ kJ - 1714.\underline{9}8\ kJ] - [184.\underline{7}6\ kJ]$

$\Delta H° = -1394.\underline{8}6 \rightarrow$ **-1394.9 kJ/mol**

7.) A 1.231 g sample of an unknown gas is measured in a 250.0 mL container at 365.2 torr and 156.1 K. Calculate the molar mass of the unknown gas.

Convert V to L, torr to atm.

$250.0\ mL \times \frac{1\ L}{1000\ mL} = 0.250\underline{0}\ L$

$365.2\ torr \times \frac{1\ atm}{760\ torr} = 0.480\underline{5}26316\ atm$

Solve for n out of the Ideal Gas Law.

$(0.480\underline{5}\ atm)(0.250\underline{0}\ L) = n\ (0.08206\ L \cdot atm/(mol \cdot K)\ )(156.1\ K)$

$0.120\underline{1}31579\ L\cdot atm = n\ (12.8\underline{0}9566\ L\cdot atm/mol)$

$n = 0.009378\underline{8}271\ mols$

Divide the grams by the mols to solve for the molar mass.

$1.231\ g/0.00937\underline{8}\ mols = 131.\underline{2}608715 \rightarrow$ **131.3 g/mol**

8.) Rank the following gases in terms of (a) *increasing* root mean square velocity and (b) *decreasing* rate of effusion at a constant temperature: H_2, H_3, Ne, Cl_2, N_2, Xe.

    a.) Xe, Cl_2, N_3, N_2, H_3, H_2 (decreasing molar mass)

    b.) H_2, H_3, N_2, Ne, Cl_2, Xe (increasing molar mass)