

PRACTICE EXAM: FINAL

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Chem20, Elementary Chemistry

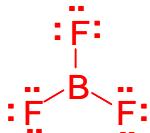
- 1.) Draw the Lewis structures for each of the following molecules, including all possible resonance structures. Give their expected **electronic** and **molecular** geometries and indicate whether they are polar or non-polar.

Element	B	F	O	Si	H	N
eN	2.0	4.0	3.5	1.9	2.1	3.1

a.) BF_3

POLAR/NONPOLAR: polar

24 total – 6 used – 20 needed = -2/2 → 1 more bond, but incomplete octet on B



$\Delta\text{eN} : (\text{F} - \text{B}) = 4.0 - 2.0 = 2.0$, polar bonds → cancel out

ELECTRONIC: trigonal planar

MOLECULAR: trigonal planar

b.) O_3

POLAR/NONPOLAR: nonpolar

18 total – 4 used – 16 needed = -2/2 → 1 more bond, resonance



$\Delta\text{eN} : (\text{O} - \text{O}) = 3.5 - 3.5 = 0$, nonpolar bonds

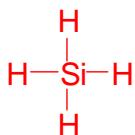
ELECTRONIC: trigonal planar

MOLECULAR: bent

c.) SiH_4

POLAR/NONPOLAR: nonpolar

8 total – 8 used – 0 needed = 0 → structure is plausible



$\Delta\text{eN} : (\text{H} - \text{Si}) = 2.1 - 1.9 = 0.2$, nonpolar bonds

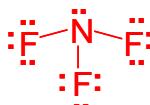
ELECTRONIC: tetrahedral

MOLECULAR: tetrahedral

d.) NF_3

POLAR/NONPOLAR: polar

26 total – 6 used – 20 needed = 0 → structure is plausible



$\Delta\text{eN} : (\text{F} - \text{N}) = 4.0 - 3.0 = 1.0$, polar bonds → do not cancel

ELECTRONIC: tetrahedral

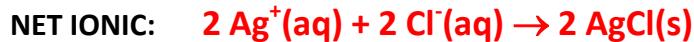
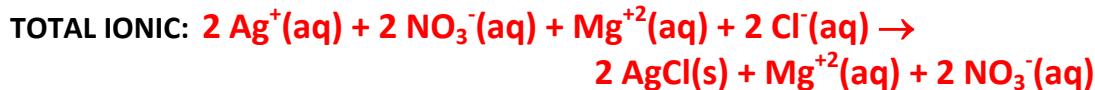
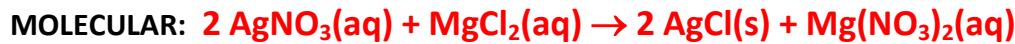
MOLECULAR: trigonal pyramidal

2.) Round the following numbers to **FIVE** significant figures. Convert to appropriate scientific notation.

- a.) 5002742.245 5.0027×10^6
- b.) 0.0000611105 6.1111×10^{-5}
- c.) 0.9830001 9.8300×10^{-1}
- d.) 10.028421 1.0028×10^1

3.) An aqueous solution of 0.225 M silver(I) nitrate is added to an aqueous solution of 203 mL of 0.125 M magnesium chloride.

- a.) Write the molecular, total ionic, and net ionic equations for the double displacement reaction.



- b.) How many mL of the 0.225 M silver nitrate solution need to be added to precipitate **all** the chloride ions? Assume a 100% yield.

Convert 203 mL MgCl_2 to mL AgNO_3 using molarity and stoichiometry.

$$203 \text{ mL } \text{MgCl}_2 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.125 \text{ mol MgCl}_2}{1 \text{ L MgCl}_2} \times \frac{2 \text{ mol AgNO}_3}{1 \text{ mol MgCl}_2} \times \frac{1 \text{ L AgNO}_3}{0.225 \text{ mol AgNO}_3} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 225.5 \rightarrow 226 \text{ mL AgNO}_3$$

3.) Convert the following measurements with appropriate significant figures.

a.) $1.52 \text{ kL} = ? \mu\text{L}$

$$1.52 \text{ kL} \times \frac{10^3 \text{ L}}{1 \text{ kL}} \times \frac{1 \mu\text{L}}{10^{-6} \text{ L}} = 1.52 \times 10^9 \rightarrow 1.52 \times 10^9 \mu\text{L}$$

b.) $942.5 \text{ cm}^2 = ? \text{ Dm}^2$

$$942.5 \text{ cm}^2 \times \left(\frac{10^{-2} \text{ m}}{1 \text{ cm}} \right)^2 \times \left(\frac{1 \text{ Dm}}{10^1 \text{ m}} \right)^2 = 9.425 \times 10^{-4} \rightarrow 9.425 \times 10^{-4} \text{ Dm}^2$$

4.) When strong acids react with metals, it results in an aqueous salt and hydrogen gas, as shown in the following **UNBALANCED** chemical reaction. Answer the following questions regarding this reaction.



a.) Balance the above chemical equation. Classify the reaction in as many ways as possible.



