

Solutions

DAWSON COLLEGE
DEPARTMENT OF CHEMISTRY & CHEMICAL TECHNOLOGY

FINAL EXAMINATION
INTRODUCTION TO COLLEGE CHEMISTRY
202-005-03 202-006-06 912-016-94

December 10, 2004
9:30 – 12:30 PM

Print your Name: _____

Student Number: _____

INSTRUCTORS: *Please circle the name of your instructor:*

J. Ali I. Dionne R. Squire

INSTRUCTIONS:

This exam set consists of **16** questions. Please ensure that you have a complete set.

Answer **all** questions in the space provided.

1. Calculators may not be shared. Programmable calculators are not permitted.
2. No books or extra paper are permitted.
3. In order to obtain full credit, you must show the method used to solve all problems involving calculations and express your answers to the correct number of significant figures. **A grade will be assigned for the proper use of significant figures and units.**
4. Your attention is drawn to the College policy on cheating.
5. A Periodic Table is provided. You may detach the Periodic Table.

USEFUL DATA:

Avogadro's number $N_A = 6.023 \times 10^{23} \text{ mol}^{-1}$

Gas constant $R = 0.08206 \text{ L}\cdot\text{atm}\cdot\text{mol}^{-1}\cdot\text{K}^{-1} = 8.314 \text{ kPa L}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$
 $= 8.314 \text{ J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$

1 atm = 760 mm Hg = 101.3 kPa = 760 torr

1 J = 1 kg·m²·s⁻²

MARK DISTRIBUTION

1. _____ /10

2. _____ /10

3. _____ /3

4. _____ /4

5. _____ /2

6. _____ /3

7. _____ /3

8. _____ /11

9. _____ /9

10. _____ /5

11. _____ /11

12. _____ /4

13. _____ /5

14. _____ /6

15. _____ /8

16. _____ /4

Significant figures _____ /1

Units _____ /1

TOTAL _____ **/100**

1. a) Carry out the following calculations. Assume the numbers represent measurements and express your answers to the proper number of significant figures.

i) $(12.688 - 10.0) \times (7.85 + 2.666) = \underline{\mathbf{28}}$

ii) $(12.61 + 0.22 + 0.037) \div 0.04 = \underline{\mathbf{3 \times 10^2}}$

- b) Express the following numbers in scientific notation.

i) $0.000771 = \underline{\mathbf{7.71 \times 10^{-4}}}$

ii) $157 = \underline{\mathbf{1.57 \times 10^2}}$

- c) Convert the following:

i) $205 \text{ K} = \underline{\mathbf{-68^\circ \text{C}}}$ °C

ii) $25.7 \text{ g} = \underline{\mathbf{2.57 \times 10^4 \text{ mg}}}$ mg

iii) $102^\circ \text{F} = \underline{\mathbf{39^\circ \text{C}}}$ °C

- d) Pumice is a volcanic rock that contains many trapped air bubbles. A 155 g sample of pumice is found to have a volume of 163 mL.

i) What is the density of pumice in $\text{g} \cdot \text{mL}^{-1}$? = $\mathbf{155 / 163 = 0.951 \text{ g/mL}}$

ii) What is the volume occupied by a 4.56 kg sample of pumice? =
 $\mathbf{4560 / 0.951 = 4.80 \times 10^3 \text{ mL}}$

- iii) Will pumice float or sink in ethyl alcohol?
(density of ethyl alcohol is 0.790 g/mL at 20°C) = $\mathbf{\text{It will sink}}$

2. a) Write the names of the following compounds:

i) FeSO_4 **iron (II) sulfate**

ii) KNO_2 **potassium nitrite**

iii) Ca(OH)_2 **calcium hydroxide**

iv) NiCO_3 **nickel (II) carbonate**

v) H_2SO_4 (aq) **sulfuric acid**

vi) MgO **magnesium oxide**

vii) H_3PO_4 (aq) **phosphoric acid**

viii) IF_3 **iodine trifluoride**

ix) $\text{Al}_2(\text{SO}_3)_3$ **aluminum sulfite**

x) BaH_2 **barium hydride**

b) Write the chemical formulas for the following compounds:

i) ammonium nitrate NH_4NO_3 _____

ii) aluminum oxide Al_2O_3 _____

iii) copper (I) sulfide Cu_2S _____

iv) perchloric acid HClO_4 _____

v) cobalt (II) bromide CoBr_2 _____

vi) nitric acid HNO_3 _____

vii) disulfur decafluoride S_2F_{10} _____

viii) silver chloride AgCl _____

ix) copper (II) chloride dihydrate $\text{CuCl}_2 \cdot 2\text{H}_2\text{O}$ _____

x) sodium cyanide NaCN _____

3. a) Determine the oxidation state (charge) of each atom in the following compounds:

