PRACTICE FINAL EXAM

Chem1A, General Chemistry I

1.) Solid magnesium will react with hydrochloric acid to produce hydrogen gas and magnesium chloride. A 0.315 g sample of a mixture containing magnesium and other unreactive materials is dissolved completely in excess hydrochloric acid. The resulting hydrogen gas was collected in a 202 mL container *over water* at 23°C. The total pressure inside the container was measured to be 752 torr. At this temperature, the vapor pressure of water is 21.07 mmHg.

a.) Write the **balanced** chemical equation for this reaction, *including* phases.

$Mg(s) + \mathbf{2}\ HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$

b.) What element is being **reduced**? **Hydrogen (+1 → 0)**

c.) What element is being **oxidized**? **Magnesium (0 → +2)**

d.) Assuming the percent yield of the reaction was 100%, calculate the mass percent of magnesium in the original sample.

Use the Ideal Gas Law to solve for the mols $H_2(g)$ produced.

$P = 752 \text{ torr} - 21.07 \text{ torr } (1 \text{ mmHg} = 1 \text{ torr}) = 73\underline{0}.93 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.96\underline{1}7 \text{ atm}$

$V = 202 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.20\underline{2} \text{ L}$; $n = ?$; $T = 23 + 273.15 = 29\underline{6}.15 \text{ K}$

$(0.96\underline{1}7 \text{ atm})(0.20\underline{2} \text{ L}) = (n)(0.08206 \text{ L*atm/(mol*K)})(29\underline{6}.15 \text{ K})$

$(0.194\underline{2} \text{ L*atm}) = (n)(24.\underline{3}0 \text{ L*atm/mol})$

$n = 0.0079\underline{9}4 \text{ mols } H_2$

Use stoichiometry to convert to g Mg originally reacted.

$0.0079\underline{9}4 \text{ mols } H_2 \times \frac{1 \text{ mol Mg}}{1 \text{ mol } H_2} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 0.19\underline{4}3 \text{ g Mg}$

Calculate the mass percent of Mg in the original sample.

$\frac{0.1943 \text{ g Mg}}{0.315 \text{ g mixture}} \times 100\% = 61.\underline{6}9 \rightarrow \mathbf{61.7\%\ Mg}$

2.) A tennis ball weighs 56.7 g. Calculate the de Broglie wavelength in m for a ball traveling at 125 mph. (1 mi. = 1.61 km)

Use the de Broglie equation to solve for waverlength (λ): $\lambda = h/(mv)$

$h = 6.626 \times 10^{-34}$ J·s

$m = 56.7 \text{ g} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 0.056\underline{7} \text{ kg}$

$v = 125 \text{ mi./hr.} \times \frac{1.61 \text{ km}}{1 \text{ mi.}} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{1 \text{ hr.}}{60 \text{ min.}} \times \frac{1 \text{ min.}}{60 \text{ sec.}} = 55.\underline{9}0 \text{ m/s}$

$\lambda = \frac{6.626 \times 10^{-34} \text{ J*s}}{(0.0567 \text{ kg})(55.90 \frac{m}{s})} = 2.0\underline{9}0 \times 10^{-34} \rightarrow \mathbf{2.09 \times 10^{-34}\ m}$

3.) Nicotine is the addictive component of tobacco. An aqueous solution is made by dissolving 1.921 g nicotine into 48.92 g of water, changing the freezing point by -0.450°C. (K_f = 1.86 °C/m)

a.) Calculate the molar mass of nicotine.

Solve for the molality of the solution.

$0.450°C = (1.86\ °C/m)(m)$ so $m = \frac{0.450\ °C}{1.86\ °C/m} = 0.24\underline{1}9\ m$

Solve for the mols of solute from the molality.