Single displacement, gas evolution, oxidation-reduction (redox)

b.) From a 6.2 M HCl(aq) solution, 25.1 mL are added to 20.3 g of Zn(s) and allowed to react. Determine the **limiting reactant** and **theoretical yield** of H₂(g), in mols.

Convert each reactant to mols of H₂.

$$20.3 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \times \frac{1 \text{ mol H}_2}{1 \text{ mol Zn}} = 0.3104 \text{ mols H}_2$$

$$25.1 \text{ mL HCl} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{6.2 \text{ mols HCl}}{1 \text{ L HCl}} \times \frac{1 \text{ mol H}_2}{2 \text{ mols HCl}} = 0.07781 \text{ mols H}_2$$

Limiting Reactant: HCl

Theoretical Yield: 0.0778 mols H₂

c.) How many g of the **reactant in excess** are left over after the reaction is complete?

Convert the theoretical yield back to g reactant in excess (Zn). Subtract from initial.

$$0.0778 \text{ mols H}_2 \times \frac{1 \text{ mol Zn}}{1 \text{ mol H}_2} \times \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}} = 5.087 \text{ g Zn used}$$

$$20.3 \text{ g Zn initial} - 5.087 \text{ g Zn used} = 15.21 \text{ g} \rightarrow \mathbf{15.2 \text{ g Zn in excess}}$$

d.) The resulting H₂(g) was collected in a 0.255 L container at 1.00 atm and 273.15 K. What was the **actual yield** of H₂(g), in mols?

Use the Ideal Gas Law and solve for n.

$$P = 1.00 \text{ atm} ; V = 0.255 \text{ L} ; n = ? ; R = 0.08206 \text{ L} \cdot \text{atm}/(\text{mol} \cdot \text{K}) ; T = 273.15 \text{ K}$$

$$(1.00 \text{ atm})(0.255 \text{ L}) = (n)(0.08206 \text{ L} \cdot \text{atm}/(\text{mol} \cdot \text{K})) (273.15 \text{ K})$$

$$0.255 \text{ L} \cdot \text{atm} = (n)(22.4146 \text{ L} \cdot \text{atm/mol})$$

$$n = 0.01137 \rightarrow \mathbf{0.0114 \text{ mols H}_2}$$

e.) What was the percent yield?

Recall: percent yield = (actual yield)/(theoretical yield) x 100%

$$\text{Percent yield} = \frac{0.0114 \text{ mols H}_2}{0.0778 \text{ mols H}_2} \times 100\% = 14.65 \rightarrow \mathbf{14.7\% \text{ yield}}$$

5.) How many grams of carbon tetrabromide would need to be added to 2.15 L of H₂O(l) to reduce the freezing point by 9.18°C? (Assume the density of water is 1.00 g/mL, K_f = 1.86°C/m)

Solve for the molality of the solution using the equation for freezing point depression.

$$9.18 \text{ } ^\circ\text{C} = (1.86 \text{ } ^\circ\text{C/m})(m) \text{ so } m = 4.935 \text{ m}$$

Convert 2.15 L H₂O into kg H₂O.

$$2.15 \text{ L water} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1.00 \text{ g water}}{1 \text{ mL}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 2.15 \text{ kg water}$$

Recall: molality = mols solute/kg solvent. Solve for mols solute.

$$4.935 \text{ m} = \frac{x \text{ mols solute}}{2.15 \text{ kg water}} \text{ so } x = (4.935 \text{ m})(2.15 \text{ kg water}) = 10.61 \text{ mols solute}$$



Convert mols solute to g.

$$10.61 \text{ mols CBr}_4 \times \frac{331.61 \text{ g CBr}_4}{1 \text{ mol CBr}_4} = 3518 \rightarrow 3520 \text{ g CBr}_4$$

6.) Write the electronic configurations for the following elements, with or without noble gas notation.

a.) Zn $[\text{Ar}]4s^23d^{10}$

b.) Cl $[\text{Ne}]3s^23p^5$

c.) Ne $1s^22s^22p^6$

7.) Complete the following table.

Atomic Symbol	Number of Electrons	Number of Protons	Number of Neutrons	Atomic Number	Mass Number	Atomic Notation
Xe	54	54	77	54	131	$^{131}_{54}\text{Xe}$
Al	10	13	14	13	27	$^{27}_{13}\text{Al}^{+3}$
Mo	40	42	54	42	96	$^{96}_{42}\text{Mo}^{+2}$
Br	36	35	45	35	80	$^{80}_{35}\text{Br}^{-1}$
Cl	18	17	18	17	35	$^{35}_{17}\text{Cl}^{-1}$

8.) A sample of an unknown compound containing only carbon, hydrogen, and oxygen is found to contain 39.99% carbon, 6.693% hydrogen, and the rest oxygen by mass. The molecular weight of the compound is 180.18 g/mol. Determine the molecular formula.