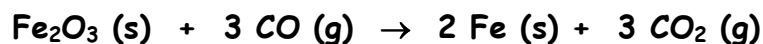
i) KMnO_4 K: +1 Mn: +7 O: -2

ii) Na_2O_2 Na: +1 O: -1

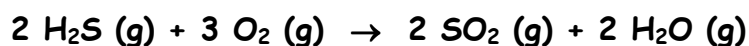
iii) $\text{Cr}_2\text{O}_7^{2-}$ Cr: +6 O: -2

4. Write and balance the chemical equations for the reaction between:

a) solid iron (III) oxide and gaseous carbon monoxide to form iron metal and gaseous carbon dioxide.

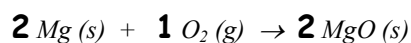


b) gaseous hydrogen sulfide and oxygen to form sulfur dioxide gas and water vapor.

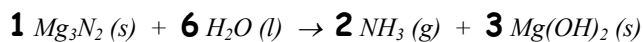


5. Balance the following chemical equations:

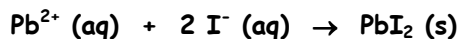
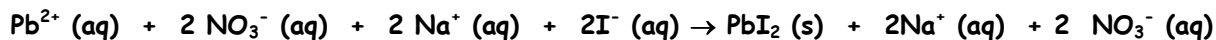
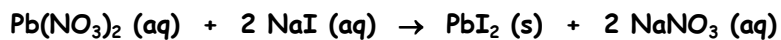
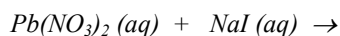
a) $Mg (s) + O_2 (g) \rightarrow MgO (s)$



b) $Mg_3N_2 (s) + H_2O (l) \rightarrow NH_3 (g) + Mg(OH)_2 (s)$



6. Write the balanced molecular, complete ionic and net ionic equations for:



7. Using check marks in the appropriate boxes classify each of the reactions given below as:

(I) Oxidation-Reduction (Redox) (II) Acid-Base (III) Precipitation (Ppt.)

Note that for each reaction more than one choice may apply.

| | I. Redox | II. Acid-Base | III. Ppt. |
|--|----------|---------------|-----------|
| a) $ZnBr_2(aq) + 2AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2AgBr(s)$ | | | ✓ |
| b) $H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow 2H_2O(l) + BaSO_4(s)$ | | ✓ | ✓ |
| c) $Ca(s) + H_2(g) \rightarrow CaH_2(s)$ | ✓ | | |
| d) $HC_2H_3O_2(aq) + CsOH(aq) \rightarrow H_2O(l) + CsC_2H_3O_2(aq)$ | | ✓ | |
| e) $Zn(s) + 2HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)$ | ✓ | | |

8. Complete the table below by providing the symbol of each atom and putting a check mark (✓) in the box that best describes the character of the atom:

| Element | Symbol of the Atom | Metal | Metalloid | Nonmetallic gas | Nonmetallic solid |
|------------|--------------------|-------|-----------|-----------------|-------------------|
| Sodium | <i>Na</i> | ✓ | | | |
| Copper | Cu | ✓ | | | |
| Carbon | C | | | | ✓ |
| Boron | B | | ✓ | | |
| Fluorine | F | | | ✓ | |
| Phosphorus | P | | | | ✓ |

- b) The following are some physical and chemical properties of metals and nonmetals. Match the stated properties in column one with the type of element (**metal** or **nonmetal**) that can exhibit the given property. State your answer in column two

| Properties | Match |
|--|-----------------|
| Have high melting point | <i>Metal</i> |
| Have no lustre | Nonmetal |
| Mostly hard but malleable | Metal |
| May combine with each other | Nonmetal |
| Have high electrical conductivity | Metal |
| Most have high densities | Metal |
| Will generally not be ductile but rather brittle | Nonmetal |