$0.24\underline{1}9\ m = \frac{x \text{ mols solute}}{48.92 \text{ g water} \times \frac{1\text{ kg}}{1000\text{ g}}}$ so $x = (0.24\underline{1}9\ m)(0.0489\underline{2} \text{ kg water}) = 0.011\underline{8}3 \text{ mols solute}$

$\frac{1.921 \text{ g nicotine}}{0.01183 \text{ mols nicotine}} = 16\underline{2}.3 \rightarrow$ **162 g/mol**

b.) Nicotine contains only carbon, hydrogen, and nitrogen. Elemental analysis revealed a composition of 74.03% C, 8.70% H, and the rest N by mass. What is the molecular formula of nicotine?

Assume 100 g of sample. Convert each percent to mols, then ratio by the smallest.

$74.03 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 6.16\underline{4}0 \text{ mols C}/1.23\underline{2}6 \text{ mols} \rightarrow 5.00$ or 5 mols C

$8.70 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 8.6\underline{3}0 \text{ mols H}/1.23\underline{2}6 \text{ mols} \rightarrow 7.00$ or 7 mols H

$(100 - 74.03 - 8.70) = 17.2\underline{7} \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 1.23\underline{2}6 \text{ mols N}/1.23\underline{2}6 \text{ mols} \rightarrow 1$ mol N

Calculate the empirical weight. Divide the molecular by the empirical.

C_5H_7N: (5)(12.01 g/mol) + (7)(1.008 g/mol) + (1)(14.01 g/mol) = 81.116 g/mol

$\frac{162 \text{ g/mol}}{81.116 \text{ g/mol}} = 1.99$ or 2, so 2 x (C_5H_7N) = $\mathbf{C_{10}H_{14}N_2}$

4.) From a 0.871 M solution of barium nitrate, 267 mL are taken and mixed with 402 mL from a 0.487 M solution of ammonium sulfate. A double displacement reaction is observed.

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

$\mathbf{Ba(NO_3)_2(aq) + (NH_4)_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ NH_4NO_3(aq)}$

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

$\mathbf{Ba^{+2}(aq) + 2\ NO_3^-(aq) + 2\ NH_4^+(aq) + SO_4^{-2}(aq) \rightarrow}$

$\mathbf{BaSO_4(s) + 2\ NH_4^+(aq) + 2\ NO_3^-(aq)}$

c.) Write the **net ionic** equation for this reaction.

$\mathbf{Ba^{+2}(aq) + SO_4^{-2}(aq) \rightarrow BaSO_4(s)}$

d.) Calculate the **theoretical yield of solid precipitate**, in mols.

Convert each of the volumes to mols $BaSO_4$

$267 \text{ mL } Ba(NO_3)_2 \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{0.871 \text{ mols } Ba(NO_3)_2}{1 \text{ L } Ba(NO_3)_2} \times \frac{1 \text{ mol } BaSO_4}{1 \text{ mol } Ba(NO_3)_2} = 0.23\underline{2}5 \rightarrow 0.234 \text{ mols } BaSO_4$

$402 \text{ mL } (NH_4)_2SO_4 \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{0.487 \text{ mols } (NH_4)_2SO_4}{1 \text{ L } (NH_4)_2SO_4} \times \frac{1 \text{ mol } BaSO_4}{1 \text{ mol } (NH_4)_2SO_4} = 0.19\underline{5}7 \rightarrow 0.196 \text{ mols } BaSO_4$

Theoretical Yield: 0.196 mols $BaSO_4$

5.) Consider the following equation:

$$3\ C(graphite) + 4\ H_2(g) \rightarrow C_3H_8\ (g) \qquad \Delta H_{rxn} = ?$$

Use Hess's Law to determine ΔH_{rxn} from the following data:

$$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l) \qquad \Delta H_{rxn} = -\ 2219.9\ kJ/mol$$
$$C(graphite) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H_{rxn} = -393.5\ kJ/mol$$
$$H_2(g) + \tfrac{1}{2}\ O_2(g) \rightarrow H_2O(l) \qquad \Delta H_{rxn} = -285.8\ kJ/mol$$

Reverse Eqn. 1: $3\ CO_2(g) + 4\ H_2O(l) \rightarrow C_3H_8(g) + 5\ O_2(g)$ $(-1)(\Delta H_1)$