Solve for the percent oxygen.

$$100\% - 39.99\% - 6.693\% = 53.317\% \text{ O}$$

Assume 100 g sample. Convert each g element to mols and ratio by the smallest.

$$39.99 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.3297 \text{ mols C} / 3.3297 \text{ mols} = 1 \text{ mol C}$$

$$6.693 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.6398 \text{ mols H} / 3.3297 \text{ mols} = 1.99 \text{ or } 2 \text{ mols H}$$

$$53.317 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.3323 \text{ mols O} / 3.3297 \text{ mols} = 1.00 \text{ or } 1 \text{ mol O}$$

Calculate the empirical weight and divide it into the molecular weight.

$$\text{CH}_2\text{O}: (1)(12.01 \text{ g/mol}) + (2)(1.008 \text{ g/mol}) + (1)(16.00 \text{ g/mol}) = 30.026 \text{ g/mol}$$

$$\frac{180.18 \text{ g/mol}}{30.026 \text{ g/mol}} = 6.00 \times (\text{CH}_2\text{O}) = \text{C}_6\text{H}_{12}\text{O}_6$$

9.) A sample of 0.52 mols of gas occupies 344 mL. If 0.83 mols of gas were added, calculate the new volume of the gas, in mL.

Use Avogadro's Law: $V_1/n_1 = V_2/n_2$

$$V_1 = 344 \text{ mL} ; n_1 = 0.52 \text{ mols} ; V_2 = ? ; n_2 = 0.52 \text{ mols} + 0.83 \text{ mols} = 1.35 \text{ mols}$$

$$\frac{344 \text{ mL}}{0.52 \text{ mols}} = \frac{x}{1.35 \text{ mols}}$$

$$(344 \text{ mL})(1.35 \text{ mols}) = (x)(0.52 \text{ mols})$$

$$(464.4 \text{ mL} * \text{mols}) = (x)(0.52 \text{ mols})$$

$$x = 893 \rightarrow 890 \text{ mL}$$

10.) If 672 mL of a 9.51 M solution of phosphoric acid is diluted to a total volume of 1.239 L, what is the new solution's concentration, in molarity (M)?

Use $M_1V_1 = M_2V_2$

$$M_1 = 9.51 \text{ M} ; V_1 = 672 \text{ mL} ; M_2 = ? ; V_2 = 1.239 \text{ L} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 1239 \text{ mL}$$

$$(9.51 \text{ M})(672 \text{ mL}) = (M_2)(1239 \text{ mL})$$

$$6390 \text{ M} \cdot \text{mL} = (M_2)(1239 \text{ mL})$$

$$M_2 = 5.157 \rightarrow 5.16 \text{ M}$$

11.) A sample of water is measured to have $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-10} \text{ M}$.

a.) What is $[\text{OH}^-]$ in the sample?

$$\text{Use } 1.00 \times 10^{-14} = [\text{H}_3\text{O}^+] [\text{OH}^-]. \text{ Solve for } [\text{OH}^-] = 1.00 \times 10^{-14}/[\text{H}_3\text{O}^+]$$

$$\frac{1.00 \times 10^{-14}}{1.0 \times 10^{-10}} = 1.0 \times 10^{-4} \rightarrow 1.0 \times 10^{-4} \text{ M}$$

b.) What is the pH of the sample?

$$\text{Use } \text{pH} = -\log([\text{H}_3\text{O}^+])$$

$$\text{pH} = -\log(1.0 \times 10^{-4}) = 10 \rightarrow 1.0 \times 10^1$$

c.) What is the pOH of the sample?

$$\text{Use } 14.00 = \text{pH} + \text{pOH}$$

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 1.0 \times 10^1 = 4 \rightarrow 4$$

12.) Classify the following compounds as molecular or ionic. Name them appropriately.

a.) BaCl_2 **ionic type I: barium chloride**

b.) $\text{Mn}(\text{NO}_3)_2$ **ionic type II: manganese(II) nitrate**

c.) PCl_5 **molecular: phosphorous pentachloride**

d.) $\text{HNO}_3(\text{aq})$ **molecular, oxyacid: nitric acid**

13.) A line in the atomic spectrum of hydrogen has a wavelength (λ) of 486 nm. (Possibly)

Useful: $c = 2.998 \times 10^8 \text{ m/s}$, $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$)

a.) Convert this wavelength into frequency (ν) in Hz.

Convert 486 nm to m.

$$486 \text{ nm} \times \frac{10^{-9} \text{ m}}{1 \text{ nm}} = 4.86 \times 10^{-7} \text{ m}$$

Use $c = \lambda * \nu$, so $\nu = c/\lambda$

$$\nu = \frac{2.998 \times 10^8 \text{ m/s}}{4.86 \times 10^{-7} \text{ m}} = 6.168 \times 10^{16} \rightarrow 6.17 \times 10^{16} \text{ Hz}$$

b.) Convert this wavelength into energy (E) in J.

Use $E = h\nu$

$$E = (6.626 \times 10^{-34} \text{ J}\cdot\text{s})(6.17 \times 10^{16} \text{ 1/s}) = 4.088 \times 10^{-17} \rightarrow 4.09 \times 10^{-17} \text{ J}$$