9. Complete the following table by providing the missing information:

| Nuclear Symbol | Atomic Number | Mass Number | Number of Neutrons | Number of Electrons | Number of Protons |
|-------------------------|---------------|-------------|--------------------|---------------------|-------------------|
| ${}_{16}^{32}\text{S}$ | 16 | 32 | 16 | 16 | 16 |
| ${}_{35}^{80}\text{Br}$ | 35 | 80 | 45 | 35 | 35 |
| ${}_{12}^{24}\text{Mg}$ | 12 | 24 | 12 | 12 | 12 |
| ${}_{3}^{7}\text{Li}$ | 3 | 7 | 4 | 3 | 3 |

10. Answer true or false for each of the following questions below (circle your choice):

- a) In a chemical reaction matter can be created and destroyed. T F
- b) Neutrons and protons are subatomic particles found in the nucleus of an atom. T F
- c) When atoms combine in a chemical reaction to form compounds they do so in simple whole number ratio. T F
- d) Atoms of one element are usually similar to atoms of other elements. T F
- e) Neutrons and protons have the same charge T F

11. a) Analysis of a sample of ibuprofen, the active ingredient in Advil, shows that the sample contains 7.568 g of carbon 0.881 g of hydrogen, and 1.551 g of oxygen, Use these data to calculate the empirical formula of ibuprofen.

| Atom | Mass | Moles | Simplest ratio | Integral Ratio |
|------|-------|---------------------------------|---------------------------------|----------------|
| C | 7.568 | $\frac{7.568}{12.011} = 0.6301$ | $\frac{0.6301}{0.09694} = 6.50$ | 13 |
| H | 0.881 | $\frac{0.881}{1.008} = 0.874$ | $\frac{0.874}{0.09694} = 9.02$ | 18 |
| O | 1.551 | $\frac{1.551}{16.00} = 0.09694$ | 1.000 | 2 |

Empirical Formula: $C_{13}H_{18}O_2$

- b) Ethylene, a compound that contains only carbon and hydrogen, is commercially used as a fruit-ripening agent. A sample of this compound is burned in a combustion analysis apparatus and 3.14 g of CO_2 and 1.29 g of H_2O are produced. What is the empirical formula of ethylene? If the molar mass of ethylene is 28 g/mol, what is the molecular formula of ethylene?

$$n_{CO_2} = \frac{3.14 \text{ g } CO_2}{44.01 \text{ g} \cdot \text{mol}^{-1}} = 0.0713 \text{ mol}$$

$$n_{H_2O} = \frac{1.29 \text{ g } H_2O}{18.02 \text{ g} \cdot \text{mol}^{-1}} = 0.0716 \text{ mol}$$

$$n_C = 0.0713 \text{ mol } CO_2 \left(\frac{1 \text{ mol } C}{1 \text{ mol } CO_2} \right) = 0.0713 \text{ mol } C$$

$$n_H = 0.0716 \text{ mol } H_2O \left(\frac{2 \text{ mol } H}{1 \text{ mol } H_2O} \right) = 0.143 \text{ mol } H$$

$$\begin{aligned} \text{Ratios: } \frac{n_C}{n_C} &= \frac{0.0713}{0.0713} = 1.00 \approx 1 \\ \frac{n_H}{n_C} &= \frac{0.143}{0.0713} = 2.01 \approx 2 \end{aligned}$$

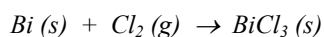
Empirical Formula CH_2

Empirical Formula Mass = $(12) + 2(1) = 14$

$$\frac{\text{Molar Mass}}{\text{Formula Mass}} = \frac{28}{14} = 2$$

Molecular Formula C_2H_4

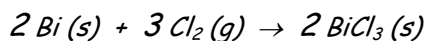
12. If 3.45 g bismuth metal, Bi, is reacted with chlorine gas according to the unbalanced chemical equation:



calculate the mass in grams of chlorine needed to completely react with the bismuth metal and the mass in grams of bismuth (III) chloride formed.