Multiply by 3 Eqn. 2: $3\ C(graphite) + 3\ O_2(g) \rightarrow 3\ CO_2(g)$ $(3)(\Delta H_2)$

Multiply by 4 Eqn. 3: $4\ H_2(g) + 4/2\ O_2(g) \rightarrow 4\ H_2O(l)$ $(4)(\Delta H_3)$

$3\ CO_2(g) + 4\ H_2O(l) + 3\ C(graphite) + 3\ O_2(g) + 4\ H_2(g) + 2\ O_2(g) \rightarrow$

$C_3H_8(g) + 5\ O_2(g) + 3\ CO_2(g) + 4\ H_2O(l)$

$3\ C(graphite) + 4\ H_2(g) \rightarrow C_3H_8(g)$

$\Delta H_T = (-1)(\Delta H_1) + (3)(\Delta H_2) + (4)(\Delta H_3) = (+2219.\underline{9}\ kJ/mol) + (-118\underline{0}.5\ kJ/mol) + (-114\underline{2}\ kJ/mol)$

$\Delta H_T = -10\underline{2}.6 \rightarrow$ **-103 kJ/mol**

6.) When potassium iodide (KI, 166.00 g/mol) is added to water, the formation of the solution is an endothermic process with an enthalpy change of 20.3 kJ/mol KI dissolved. At 23.5°C, enough KI is dissolved in water to make 150.0 mL of a 2.50 M KI solution. Calculate the final temperature of the water, in °C, given the heat capacity of liquid water is 4.184 J/(g °C) and the density of the solution is 1.72 g/mL.

Solve for the mols and grams of KI present in the solution.

$150.0\ \text{mL solution} \times \frac{1\ \text{L}}{1000\ \text{mL}} \times \frac{2.50\ \text{mols KI}}{1\ \text{L KI}} = 0.37\underline{5}\ \text{mols KI} \times \frac{166.00\ \text{g KI}}{1\ \text{mol KI}} = 62.\underline{2}5\ \text{g KI}$

Convert to the J of heat absorbed by the KI dissolving.

$0.375\ \text{mols KI} \times \frac{20.3\ \text{kJ}}{1\ \text{mol KI}} \times \frac{1000\ \text{J}}{1\ \text{kJ}} = 76\underline{1}2\ \text{J heat absorbed}$

Heat absorbed by KI ($76\underline{1}2$ J) = - heat lost by solution ($-76\underline{1}2$ J) = $m\ C_s\ \Delta T$

Find the mass of water via the density of total solution.

$150.0\ \text{mL solution} \times \frac{1.72\ \text{g solution}}{1\ \text{mL solution}} = 25\underline{8}\ \text{g solution} - 62.\underline{2}5\ \text{g KI} = 19\underline{5}.7\ \text{g water}$

Solve for T_f.

$-76\underline{1}2\ J = (19\underline{5}.7\ g)(4.184\ J/(g\ °C))(T_f - 23.5°C)$

$-76\underline{1}2\ J = (81\underline{9}.0\ J/°C)(T_f - 23.5\ °C)$

$-9.2\underline{9}4\ °C = T_f - 23.5°C$

$T_f = 14.\underline{2}0 \rightarrow$ **14.2 °C**

7.) Benzene (C_6H_6) is an aromatic organic hydrocarbon.

a.) **Balance** the following equation.

__2__ $C_6H_6(l)$ + __**15**__ $O_2(g) \rightarrow$ __**12**__ $CO_2(g)$ + __**6**__ $H_2O(l)$

b.) The enthalpy change associated with the above reaction of 2 mols of C_6H_6 is -6535 kJ/mol. Given the following information, calculate the standard enthalpy of formation for 1 mol of $C_6H_6(l)$, in kJ/mol.

Compound	$CO_2(g)$	$H_2O(l)$
$\Delta H_f°$ (kJ/mol)	-393.5	-285.8

$\Delta H_{rxn} = \sum$ m (products) - $\sum$ n (reactants)

ΔH_{rxn} = [(12)(ΔH $CO_2(g)$) + (6)(ΔH $H_2O(l)$)] – [(2)(ΔH $C_6H_6(l)$) + (15) (ΔH $O_2(g)$)

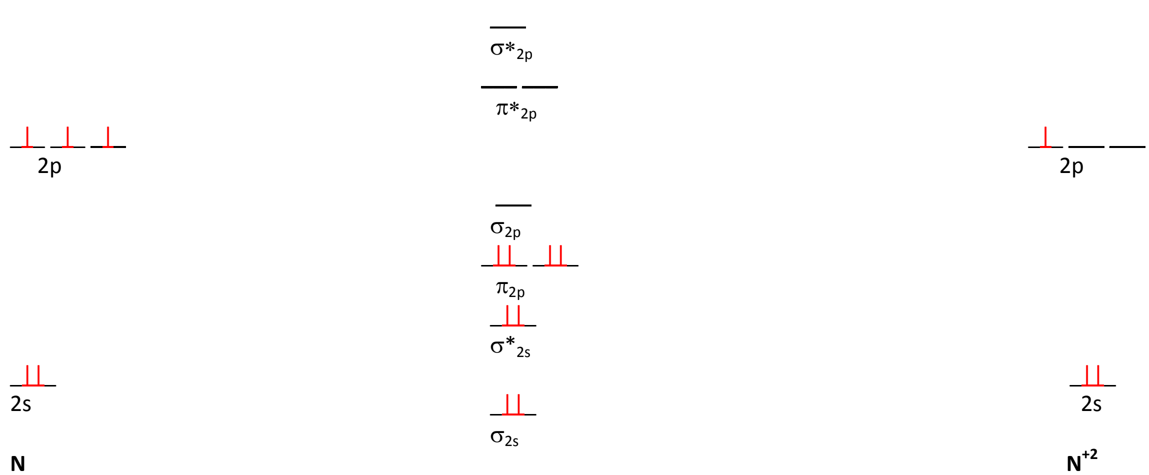
-6535 kJ/mol = [(12)(-393.5 kJ/mol) + (6)(-285.8 kJ/mol)] – [(2)(x) + (15)(0)]

-6535 kJ/mol = [-4722 kJ/mol – 1714.8 kJ/mol] – [2x]

-98.2 kJ/mol = - 2x

x = 49.1 kJ/mol → **49 kJ/mol**

8.) Draw the molecular orbital diagram for N_2^{+2}, ignoring core electrons. Calculate the bond order and determine whether the molecule is paramagnetic or diamagnetic.

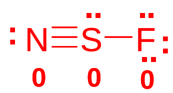


Bond order: ½(6 – 2) = **2**, stable. **Diamagnetic**

9.) Draw the most plausible Lewis structures for the following molecules, including all **resonance structures** and **formal charges**. Indicate the **electronic** and **molecular** geometry expected. Determine the **polarity** of the molecule, give the **hybridization** around each central atom and determine the **number of σ and π bonds** in the structure. (8 pts)

Element	N	S	F	I	Cl
eN	3.0	2.5	4.0	2.5	3.0

a.) NSF 18 total – 4 used – 16 needed = -2/2 → 1 more bond

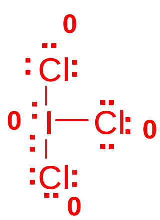


(expanded octet minimizes formal charges)

EG: **Trigonal Planar** ; MG: **Bent** ; **polar** (N-S (0.5, polar) and S-F (1.5, polar) don't cancel)

S: sp^2 hybridized; 2 σ bonds, 2 π bonds

b.) ICl_3 28 total – 6 used – 20 needed = +2 → expanded octet on iodine



EG: **Trigonal Bipyramidal** ; MG: **T-shaped** ; **polar** (I-Cl (0.5, polar) don't cancel)

I: sp^3d hybridized, 3 σ bonds

10.) One mole of photons contains 1799 kJ of energy.

a.) Calculate the energy per one photon, in J.