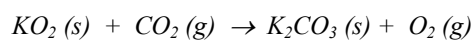
Molar Mass: Bi = 208.98 Cl₂ = 2(35.453) = 70.906

BiCl₃ = (208.98) + 3(35.453) = 315.34



$$\left. \begin{aligned} n_{\text{Bi}} &= \frac{3.45 \text{ g}}{208.98 \text{ g} \cdot \text{mol}^{-1}} = 0.0165 \text{ mol Bi} \\ n_{\text{Cl}_2} &= 0.0165 \text{ mol Bi} \left(\frac{3 \text{ mol Cl}_2}{2 \text{ mol Bi}} \right) = 0.0248 \text{ mol} \\ n_{\text{BiCl}_3} &= 0.0165 \text{ mol Bi} \left(\frac{2 \text{ mol BiCl}_3}{2 \text{ mol Bi}} \right) = 0.0165 \text{ mol} \\ m_{\text{Cl}_2} &= 0.0248 \text{ mol} \left(\frac{70.906 \text{ g Cl}_2}{1 \text{ mol}} \right) = 1.76 \text{ g Cl}_2 \\ m_{\text{BiCl}_3} &= 0.0165 \text{ mol} \left(\frac{315.34 \text{ g BiCl}_3}{1 \text{ mol}} \right) = 5.20 \text{ g BiCl}_3 \end{aligned} \right\}$$

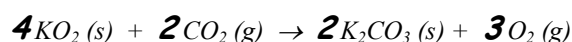
13. When 2.50 g potassium superoxide, KO_2 , reacts with 4.50 g carbon dioxide according to the unbalanced chemical equation:



0.799 g oxygen gas are produced. Calculate:

- The theoretical yield of oxygen.
- The percent yield of oxygen in this reaction.

Molar Mass: $\text{KO}_2 = 71.10$ $\text{CO}_2 = 44.01$ $\text{O}_2 = 32.00$ $\text{K}_2\text{CO}_3 = 138.21$



$$n_{\text{KO}_2} = \frac{2.50 \text{ g}}{71.10 \text{ g} \cdot \text{mol}^{-1}} = 0.0352 \text{ mol KO}_2$$

$$n_{\text{CO}_2} = \frac{4.50 \text{ g}}{44.01 \text{ g} \cdot \text{mol}^{-1}} = 0.102 \text{ mol CO}_2$$

Assume KO_2 is the Limiting Reactant

$$\text{Number of moles CO}_2 \text{ needed } n_{\text{CO}_2} = 0.0352 \text{ mol KO}_2 \left(\frac{2 \text{ mol CO}_2}{4 \text{ mol KO}_2} \right) = 0.0176 \text{ mol CO}_2$$

Number of moles CO_2 needed (0.0176 mol) < Number of moles present (0.102 mol)

There is sufficient CO_2 so the assumption is correct.

- a) Theoretical yield of oxygen:**

$$n_{\text{O}_2} = 0.0352 \text{ mol KO}_2 \left(\frac{3 \text{ mol O}_2}{4 \text{ mol KO}_2} \right) = 0.0264 \text{ mol O}_2$$

$$m_{\text{O}_2} = 0.0264 \text{ mol O}_2 \left(\frac{32.00 \text{ g}}{\text{mol}} \right) = 0.844 \text{ g O}_2$$

- b) Percentage yield**

$$\% \text{yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} 100\% = \frac{0.799 \text{ g}}{0.844 \text{ g}} 100\% = 94.7\%$$

14. a) Perform the following molar concentration calculations:

i) Calculate the molar concentration of 5.55 g CaCl_2 in 125 mL of solution.

Molar Mass: $\text{CaCl}_2 = (40.08) + 2(35.453) = 110.99$

$$n_{\text{CaCl}_2} = \frac{5.55 \text{ g}}{110.99 \text{ g} \cdot \text{mol}^{-1}} = 0.0500 \text{ mol}$$

$$V = 125 \text{ mL} \left(\frac{10^{-3} \text{ L}}{\text{mL}} \right) = 0.125 \text{ L}$$

$$[\text{CaCl}_2] = \frac{0.0500 \text{ mol}}{0.125 \text{ L}} = 0.400 \text{ M}$$

ii) Calculate the molar concentration of ammonium ion in a 0.333 M solution of ammonium phosphate.

Formula of ammonium phosphate $(\text{NH}_4)_3\text{PO}_4$

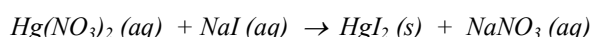
$$[\text{NH}_4^+] = \frac{0.333 \text{ mol } (\text{NH}_4)_3\text{PO}_4}{\text{L solution}} \left(\frac{3 \text{ mol NH}_4^+}{\text{mol } (\text{NH}_4)_3\text{PO}_4} \right) = 0.999 \text{ M}$$

b) Concentrated nitric acid is available as a 16 M solution. What volume of concentrated nitric acid must be diluted with distilled water to prepare 2.25 L of 0.10 M HNO_3 ?