Divide by Avogadro's Number and convert to J.

$$\frac{1799 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ photons}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 2.9873 \times 10^{-18} \rightarrow \mathbf{2.987 \times 10^{-18}\ J/photon}$$

b.) Calculate the frequency of one photon, in Hz.

Use E = hν, so ν = E/h

$$\nu = \frac{2.987 \times 10^{-18} \text{ J}}{6.626 \times 10^{-34} \text{ J*s}} = 4.5079 \times 10^{15} \rightarrow \mathbf{4.508 \times 10^{15}\ Hz}$$

c.) Calculate the wavelength of one photon, in nm.

Use c = λ * ν, so λ = c/ν

$$\lambda = \frac{2.998 \times 10^{8} \text{ m/s}}{4.508 \times 10^{15} \text{ 1/s}} \times \frac{1 \text{ nm}}{10^{-9} \text{ m}} = 66.503 \rightarrow \mathbf{66.50\ nm}$$

11.) A gaseous hydrocarbon weighs 0.231 g and occupies a volume of 102 mL at 23.0°C at 749 mmHg. Calculate the molar mass of the unknown, in g/mol.

Use the Ideal Gas Law to solve for mols of unknown.

$P = 749 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.98\underline{5}5 \text{ atm}$; $V = 102 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.10\underline{2} \text{ L}$

$n = ?$; $T = 23.0 + 273.15 = 296.\underline{1}5 \text{ K}$

$(0.98\underline{5}5 \text{ atm})(0.10\underline{2} \text{ L}) = (n)(0.08206 \text{ L·atm/(mol·K)})(296.\underline{1}5 \text{ K})$

$(0.10\underline{0}5 \text{ L·atm}) = (n)(24.\underline{3}0 \text{ L·atm/mol})$

$n = 0.0041\underline{3}6$ mols

Divide grams unknown by mols unknown.

$(0.231 \text{ g})/(0.0041\underline{3}6 \text{ mols}) = 55.\underline{8}4 \rightarrow$ **55.8 g/mol**

12.) A "coffee-cup" calorimeter contains 100.0 mL of 0.300 M HCl at 20.3°C. When 1.82 g Zn(s) is added, a single displacement reaction is observed and the temperature rises to 30.5°C. Calculate the heat of reaction per mol Zn, in kJ/mol, assuming that the specific heat capacity of the solution is 4.184 J/(g °C), the density of hydrochloric acid is 1.18 g/mL, and no heat is lost.

Reaction: $Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Determine the limiting reactant.

$100.0 \text{ mL HCl} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.300 \text{ mols HCl}}{1 \text{ L HCl}} \times \frac{1 \text{ mol } ZnCl_2}{2 \text{ mols HCl}} = 0.015\underline{0} \text{ mols } ZnCl_2$ *theoretical yield*

$1.82 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \times \frac{1 \text{ mol } ZnCl_2}{1 \text{ mol Zn}} = 0.027\underline{8}3 \text{ mols } ZnCl_2$

Convert the theoretical yield back to mols Zn used.

$0.015\underline{0} \text{ mols } ZnCl_2 \times \frac{1 \text{ mol Zn}}{1 \text{ mol } ZnCl_2} = 0.015\underline{0} \text{ mols Zn used}$

Calculate the total mass of the solution.

$100.0 \text{ mL soln} \times \frac{1.18 \text{ g soln}}{1 \text{ mL soln}} = 11\underline{8} \text{ g soln}$

Use $q = mC_s\Delta T$ to solve for q.

$q = (11\underline{8} \text{ g})(4.814 \text{ J/(g °C)})(30.5°C - 20.3°C) = 50\underline{3}5 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 5.0\underline{3}5 \text{ kJ}$

Divide q/mols Zn used to find the heat of reaction.

$\Delta H = \frac{5.035 \text{ kJ}}{0.0150 \text{ mols}} = 33\underline{5}.7 \rightarrow$ **336 kJ/mol**