Number of moles HNO_3 , $n_{\text{HNO}_3} = (0.10 \text{ M})(2.25 \text{ L}) = 0.225 \text{ mol}$

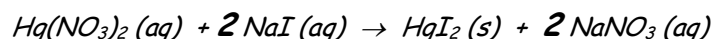
$$V_{\text{HNO}_3} = \frac{n_{\text{HNO}_3}}{[\text{HNO}_3]} = \frac{0.225 \text{ mol}}{16 \text{ M}} = 0.014 \text{ L} = 14 \text{ mL}$$

15. a) Given that 24.0 mL of 0.170 M sodium iodide reacts with 0.209 M mercury (II) nitrate according to the unbalanced equation:



- i) What volume of $\text{Hg}(\text{NO}_3)_2$ is required for complete precipitation of mercury in the form of HgI_2 ?
 ii) What is the mass of the precipitate formed?

$$n_{\text{NaI}} = (0.170\text{M})(0.0240\text{L}) = 4.08 \times 10^{-3} \text{ mol}$$



i)
$$n_{\text{Hg}(\text{NO}_3)_2} = 4.08 \times 10^{-3} \text{ mol} \left(\frac{1 \text{ mol Hg}(\text{NO}_3)_2}{2 \text{ mol NaI}} \right) = 2.04 \times 10^{-3} \text{ mol}$$

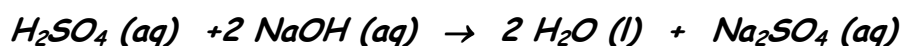
$$V_{\text{Hg}(\text{NO}_3)_2} = \frac{2.04 \times 10^{-3} \text{ mol}}{0.209 \text{ M}} = 9.76 \times 10^{-3} \text{ L} = 9.76 \text{ mL}$$

$$n_{\text{Hg}(\text{NO}_3)_2} = n_{\text{HgI}_2} = 2.04 \times 10^{-3} \text{ mol}$$

ii) **Molar Mass:** $\text{HgI}_2 = (200.59) + 2 \times (126.90) = 454.39$

$$m_{\text{HgI}_2} = 2.04 \times 10^{-3} \text{ mol} \left(\frac{454.39 \text{ g HgI}_2}{\text{mol}} \right) = 0.926 \text{ g}$$

- b) A volume of 39.48 mL of 0.115 M NaOH solution is required to neutralize completely a 10.00 mL sample of sulfuric acid. Calculate the concentration of the sulfuric acid solution.



$$n_{\text{NaOH}} = (0.115 \text{ M})(0.03948 \text{ L}) = 4.54 \times 10^{-3} \text{ mol}$$

$$n_{\text{H}_2\text{SO}_4} = 4.54 \times 10^{-3} \text{ mol NaOH} \left(\frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} \right) = 2.27 \times 10^{-3} \text{ mol}$$

$$[\text{H}_2\text{SO}_4] = \frac{2.27 \times 10^{-3} \text{ mol}}{0.01000 \text{ L}} = 0.227 \text{ M}$$

16. a) A 5.00 L sample of krypton gas contains 1.51×10^{24} atoms at 25°C . What is the pressure of the krypton gas in units of atm?

$$n_{\text{Kr}} = 1.51 \times 10^{24} \text{ atoms Kr} \left(\frac{\text{mol}}{6.022 \times 10^{23} \text{ atoms}} \right) = 2.51 \text{ mol}$$

$$V = 5.00 \text{ L}$$

$$T = 25 + 273 = 298 \text{ K}$$

$$P = \frac{nRT}{V} = \frac{(2.51 \text{ mol})(0.08206 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})(298 \text{ K})}{5.00 \text{ L}} = 12.3 \text{ atm}$$

- b) A sample of unknown gas weighs 1.95 g and occupies 3.00 L at 1.25 atm and 20°C . What is the molar mass of the unknown gas?

$$T = 20 + 273 = 293 \text{ K}$$

$$MM = \frac{mRT}{PV} = \frac{(1.95 \text{ g})(0.08206 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})(293 \text{ K})}{(1.25 \text{ atm})(3.00 \text{ L})} = 12.5 \text{ g} \cdot \text{mol}^{-1}$$